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

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

Is the salt acid, basic, or neutral?

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

This is the WEAK BASE ammonia. Ammonia is fairly stable in water.

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

The conjugate of either a strong acid or strong base is NEUTRAL

This is the STRONG ACID hydrochloric acid, It completely ionizes in water, meaning that the molecule doesn't exist in water - and that chloride ion is NOT a good proton acceptor.

We need to solve the following equilibrium:

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

This is the only reaction from the two above that will actually influence the pH (since the second one doesn't actually occur!)

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

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

Solve the equilbirum the same way we'd solve any other weak acid equilibrium

[Species]	[Init]	Δ	[Equil]
NH_3	0	+x	x
H_3O^+	0	+x	x
NH_4^+	0.100	-x	0.100-x

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

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

We find the Kb for ammonia in our data tables, but not the Ka for ammonium.

We can convert one to the other, though.

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

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

So, $K_a = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}}$

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

Assume x << 0.100

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

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

So, $\boxed{pH = 5.13}$

Compare:

pH = 1.00 for 0.1 M strong acid

pH = 7.00 for pure water

pH = 2.16 for 0.100 M nitrous acid

$0.100\ M\ NaC_2H_3O_2$, Find pH

$$NaC_2H_3O_2 \rightarrow Na^+ + C_2H_3O_2^-$$

Check the ions formed to see whether they're acidic or basic.

Na^+: Not a B-L acid (since it has no H to donate). Not likely to be a proton acceptor either (a B-L base) due ti the positive charge and the fact that it's a simple metal ion.

$C_2H_3O_2^-$: Has hydrogens, but also has a negative charge. Since it's negatively charged, it may also be a proton acceptor - a base.

Acetic acid - a WEAK ACID!

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

Acetic acid is stable in water, so we expect that the acetate ion can function as a proton acceptor.

$$K_{b,\,C_2H_3O_2^-} = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

Kb for acetate isn't in the chart at the back of the book, but Ka for acetic acid (its conjugate) is!

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

$$K_a \times K_b = 1.0 \times 10^{-14}, \text{ so } K_b = 5.88 \times 10^{-10}$$

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

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

Solve this just like solving for the pH of any other weak base:

Species	[Init]	Δ	[Equil]
$HC_2H_3O_2$	0	$+x$	x
OH^-	0	$+x$	x
$C_2H_3O_2^-$	0.100	$-x$	$0.100-x$

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

$x << 0.100$

$0.100 - x \approx 0.100$

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

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

We need pH ... but keep in mind that we just calculated HYDROXIDE concentration!

$pOH = -\log(7.67 \times 10^{-6}) = 5.12$

Since $pH + pOH = 14$, ...

$$\boxed{pH = 8.88}$$

Compare:
pH = 7.00 for pure water
pH = 11.13 for 0.100 M ammonia
pH = 13.00 for 0.100 M strong base (like NaOH)

0.100 M NaCl, Find pH

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

Check the ions formed to see if they;re acidic or basic:

Na^+: Sodium ion has no H, so it's not capable of being a B-L acid. Unlikely to be a B-L base due to the positive charge.

Cl^-: Chloride ion can't be a B-L acid (no H to donate). Might be a B-L base, since it's negatively charged and could conceivably attract protons.

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

This is a strong acid (which does not exist as a stable molecule in water)l. So, we conclude that chloride would not be a good proton acceptor.

Since neither sodium ion nor chloride ion affect the water equilibrium, the water equilibrium itself sets the pH, and the pH will be the same as the pH of pure water: 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.

Solving the equilibrium of phosphoric acid's first proton:

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

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

Species	[Init]	Δ	[Equil]
$H_2PO_4^-$	0	$+x$	x
H_3O^+	0	$+x$	x
H_3PO_4	0.10	$-x$	$0.10 - x$

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

This time, we'll solve with the quadratic equation. We're not as confident that 'x' is small... (compare 0.1 to Ka)

$$x^2 = 0.00069 - 0.0069x$$

$$x^2 + 0.0069x - 0.00069 = 0$$

$$a = 1 \quad b = 0.0069 \quad c = -0.00069$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

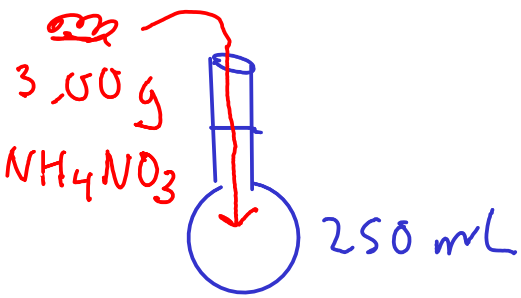
$$x = \frac{-0.0069 \pm 0.0529868852}{2}$$

Ignore the negative root ... negative values for x make no sense here...

$$x = 0.0230 = [H_3O^+]$$

$$\boxed{pH = 1.64}$$

171 Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate solid into enough water to make 250. mL of solution.



NH_4NO_3: N : 2 x 14.01
H : 4 x 1.008
O : 3 x 16.00

80.052 g NH_4NO_3 = 1 mol NH_4NO_3

Find out the nature of the salt. Acidic? Basic? Neutral?

$$NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$$

$$NO_3^- + H_2O \rightleftharpoons \underline{HNO_3} + OH^-$$

Nitric acid is a strong acid, so NITRATE ION should be NEUTRAL

$$NH_4^+ + H_2O \rightleftharpoons \underline{NH_3} + H_3O^+$$

Ammonia is a WEAK BASE, so AMMONIUM ION should be ACIDIC

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

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

$\underline{K_{a,NH_4^+}}$

$K_{b,NH_3} = 1.8 \times 10^{-5}$ (p A-14)

$(K_{a,NH_4^+})(1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$

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

$$\frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56\times10^{-10}$$

Species	[Init]	Δ	[Equil]
NH_3	0	$+x$	
H_3O^+	0	$+x$	
NH_4^+	0.14990	$-x$	

Calculate initial ammonium concentration

$$3.00\,g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.052\,g\ NH_4NO_3} =$$

$$= 0.037475640\,8\ mol\ NH_4NO_3$$

$$M = \frac{0.0374756408\ mol\ NH_4NO_3}{0.250\,L}$$

$$M = 0.14990\ \underline{M}$$

Solve...

$$\frac{x^2}{0.14990 - x} = 5.56\times10^{-10}$$

$x << 0.14990$

$$\frac{x^2}{0.14990} = 5.56\times10^{-10}$$

$x = 9.13\times10^{-6} = [H_3O^+]$

$$\boxed{pH = 5.04}$$

Seems okay for an acidic salt pH

THE COMMON-ION EFFECT

- is the effect on the ionization of a compound caused by the presence of an ion involved in the equilibrium

- is essentially Le Chateleir's Principle applied to equilibria involving ions

ex: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$; $K_b = 1.8 \times 10^{-5}$

From previous calculations, we know that an 0.10 M solution of ammonia has a pH of 11.13 .

What would happen to the pH if we dissolved ammonium chloride into the solution?

$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$

The ammonium chloride provides the ammonium ion. According to Le Chateleir's principle, this would shift the ammonia equilibrium to the LEFT!

What would happen to the pH? Let's find out!