143 Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

Salt: Look at the ions this salt will make when it dissolves. $NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$

$NH_4^+$: $NH_4^+ + H_2O \rightleftharpoons \boxed{NH_3} + H_3O^+$ Possibly a proton donor (acid) ...

weak base. stable in $H_2O$. ✓

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

strong acid. Not stable in $H_2O$. ✗

Since the ammonium ion can dontate a proton, but the nitrate ion can't either donate or accept, the equilibrium we'll need to solve is ammonium ion's ACID IONIZATION. We will need a Ka expression for ammonium ion, a value for the Ka, and the initial MOLAR concentrations of the species in ammonium's equilibrium.

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

$K_{b,NH_3} = 1.8 \times 10^{-5}$ (p1249, 05)

$K_a K_b = K_w$

$K_a (1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$

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

$[NH_4NO_3]_{nominal}$:

$$3.00\,g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.052\,g\ NH_4NO_3} = 0.03747564 08\ mol\ NH_4NO_3$$

$$\frac{0.0374756408\ mol\ NH_4NO_3}{0.250\,L} = 0.149025633\ M \approx 0.150\ M\ NH_4NO_3$$

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ $\quad K_{a,NH_4^+} = \dfrac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56\times10^{-10}$

| Species | [Initial] | Δ | [Equilibrium] |
|---|---|---|---|
| $H_3O^+$ | 0 | +x | x |
| $NH_3$ | 0 | +x | x |
| $NH_4^+$ | 0.150 | −x | 0.150 − x |

Let "x" equal the change in hydronium ion concentration.

$$\frac{(x)(x)}{(0.150-x)} = 5.56\times10^{-10}$$

$$\frac{x^2}{0.150-x} = 5.56\times10^{-10}$$

$x << 0.150$

$0.150 - x \approx 0.150$

$$\frac{x^2}{0.150} = 5.56\times10^{-10}$$

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

$pH = \boxed{5.04}$ $\quad pH = -\log_{10}[H_3O^+]$

# 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!

Calculate the pH of a solution which contans 0.10 M ammonia AND 0.10 M ammonium chloride.

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ ; $K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$

| Species | [Initial] | Δ | [Equilibrium] |
|---|---|---|---|
| $OH^-$ | 0 | +x | x |
| $NH_4^+$ | 0.10 | +x | 0.10 + x |
| $NH_3$ | 0.10 | −x | 0.10 − x |

Let "x" equal the change in hydroxide ion concentration.

$$\frac{(0.10+x)(x)}{(0.10-x)} = 1.8 \times 10^{-5}$$

$x << 0.10$, so

$0.10 - x \approx 0.10$

$0.10 + x \approx 0.10$

$$\frac{(0.10)(x)}{0.10} = 1.8 \times 10^{-5}$$

$x = 1.8 \times 10^{-5} = [OH^-]$

$pOH = 4.74$ $\quad pOH = -\log_{10}[OH^-]$

$pH + pOH = 14.00$

$pH + 4.74 = 14.00$

$pH = \boxed{9.26}$

# BUFFERS

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

Made in one of two ways:

$HC_2H_3O_2$ $NaC_2H_3O_2$

1. Make a mixture of a weak acid and its conjugate base (as the SALT)

$NH_3$ $NH_4Cl$

2. Make a mixture of a weak base and its conjugate acid (as the SALT)

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.