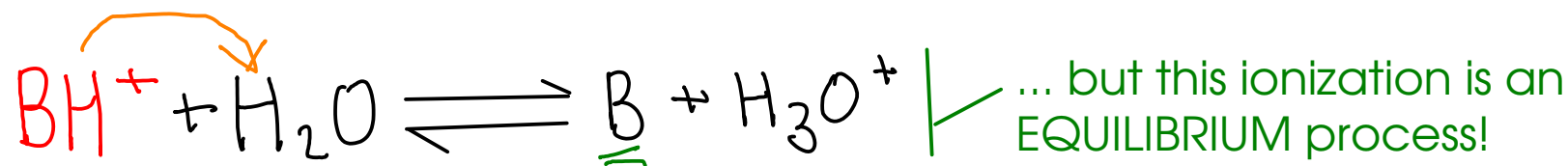
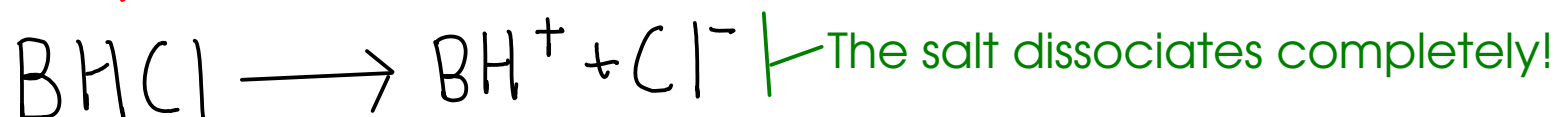


SALT OF A WEAK BASE

ex: NH_4Cl 

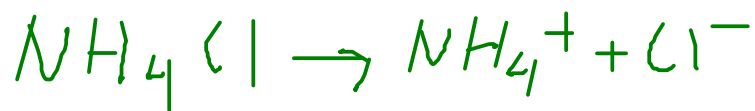
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for } \text{BH}^+ \right.$$

$$K_w = (K_{a, \text{BH}^+})(K_{b, \text{B}})$$

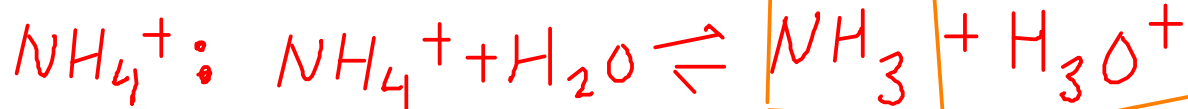
1.0×10^{-14}

Find the pH for salt solutions just like you would find pH for any other weak acid or weak base solutions. Only trick is to find out whether the salt is actually acidic or basic!

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



Is the salt acidic, basic, or neutral? Look at the IONS the salt releases when it dissolves!



Since ammonia is a WEAK base, that means it will exist in water as molecules ... meaning the reaction we wrote can occur and ammonium ion can act as an acid.



HCl is a strong acid. Since it completely ionizes, that means chloride ion isn't able to keep a proton attached, and isn't a base.

So the equilibrium we'll need to look at is the one with the acidic ammonium ion.



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

$K_{a, \text{NH}_4^+} = ?$ We'll need to find K_a for ammonium by looking at K_b for its conjugate, ammonia.

$$K_{b, \text{NH}_3} = 1.8 \times 10^{-5} \quad (\text{p}K_b = 14)$$

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

$$K_{a, \text{NH}_4^+} = 5.56 \times 10^{-10}$$



Set up and solve this equilibrium like any other weak acid problem!

Species	[Initial]	Δ	[Equilibrium]
NH_3	0	+x	x
H_3O^+	0	+x	x
NH_4^+	0.100	-x	0.100 - x

Let "x" equal the change in ammonia concentration...

$$\frac{(x)(x)}{(0.100 - x)} = 5.56 \times 10^{-10}$$

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

$$\begin{array}{l} \downarrow x \ll 0.100, \\ \text{so } 0.100 - x \approx 0.100 \end{array}$$

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

$$x = 7.45 \times 10^{-6} \text{ M} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log_{10}(7.45 \times 10^{-6})$$

$$\boxed{\text{pH} = 5.13}$$

For comparison,

pH = 1.00 for a strong acid like HCl

pH = 2.17 for 0.10 M nitrous acid

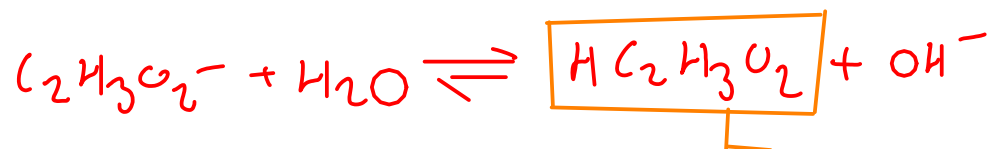
pH = 7.00 for distilled water

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



Na^+ : Cannot be B-L acid, as it's missing an H^+ to donate. Has a positive charge, so unlikely to be a B-L base (would repel proton). Neutral.

$\text{C}_2\text{H}_3\text{O}_2^-$: Has hydrogens, but also has a negative charge. Negative charge is attractive to protons, so a base?



This is ACETIC ACID, a WEAK acid. Since acetic acid is weak (and stable in water), we expect the acetate ion to be BASIC.

So, we need to solve the equilibrium of acetate ion reacting as a base with water.



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

$$K_{a, \text{HC}_2\text{H}_3\text{O}_2} = 1.7 \times 10^{-5}$$

Since $K_a \times K_b = K_w$

$$(1.7 \times 10^{-5}) K_b = (1.0 \times 10^{-14})$$

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

(data from pages A-13/A-14 in Ebbing)



(0.100 M $\text{Na}(\text{C}_2\text{H}_3\text{O}_2)$)

$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5.88 \times 10^{-10}$$

Set up an equilibrium chart and solve

Species	[Initial]	Δ	[Equilibrium]
$\text{HC}_2\text{H}_3\text{O}_2$	0	+x	x
OH^-	0	+x	x
$\text{C}_2\text{H}_3\text{O}_2^-$	0.100	-x	0.100 - x

Let "x" equal the change in acetic acid concentration.

$$\frac{(x)(x)}{(0.100 - x)} = 5.88 \times 10^{-10}$$

Assume $x \ll 0.100$,
so $0.100 - x = 0.100$

$$\frac{x^2}{0.100} = 5.88 \times 10^{-10}$$

$$x = 7.6696 \times 10^{-6} = [\text{OH}^-]$$

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

$$\text{pOH} = -\log_{10}(7.6696 \times 10^{-6}) = 5.12$$

$$\text{pH} = 14 - 5.12 = \boxed{8.88}$$

$$\curvearrowright (\text{pH} + \text{pOH} = 14.00)$$

For comparison:

0.100 M sodium acetate, pH = 8.88

0.100 M ammonia, pH = 11.13

0.100 M NaOH (strong base), pH = 13.00

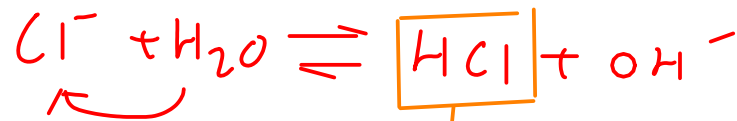
The acetate ion is basic, but it's a very weak base!

0.100 M NaCl, Find pH



Na^+ ? Neither a B-L acid nor base (no protons to donate, and the positive charge hinders acceptance of a proton). Neutral.

Cl^- ? Base?



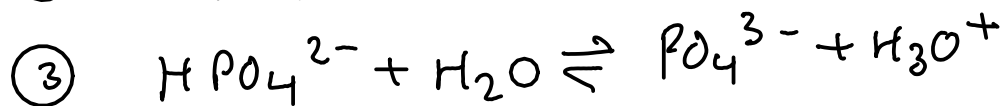
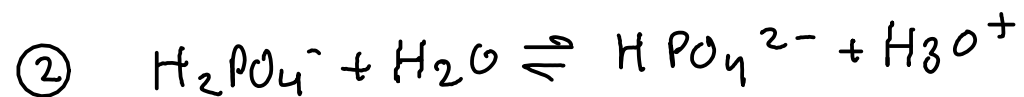
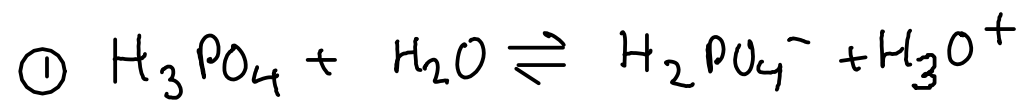
STRONG ACID, so chloride ion will not be a proton acceptor!

... so chloride ion is also neutral.

In this case, the pH of the solution is controlled by water's own ionization, so the solution would be pH = 7.00.

Find pH of 0.10 M H_3PO_4

... what's special about phosphoric acid?



Phosphoric acid has THREE acidic protons!

$$K_{a1} = 6.9 \times 10^{-3}$$

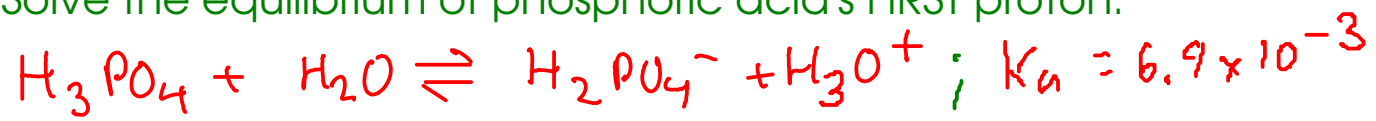
$$K_{a2} = 6.2 \times 10^{-8}$$

$$K_{a3} = 4.8 \times 10^{-13}$$

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.

Solve the equilibrium of phosphoric acid's FIRST proton:



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

Species	[Initial]	Δ	[Equilibrium]
H_2PO_4^-	0	+x	x
H_3O^+	0	+x	x
H_3PO_4	0.10	-x	0.10 - x

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

assume $x \ll 0.10$
so $0.10 - x \approx 0.10$

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

$$x = 0.0262678511 = [\text{H}_3\text{O}^+]$$

$$\text{So, } \text{pH} = -\log(0.0262678511)$$

$$\text{pH} = 1.58$$