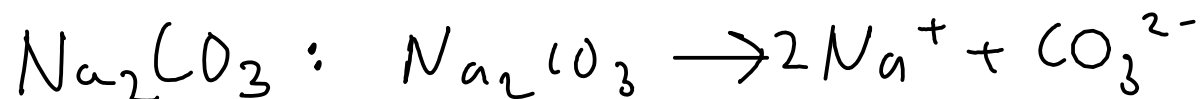


SALTS

- Compounds that result from the reaction of an acid and a base.
- Salts are strong electrolytes (completely dissociate in water) IF SOLUBLE (not all salts dissolve appreciably).
- Most ionic compounds are considered salts (they can be made by some reaction between the appropriate acid and base)
- Salts have acidic and basic properties! The ions that form when salts are dissolved can be acidic, basic, or neutral.
- Salts made from WEAK ACIDS tend to form BASIC solutions
- Salts made from WEAK BASES tend to form ACIDIC solutions



Do any of these ions have acidic or basic properties?

Na^+ : neutral. Not a proton donor or a proton acceptor

CO_3^{2-} : BASIC, since it can accept protons to form the weak acid CARBONIC ACID in solution.



ACID

BASE

SALT OF A WEAK ACID

ex: $\text{NaC}_2\text{H}_3\text{O}_2$ 

For this reaction to occur, HA MUST be stable in water. In other words, a weak acid.



The anion is a BASE. It can accept a proton from water to form the weak (therefore stable as a molecule!) acid HA

$$K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]} \quad \left| \text{--- This is the base ionization constant for } \text{A}^- \right.$$

Since A^- and HA are a conjugate pair, the ionization constants are related!

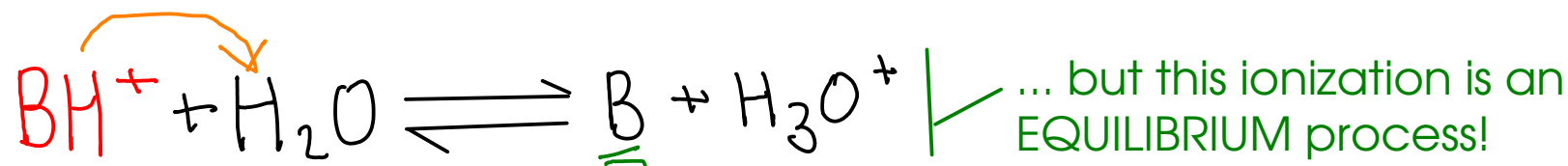
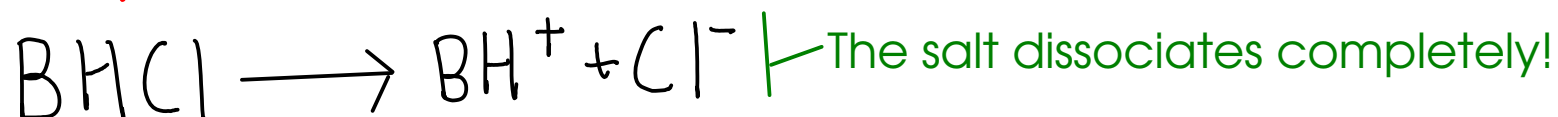
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

1.0×10^{-14}

$$14 = \text{p}K_a + \text{p}K_b$$

You will generally not find both the K_a AND K_b for a conjugate pair in the literature, since one can be easily converted to the other!

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

Find the ions produced by the salt. Are they acids or bases? $\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$

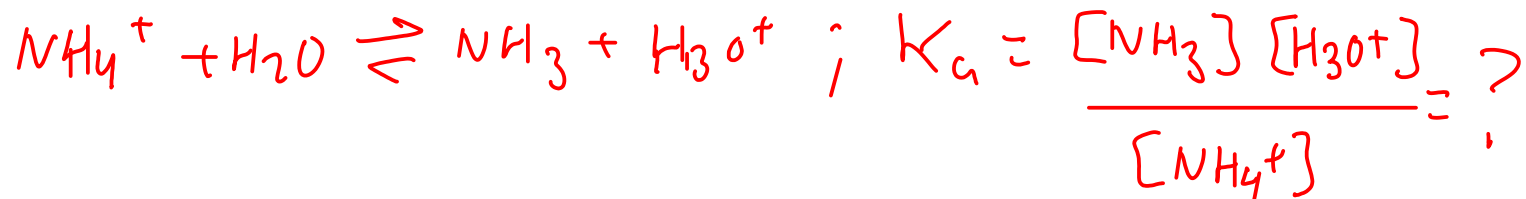


↳ This is ammonia, a WEAK BASE. It's water-stable, so this reaction is viable.



↳ This is hydrochloric acid, a STRONG ACID. It's NOT water-stable, since it fully ionizes in water. This reaction won't go.

We need to solve the equilibrium of the ammonium ion.



We can't find the K_a for ammonium ion in OpenStax Appendix H. We can, however, find the K_b of its conjugate (ammonia) in Appendix I:

$$K_b, \text{NH}_3 = 1.8 \times 10^{-5} \quad K_a \times K_b = 1.0 \times 10^{-14}$$

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

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



Now that we have K_a , solve as a normal weak acid.

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

Let "x" equal the change in hydronium ion concentration

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

$$x \ll 0.100, \text{ so}$$

$$0.100 - x \approx 0.100$$

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

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

$$\text{pH} = -\log_{10}(7.45 \times 10^{-6}) = \boxed{5.12}$$

Compare:

pH of 0.100 M nitric acid: 1.00

pH of 0.100 M nitrous acid: 2.17

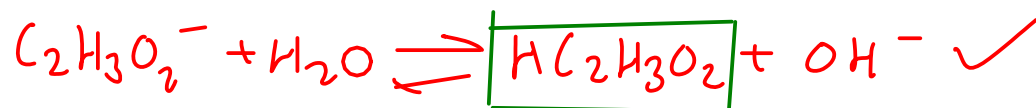
pH of DI water: 7.00

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



Na^+ : Not a B-L acid (no H^+ to donate), nor is it basic ... the positive charge should REPEL protons!

$\text{C}_2\text{H}_3\text{O}_2^-$: Does have hydrogens, but also has a negative charge. Let's see if this is basic:



Acetic acid (weak!)

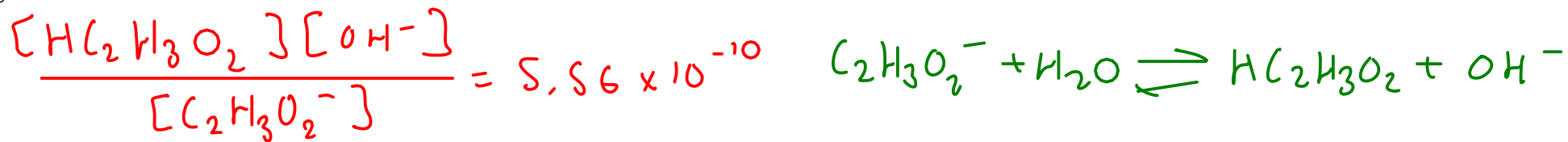
Solve equilibrium $\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$; $K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]}$

Look at OpenStax appendicies: Chart doesn't have acetate ion, but DOES have acetic acid!

$$K_a, \text{HC}_2\text{H}_3\text{O}_2 = 1.8 \times 10^{-5}; (1.8 \times 10^{-5}) K_b = 1.0 \times 10^{-14} \quad (K_a K_b = K_w)$$

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

$$\frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5.56 \times 10^{-10}$$



Solve like any other base!

Species	[Initial]	Δ	[Equilibrium]
OH^-	0	+x	x
$HC_2H_3O_2$	0	+x	x
$C_2H_3O_2^-$	0.100	-x	0.100 - x

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

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

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

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

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

$$x = 7.45 \times 10^{-6} = [OH^-]$$

$$pOH = 5.13$$

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

$$pH = 14.00 - 5.13 = \boxed{8.87} \quad pH + pOH = 14.00$$

For comparison:

0.100 M sodium acetate, pH = 8.87

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^+ : Not a B-L acid (no H^+ to donate), nor is it basic ... the positive charge should REPEL protons!

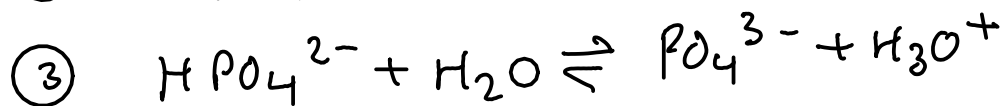
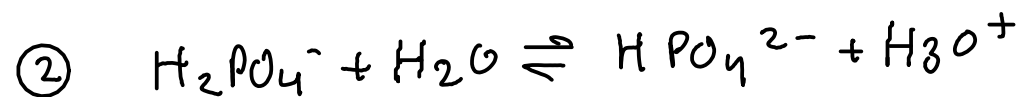
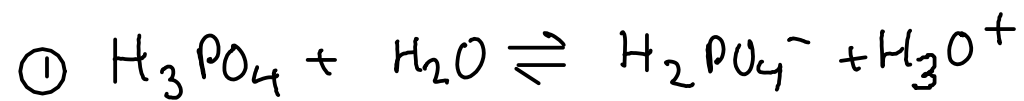
Cl^- ; Not a B-L acid (no H^+ to donate). Base? $\text{Cl}^- + \text{H}_2\text{O} \rightleftharpoons \boxed{\text{HCl}} + \text{OH}^-$

Strong base, so this is NOT stable in water. Cl^- cannot hold onto the proton!

Since neither ion acts as an acid or base, then water's own ionization controls pH. The pH is 7.00.

Find pH of 0.10 M H_3PO_4

... what's special about phosphoric acid?



Phosphoric acid has **THREE** acidic protons!

$$K_{a1} = 7.5 \times 10^{-3}$$

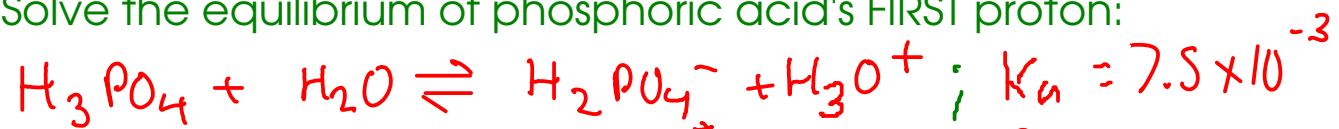
$$K_{a2} = 6.2 \times 10^{-8}$$

$$K_{a3} = 4.2 \times 10^{-13}$$

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:



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{H}_2\text{PO}_4^-]}{[\text{H}_3\text{PO}_4]} = 7.5 \times 10^{-3}$$

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

Let 'x' equal the change in hydronium ion concentration

$$\frac{(x)(x)}{(0.10-x)} = 7.5 \times 10^{-3}$$

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

$$\frac{x^2}{0.10} = 7.5 \times 10^{-3}$$

$$x = 0.0273861279 = [\text{H}_3\text{O}^+]$$

$$\text{pH} = 1.56$$

Check this in experiment 16A. You'll measure the pH of this same concentration of phosphoric acid yourself. It may be slightly lower, since we ignored the other two acid ionizations.