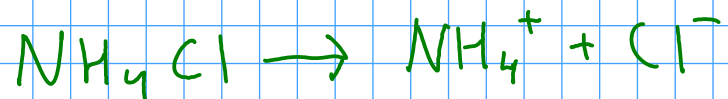
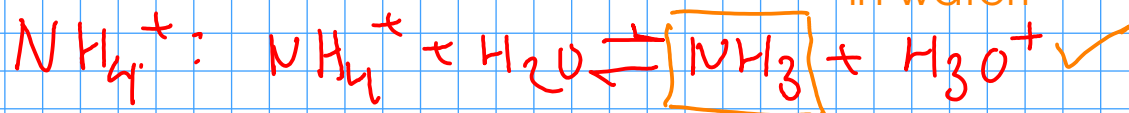


0.100 M NH_4Cl ... Find the pH of the solution



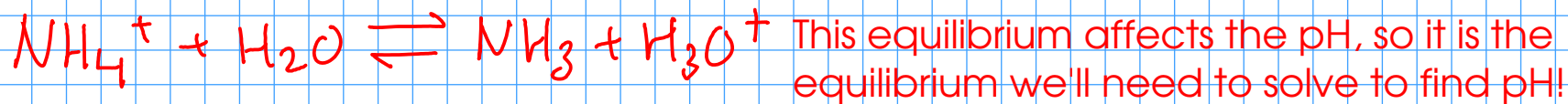
Acidic, basic, or neutral salt?

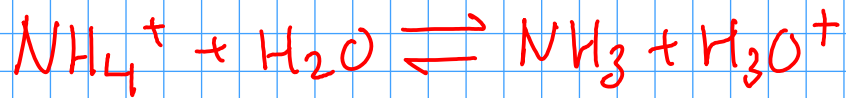
This is the WEAK BASE ammonia. Stable in water.



This is a STRONG ACID, which does not exist as a stable molecule in water.

The conjugate of a strong acid or base is NEUTRAL - does not affect pH!





$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]} = 5.56 \times 10^{-10}$$

Where do we get this K_a ?
 $K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$
 $K_a \times K_b = 1.0 \times 10^{-14}$

	initial	Δ	equilibrium
$[\text{NH}_3]$	0	+x	x
$[\text{H}_3\text{O}^+]$	0	+x	x
$[\text{NH}_4^+]$	0.100	-x	0.100 - x

$$\frac{x^2}{0.100 - x} \approx 5.56 \times 10^{-10}$$

$x \ll 0.100$

$$\frac{x^2}{0.100} \approx 5.56 \times 10^{-10}$$

$$x = 7.46 \times 10^{-6}$$

We defined "x" as the concentration of hydronium ion made by the equilibrium!

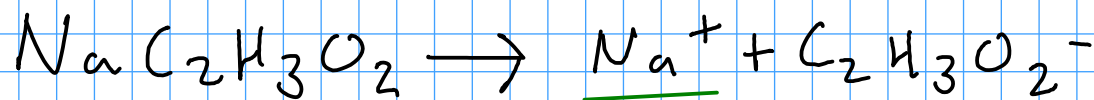
$$x = 7.46 \times 10^{-6} \approx [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log_{10}(7.46 \times 10^{-6}) = 5.13 = \text{pH}$$

Compare:

- pH = 1.00 for 0.100 M strong acid
- pH = 2.16 for 0.100 M nitrous acid
- pH = 7.00 for pure distilled water

0.100 M $\text{NaC}_2\text{H}_3\text{O}_2$, Find pH



conjugate of acetic acid
 $\text{HC}_2\text{H}_3\text{O}_2$

neutral



Since acetic acid is a weak acid and stable in water, we would expect acetate ion to be BASIC.

$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = ?$$

$$K_{a, \text{HC}_2\text{H}_3\text{O}_2} = 1.7 \times 10^{-5}; \quad K_a \times K_b = 1.0 \times 10^{-14} \quad \dots \text{true for conjugate pairs}$$

$$K_b = 5.88 \times 10^{-10}$$



$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5,88 \times 10^{-10}$$

	Initial	Δ	equilibrium
$[\text{HC}_2\text{H}_3\text{O}_2]$	0	+x	x
$[\text{OH}^-]$	0	+x	x
$[\text{C}_2\text{H}_3\text{O}_2^-]$	0,100	-x	0,100 - x

Define "x" as the concentration of hydroxide produced by the equilibrium, since that's closely related to pH

$$\frac{x^2}{0,100 - x} = 5,88 \times 10^{-10}$$

$x \ll 0,100$

$$\frac{x^2}{0,100} = 5,88 \times 10^{-10} ; x = 7,67 \times 10^{-6}$$

$$x = 2.67 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = \log_{10}(2.67 \times 10^{-6})$$

$\text{pOH} = 5.12$... this is pOH, but we need pH. Luckily for us, they are related very simply.

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

$$\text{so, } \text{pH} = 14.00 - 5.12$$

$$\text{pH} = \boxed{8.88}$$

Compare:

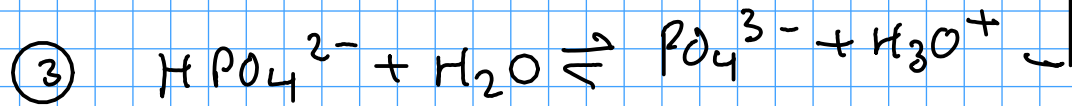
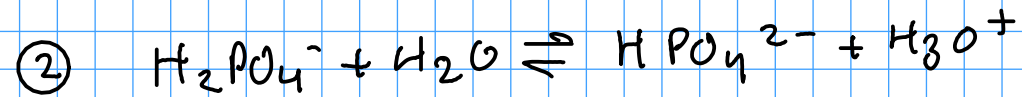
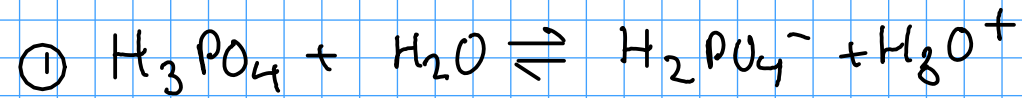
pH = 13.00 for 0.100 M strong base

pH = 7.00 for pure distilled water

POLYPROTIC ACIDS

Find pH of 0,10 M H_3PO_4

... what's special about phosphoric acid?



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 dissociation 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:



$$K_a = 6.9 \times 10^{-3} = \frac{[\text{H}_2\text{PO}_4^-][\text{H}_3\text{O}^+]}{[\text{H}_3\text{PO}_4]}$$

	initial	Δ	equilibrium
$[\text{H}_2\text{PO}_4^-]$	0	+x	x
$[\text{H}_3\text{O}^+]$	0	+x	x
$[\text{H}_3\text{PO}_4]$	0.10	-x	0.10 - x

$$\frac{x^2}{0.10 - x} = 6.9 \times 10^{-3}$$

For this problem, the assumption that $x \ll 0.10$ might not be safe (K_a is fairly large for a weak acid), so we should solve the quadratic instead...

$$x = 0.023043 \text{ or } -0.029943$$

$$[\text{H}_3\text{O}^+] = 0.023043 \text{ M}$$

$$\text{pH} = 1.64$$

Less than 0.10, but not MUCH less!