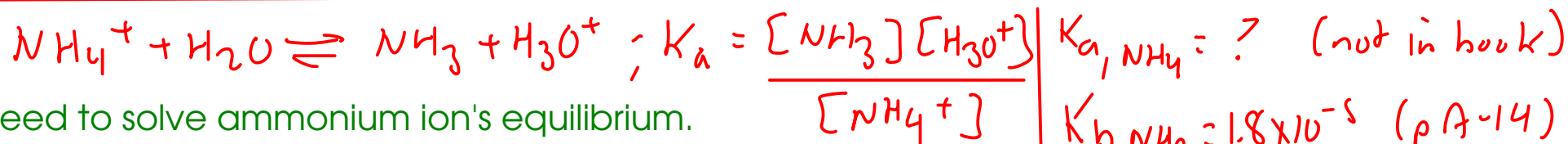
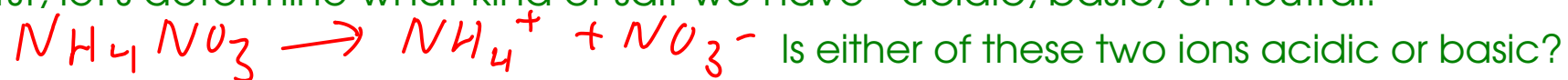


163 Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

First, let's determine what kind of salt we have - acidic, basic, or neutral:



$K_a, \text{NH}_4 = ?$  (not in book)

$K_b, \text{NH}_3 = 1.8 \times 10^{-5}$  (pA-14)

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

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

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

We need to solve ammonium ion's equilibrium.

First, find  $K_a$  (you'll need to get it from the  $K_b$  of the conjugate ---->)

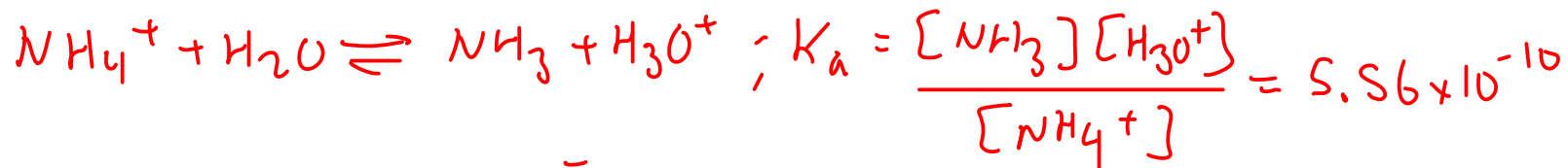
We need to set up an equilibrium chart, but those concentrations are in molarity units. So let's get the initial ammonium nitrate concentration in molarity units!

$$3.00 \text{ g NH}_4\text{NO}_3 \times \frac{\text{mol NH}_4\text{NO}_3}{80.052 \text{ g NH}_4\text{NO}_3} = 0.0374756400 \text{ mol NH}_4\text{NO}_3$$

$$\frac{0.0374756400 \text{ mol NH}_4\text{NO}_3}{0.250 \text{ L}} = 0.1499025633 \text{ M NH}_4\text{NO}_3$$

0.250 L  
(250 mL)

$$= 0.1499025633 \text{ M NH}_4^+$$



0.1499025633 M  $\text{NH}_4^+$  ← concentration of acid

Set up equilibrium chart and solve!

Species	[Initial]	$\Delta$	[Equilibrium]
$\text{H}_3\text{O}^+$	0	+ X	X
$\text{NH}_3$	0	+ X	X
$\text{NH}_4^+$	0.1499025633 M	- X	0.1499025633 - X

Let "x" equal the change in hydronium ion concentration

$$\frac{(x)(x)}{(0.1499025633 - x)} = 5.56 \times 10^{-10}$$

$$\frac{x^2}{0.1499025633 - x} = 5.56 \times 10^{-10}$$

Assume  $x \ll 0.1499$

$$\frac{x^2}{0.1499025633} = 5.56 \times 10^{-10}$$

$$x = 9.129393475 \times 10^{-6} \text{ M } \text{H}_3\text{O}^+$$

$$\text{pH} = \boxed{5.04} \quad (\text{pH} = -\log_{10} [\text{H}_3\text{O}^+])$$

## 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 .

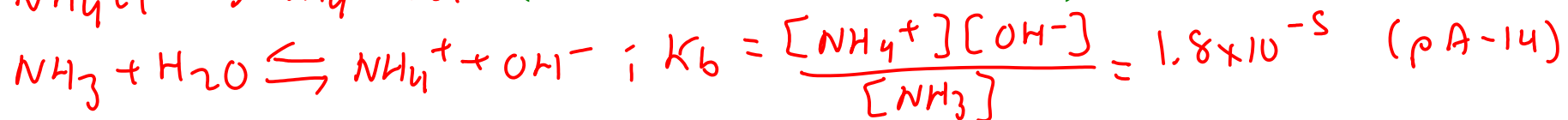
What would happen to the pH if we dissolved ammonium chloride into the solution?



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.



Set up and solve the ammonia equilibrium...

Species	[Initial]	$\Delta$	[Equilibrium]
$\text{OH}^-$	0	+x	x
$\text{NH}_4^+$	0.10	+x	0.10 + x
$\text{NH}_3$	0.10	-x	0.10 - x

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

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

Assume  $x \ll 0.10$   
 $0.10 - x \approx 0.10$   
 $0.10 + x \approx 0.10$

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

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

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

$$\boxed{\text{pH} = 9.26} \quad (\text{pH} + \text{pOH} = 14.00)$$

(\*Agrees with our expectations on what the pH would be based on the common ion effect - the pH should be more acidic than before, and it is...)

## BUFFERS

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

Made in one of two ways:

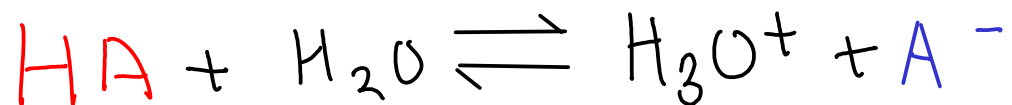
- ① Make a mixture of a weak acid and its conjugate base (as the SALT)  
 $\text{HC}_2\text{H}_3\text{O}_2$                        $\text{NaC}_2\text{H}_3\text{O}_2$
- ② Make a mixture of a weak base and its conjugate acid (as the SALT)  
 $\text{NH}_3$                                        $\text{NH}_4\text{Cl}$

For a weak acid, you would:



- Add HA (weak acid)
- Add a salt containing  $\text{A}^-$  (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.



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

① Take log of both sides

② Multiply by -1

③ Rearrange, solving for pH

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) \quad \left| \begin{array}{l} \text{Henderson-} \\ \text{Hasselbalch} \\ \text{Equation} \end{array} \right.$$

$[A^-]$  ... 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.