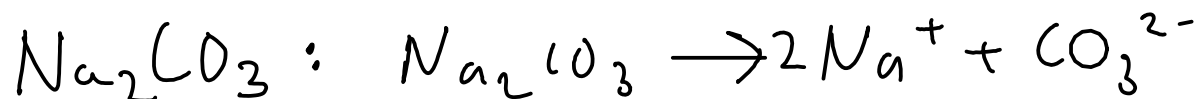


SALTS

- 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!

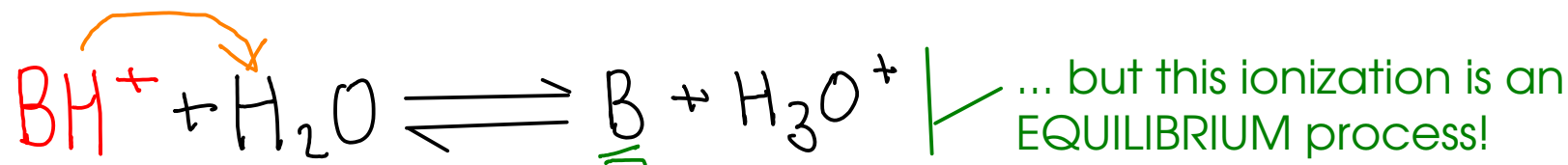
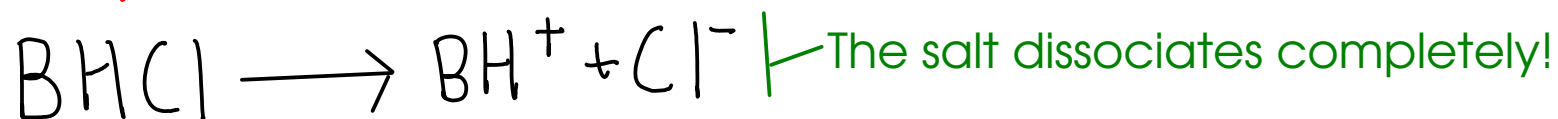
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

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 

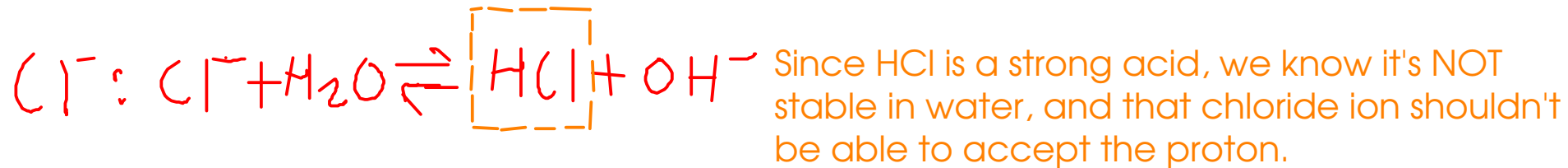
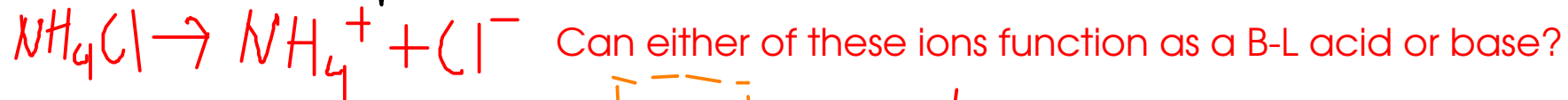
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for } \text{BH}^+ \right.$$

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

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



So, to figure out the pH of the solution, we need to look at ammonium ion's equilibrium $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$; $K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$

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

$$\frac{(x)(x)}{(0.100 - x)} = K_a$$

$$\frac{(x)(x)}{(0.100-x)} = K_a$$

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

$$\downarrow \begin{array}{l} x \ll 0.100, 50 \\ 0.100-x \approx 0.100 \end{array}$$

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

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

$$\text{pH} = 5.12$$

We need K_a for ammonium ion. No value for this K_a on page A-13, but on page A-14 we find the K_b for ammonia, the conjugate of ammonium ion... $K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$

$$K_a \times K_b = K_w$$

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

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

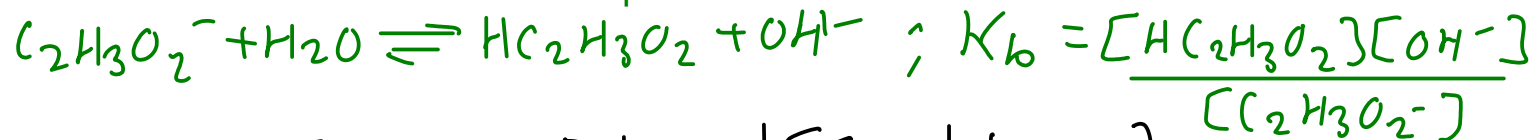
0.100 M $\text{NaC}_2\text{H}_3\text{O}_2$, Find pH



Na^+ : Neutral. No proton (can't be acid), positive charge (not likely to attract H^+)

$\text{C}_2\text{H}_3\text{O}_2^-$: base? $\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$ Acetic acid is WEAK (meaning it's stable in water)... Acetate ion, then, is BASIC.

We need to solve the BASE equilibrium of acetate ion...



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

$$\frac{(x)(x)}{0.100 - x} = K_b$$

$$\frac{x^2}{0.100 - x} = K_b$$

$$\frac{x^2}{0.100-x} = K_b, \text{C}_2\text{H}_3\text{O}_2^-$$

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

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

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

$$x = 7.67 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pOH} = 5.12$$

$$\text{pH} = 14.00 - 5.12 = \boxed{8.88}$$

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

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

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

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

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

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^+ : neutral - no H^+ (not B-L acid), not a likely base either!

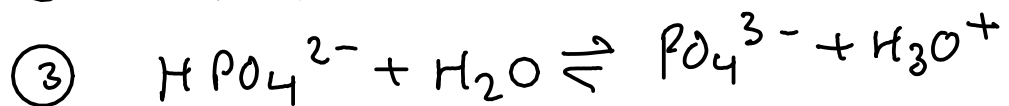
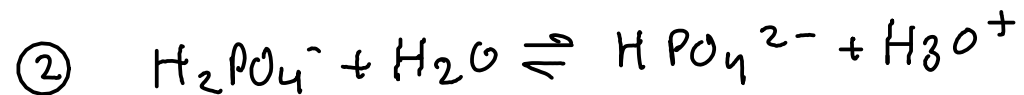
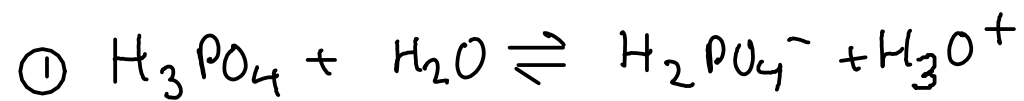


↑ strong acid, completely ionizes,
so Cl^- can't hang on to protons!

Since neither Na^+ nor Cl^- can function as acid or base, the water equilibrium itself will set the pH ... which will be $\text{pH} = 7.00$...

Find pH of 0.10 M H_3PO_4

... what's special about phosphoric acid?



Phosphoric acid has **THREE** acidic protons!

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

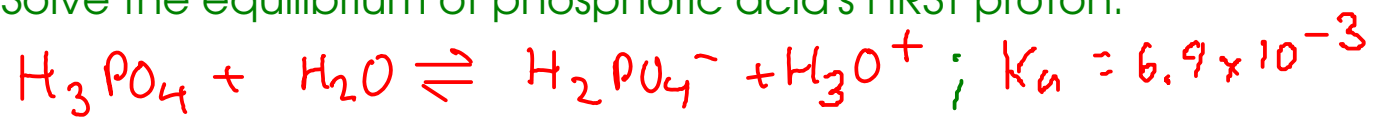
$$K_{a2} = 6.2 \times 10^{-8}$$

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

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 $0.10 - x \approx 0.10$

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

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

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

$$\boxed{\text{pH} = 1.58}$$

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!