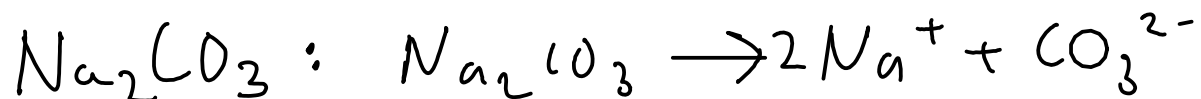


SALTS

- 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!

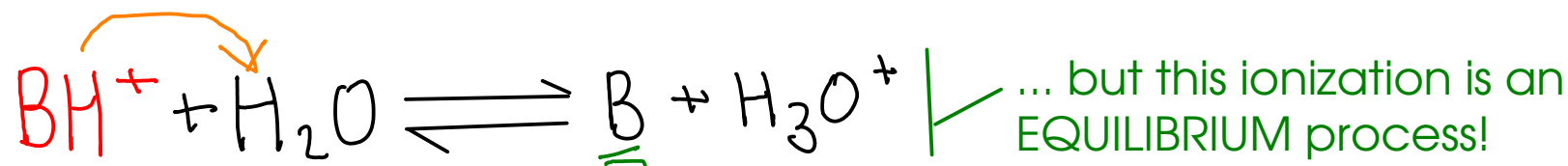
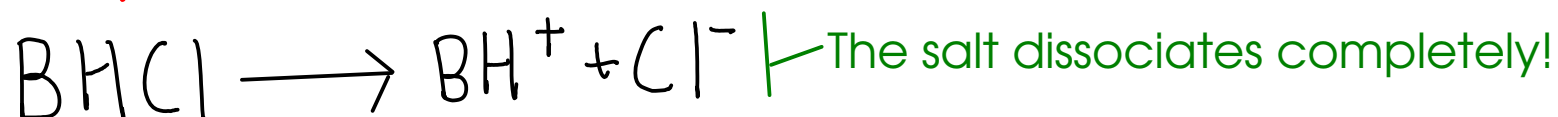
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

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? $\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$

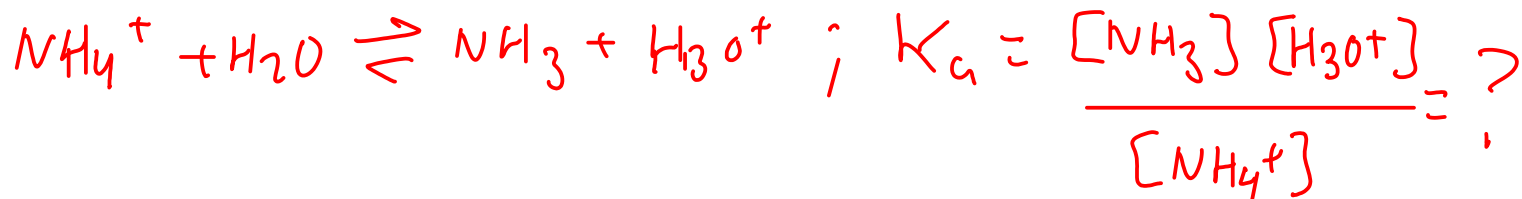


↳ This is ammonia, a WEAK BASE. It's water-stable, so this reaction is viable.



↳ 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

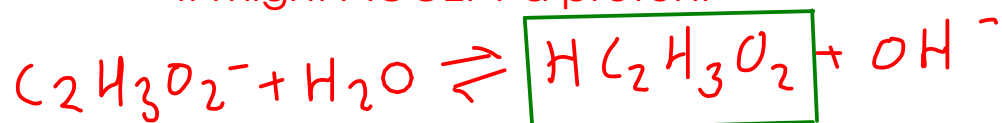
0.100 M $\text{NaC}_2\text{H}_3\text{O}_2$, Find pH



First, find out whether sodium acetate is acidic, basic, or neutral. Look at the IONS.

Na^+ : B-L acid? No; since it has no proton (H^+) to donate. Unlikely to be a base due to the positive charge which should repel protons.

$\text{C}_2\text{H}_3\text{O}_2^-$: Does have hydrogens, but also has a negative charge which means it might ACCEPT a proton.



Acetic acid (conjugate of acetate ion).

WEAK ACID. That means that acetate ion can hold on to a proton, and is BASIC.

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

We need a value for K_b . Let's consult the charts (OpenStax appendices).

The charts do not have a K_b for acetate ion, but they do have a K_a for

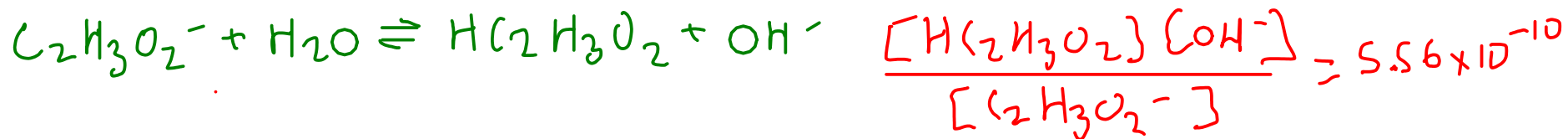
its conjugate, acetic acid: $K_a, \text{HC}_2\text{H}_3\text{O}_2 = 1.8 \times 10^{-5}$; $K_a \cdot K_b = K_w$

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

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

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





Solve this just like we'd solve any weak base pH calculation!

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

Let "x" equal the change in HYDROXIDE ion concentration

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

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

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

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

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

$$\text{pOH} = 5.13 \quad (\text{pOH} = -\log_{10} [\text{OH}^-])$$

$$\text{pH} = 14.00 - 5.13 \quad (\text{pH} + \text{pOH} = 14.00)$$

$$= \boxed{8.87}$$

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!

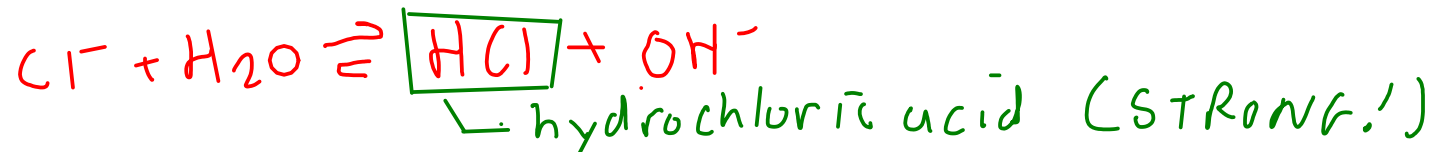
0.100 M NaCl, Find pH



Find out whether this salt is acid, basic, or neutral. Look at what the IONS do in water.

Na^+ : Has no protons it can donate, so not a B-L acid. Has a positive charge, which means it's unlikely to accept a proton. Neutral ... no effect on pH.

Cl^- : Has no protons to donate, but might accept one due to the negative charge.



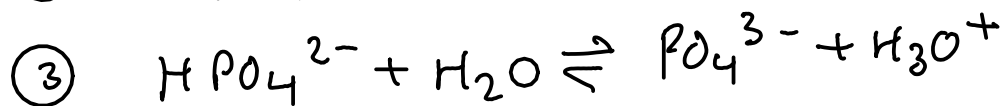
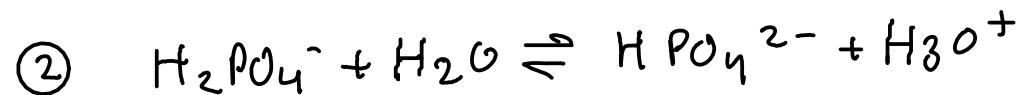
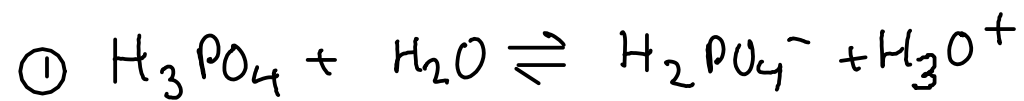
... so Cl^- can't accept a proton. Not a base.

Neither the sodium ion nor the chloride ion will affect the pH of the solution, so what DOES set the pH?

Since this is an aqueous solution, the water equilibrium itself sets the pH, and at room temperature, that 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!

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

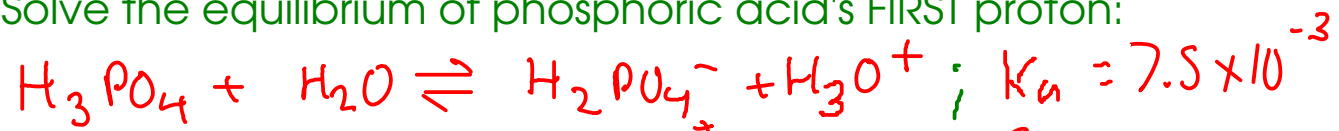
$$K_{a2} = 6.2 \times 10^{-8}$$

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

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 $0.10 - x \approx 0.10$

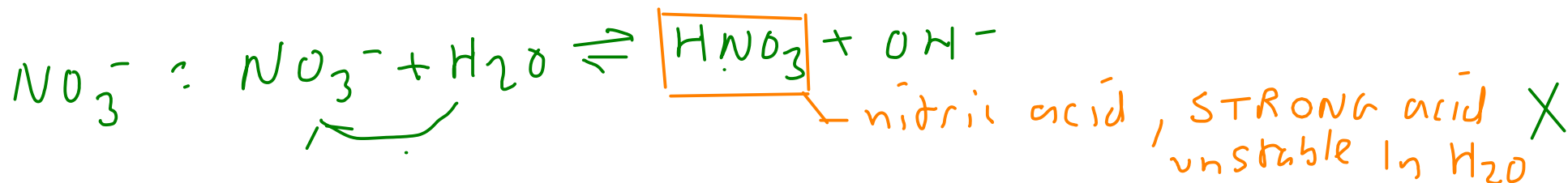
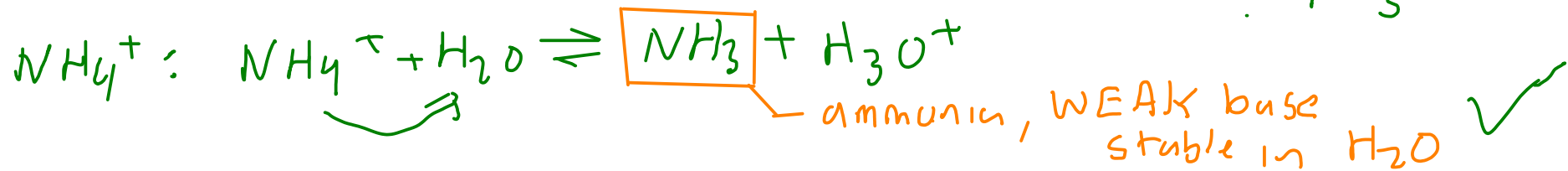
$$\frac{x^2}{0.10} = 7.5 \times 10^{-3}$$

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

$$\text{pH} = 1.56$$

143 Find the pH of an aqueous 0.150 M solution of ammonium nitrate (FW=80.052 g/mol).

This is a salt, so we'll need to see if its IONS are acidic, basic, or neutral: NH_4NO_3



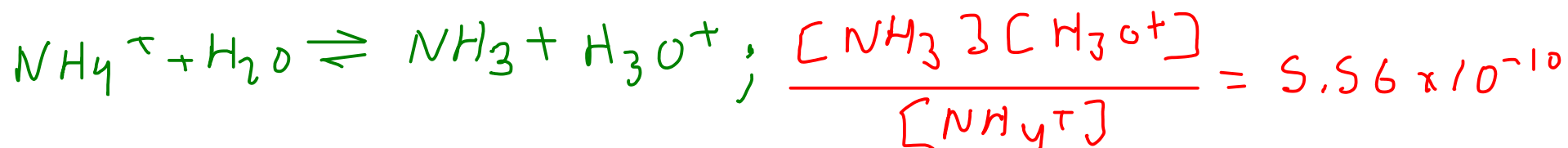
Since the ammonium ion can donate a proton and give a stable product while the nitrate ion can neither donate nor receive a proton, the equilibrium we need to solve is the ammonium ion equilibrium. We will need the equilibrium expression and a value for K_a .

$$K_{a, \text{NH}_4^+} = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

$K_{b, \text{NH}_3} = 1.8 \times 10^{-5}$
 $K_a K_b = K_w$
 $K_a (1.8 \times 10^{-5}) = 1.0 \times 10^{-14}$
 $K_a = 5.56 \times 10^{-10}$

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

<-- Solve this equilibrium expression



Set up an equilibrium chart and solve:

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	+x	x
NH_3	0	+x	x
NH_4^+	0.150	-x	0.150 - x

Let "x" equal the change in hydronium ion concentration

$$\frac{(x)(x)}{0.150 - x} = 5.56 \times 10^{-10}$$

$$\frac{x^2}{0.150 - x} = 5.56 \times 10^{-10}$$

$$\downarrow \begin{array}{l} x \ll 0.150 \\ 0.150 - x \approx 0.150 \end{array}$$

$$\frac{x^2}{0.150} = 5.56 \times 10^{-10}$$

$$x = 9.13 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = 5.04$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

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 .

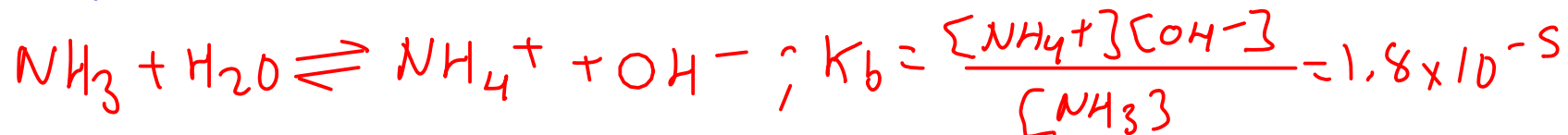
What would happen to the pH if we dissolved ammonium chloride into the solution?



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.



Species	[Initial]	Δ	[Equilibrium]
OH^-	0	+x	x
NH_4^+	0.10	+x	0.10 + x
NH_3	0.10	-x	0.10 - x

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

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

$$\begin{aligned} x &<< 0.10, 50 \\ 0.10 - x &\approx 0.10 \\ 0.10 + x &\approx 0.10 \end{aligned}$$

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

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

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

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

$$\text{pH} + 4.74 = 14.00$$

$$\text{pH} = \boxed{9.26}$$