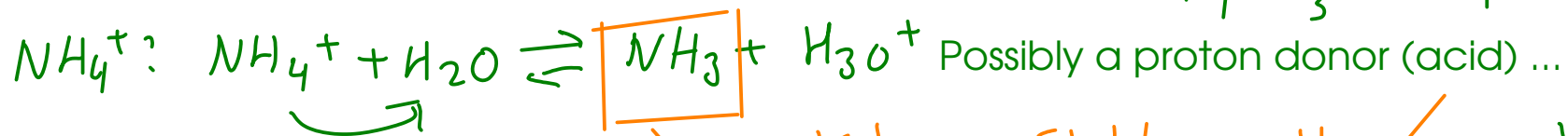
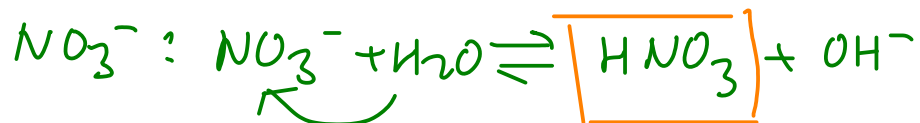


143 Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

Salt: Look at the ions this salt will make when it dissolves. $\text{NH}_4\text{NO}_3 \rightarrow \text{NH}_4^+ + \text{NO}_3^-$



weak base. stable in H_2O . ✓



strong acid, not stable in H_2O . ✗

Since the ammonium ion can donate 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 K_a expression for ammonium ion, a value for the K_a , and the initial MOLAR concentrations of the species in ammonium's equilibrium.

$$K_{a, \text{NH}_4^+} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = ?$$

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5} \quad (\text{p}1249, 05)$$

$$K_a K_b = K_w$$

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

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

$[\text{NH}_4\text{NO}_3]_{\text{nominal}}$:

$$3.00 \text{ g } \text{NH}_4\text{NO}_3 \times \frac{\text{mol } \text{NH}_4\text{NO}_3}{80.052 \text{ g } \text{NH}_4\text{NO}_3} = 0.0374756408 \text{ mol } \text{NH}_4\text{NO}_3$$

$$\frac{0.0374756408 \text{ mol } \text{NH}_4\text{NO}_3}{0.250 \text{ L}} = 0.1499025633 \text{ M} \approx 0.150 \text{ M } \text{NH}_4\text{NO}_3$$

0.250 L



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.

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

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

$$\downarrow \begin{array}{l} x \ll 0.150 \\ 0.150-x \approx 0.150 \end{array}$$

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

$$x = 9.13 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = \boxed{5.04}$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

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 Chateleur's Principle applied to equilibria involving ions



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

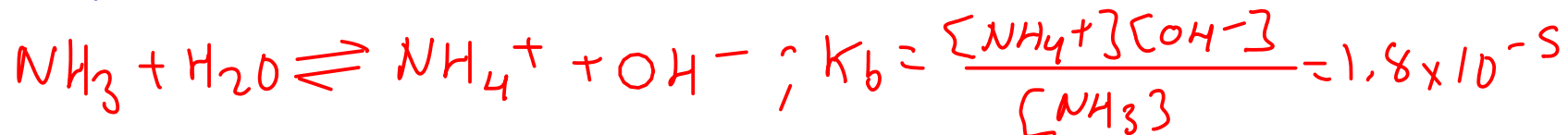
What would happen to the pH if we dissolved ammonium chloride into the solution?



The ammonium chloride provides the ammonium ion. According to Le Chateleur'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 contains 0.10 M ammonia AND 0.10 M ammonium chloride.



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.

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

$$\begin{aligned} & \downarrow x \ll 0.10, 50 \\ & 0.10 - x \approx 0.10 \\ & 0.10 + x \approx 0.10 \end{aligned}$$

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

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

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

$$\swarrow \text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 4.74 = 14.00$$

$$\text{pH} = \boxed{9.26}$$

BUFFERS

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

Made in one of two ways:

- ① Make a mixture of a weak acid and its conjugate base (as the SALT)
 $\text{HC}_2\text{H}_3\text{O}_2$ $\text{NaC}_2\text{H}_3\text{O}_2$
- 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:



- 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.