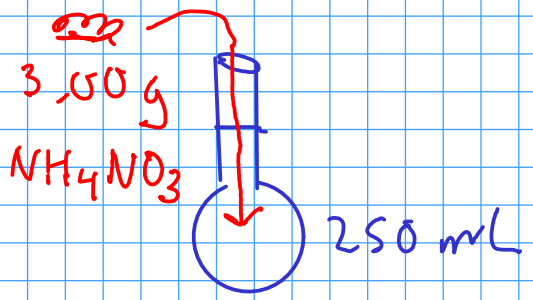
Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate solid into enough water to make 250. mL of solution.



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

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

✓ $NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$ Since ammonia is a weak base, its conugate (ammonium) is a weak acid!

X $NO_3^- + H_2O \rightarrow HNO_3 + OH^-$ Nitric acid is a STRONG acid, so it doesn't exist as the molecule in water. Nitrate ion, therefore, can't act as a base.

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

$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = ?$

$(K_{a,NH_4^+})(K_{b,NH_3}) = 1.0 \times 10^{-14}$ ($K_{b,NH_3} = 1.8 \times 10^{-5}$)

... conjugate pairs!

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

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

	initial	Δ	equilibrium
$[NH_3]$	0	$+x$	x
$[H_3O^+]$	0	$+x$	x
$[NH_4^+]$	0.1499	$-x$	$0.1499 - x$

We need to calculate the initial concentration of the ammonium ion!

$$3.00\,g\ NH_4NO_3 \times \frac{1\ mol\ NH_4NO_3}{80.052\,g\ NH_4NO_3} \times \frac{}{0.250\,L} = 0.1499\,\underline{M}$$

$$\underline{M} = \frac{mol}{L}$$

Now,, solve the equilibrium:

$$\frac{x^2}{0.1499 - x} = 5.56 \times 10^{-10}$$

$$\downarrow \quad x \ll 0.1499$$

$$\frac{x^2}{0.1499} = 5.56 \times 10^{-10} \longrightarrow x = 9.13 \times 10^{-6} = [H_3O^+]$$

$$\boxed{pH = 5.04}$$

BUFFERS

- resist pH change caused by either the addition of strong acid/base OR by dilution

Made in one of two ways:

1. Make a mixture of a weak acid and its conjugate base (as the SALT) — $HC_2H_3O_2$ / $NaC_2H_3O_2$
2. Make a mixture of a weak base and its conjugate acid (as the SALT) — NH_3 / NH_4Cl

For a weak acid, you would:

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

- Add HA (weak acid)
- Add a salt containing A^- (example: NaA)

- This solution actually contains an acid and a base at equilibrium, with a significant concentration of BOTH.

- The acid in the buffer can neutralize bases, while the base can neutralize acids.

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

1. Take log of both sides
2. Multiply by -1
3. Rearrange, solving for pH

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

Henderson-Hasselbalch Equation

$[A^-]$ ... from the salt

$[HA]$ ... from the weak acid

- We ASSUME that the initial concentrations of both the acid and its conjugate are equal to the equilibrium concentrations. Valid IF there are significant amounts of both species initially.

$$pH = pK_{a,\,acidic} + \log\left(\frac{[\text{basic species}]}{[\text{acidic species}]}\right)$$

Henderson-Hasselbalch Equation

ex: acidic buffer

$HC_2H_3O_2 \,/\, NaC_2H_3O_2$

$$pH = pK_{a,\,HC_2H_3O_2} + \log\left(\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}\right)$$

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ex: basic buffer

$NH_3 \,/\, NH_4NO_3$

$$pH = pK_{a,\,NH_4^+} + \log\left(\frac{[NH_3]}{[NH_4^+]}\right)$$

$pK_a + pK_b = 14{,}00$ ... is the -log of $K_a \times K_b = K_w$

Calculate the pH of a buffer made from 3.2 grams of ammonium chloride and 29 mL of 18.1 M ammonia diluted to 50.0 mL with water.

$$pH = pK_{a,\,acidic} + \log\left(\frac{[\text{basic species}]}{[\text{acidic species}]}\right)$$ — Henderson-Hasselbalch Equation

$[NH_3] = ?$ Dilution! $\quad M_1V_1 = M_2V_2$

$(18.1\,M)(29\,mL) = M_2(50.0\,mL)$

$M_2 = [NH_3] = 10.498\ \underline{M}$

$[NH_4^+] = ?$ $\quad NH_4Cl$: 53.492 g/mol

$$3.00\,g\,NH_4Cl \times \frac{mol\ NH_4Cl}{53.492\,g\,NH_4Cl} \times \frac{}{0.0500\,L} = 1.12\ \underline{M}$$

$pK_{a,NH_4^+} = -\log(5.56\times10^{-10}) = 9.25$ ← from a previous problem and $K_a \times K_b = K_w$

$$pH = 9.25 + \log\left(\frac{10.498}{1.12}\right) = 10.22 = pH$$

$\log(10.498/1.12) = 0.97$