$0.100\ M\ NaC_2H_3O_2$, Find pH

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

Check the ions formed by the salt to see if they have acidic or basic properties:

Na^+: Not a B-L acid, since it has no protons to donate. Not likely to be a B-L base, either. It's a simple alkali metal ion - positively charged.

$C_2H_3O_2^-$: Has hydrogens, but also a negative charge. It's probably a proton acceptor (base) rather than a donor due to the negative charge.

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

Acetic acid. Acetic acid is a WEAK ACID and stable in water, so we expect the ACETATE ION to be able to function as a base.

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

Kb for acetate ion is not available in our chart on page A-14, BUT we can find the Ka for acetic acid - its conjugate - on the previous page!

$K_{a,HC_2H_3O_2} = 1.7 \times 10^{-5}$; $K_a K_b = 1.0 \times 10^{-14}$; $K_b = 5.88 \times 10^{-10}$

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

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

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

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

$x << 0.100$

$0.100 - x \approx 0.100$

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

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

We need pH, but we've calculated HYDROXIDE concentration.

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

$pOH = 5.12$

Convert to pH:

$pH + pOH = 14.00$

$pH + 5.12 = 14.00$

$pH = 8.88$

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

0.100 M NaCl, Find pH

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

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

Na^+: Not a B-L acid, since it has no protons to donate. Not likely to be a B-L base, either. It's a simple alkali metal ion - positively charged.

Cl^-: Not a B-L acid, since it has no protons (H) to donate. Might be a proton acceptor due to its negative charge.

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

This is hydrochloric acid, a STRONG ACID. it completely ionizes in water, meaning that the chlloride ion must not be a good proton acceptor.

... so chloride ion should be NEUTRAL.

Since neither sodium ion nor chloride ion affect the water equilibrium, the pH of the solution will be the same as that of pure distilled 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:

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

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

Species	[Initial]	Δ	[Equilibrium]
H_3PO_4	0.10	$-x$	$0.10 - x$
$H_2PO_4^-$	0	$+x$	x
H_3O^+	0	$+x$	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'<<0.10 as we were in the previous examples.

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

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

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

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

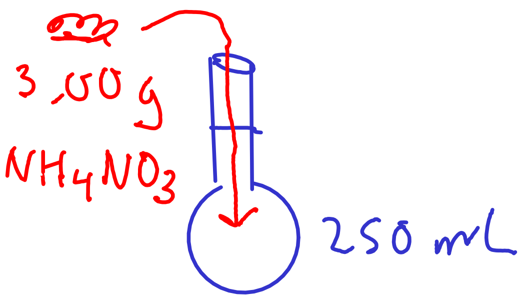
$$\frac{-0.0069 \pm 0.052986852}{2}$$

Discarding the negative root (does not make physical sense)...

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

$$pH = 1.64$$

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?

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

$$NO_3^- + H_2O \rightleftharpoons HNO_3 + OH^-$$

Nitric acid is a STRONG ACID, so NITRATE ION will be neutral.

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

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

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

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

K_{a,NH_4^+}:

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

$(K_{a,NH_4^+})(K_{b,NH_3}) = 1.0 \times 10^{-14}$

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

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

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	+x	x
H_3O^+	0	+x	x
NH_4^+	0.14990	−x	0.14990 − x

Calculate the initial ammonium nitrate concentration

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

$$= 0.0374756408\ mol\ NH_4NO_3$$

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

$$= 0.149902563\underline{3}\ \underline{M}$$

Solve:

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

$x << 0.14990$

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

$$x = 9.13 \times 10^{-6} = [H_3O^+]$$

$$pH = 5.04$$

This pH seems okay for an acidic salt.