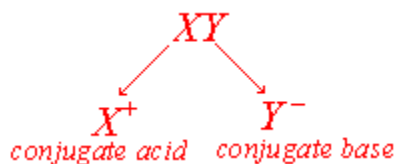


Conjugate Salts

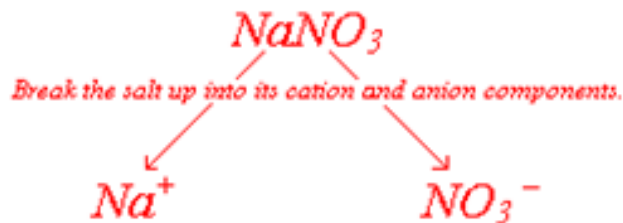
1. A salt is made up of a cation and an anion. The cation is typically a **conjugate acid** and the anion is typically a **conjugate base**. The strength of these **conjugates** depends on the strength of the substance from which it was derived.



2. This being the case, there are three types of salt solutions that can be formed
 - a. **Neutral**
 - b. **Acidic**
 - c. **Basic**
3. How do you determine which solution you have?
 - a. **Neutral – no effect on the pH.**

If the salt is made up of the conjugates of strong acids/bases then the salt is neutral. Remember that these conjugate have acid/base strength weaker than water – so in an aqueous solution they don't react.

For Example:



Na^+ , like all Group 1A and Group 2A elements are neutral in solution.

NO_3^- is the conjugate of a strong acid – meaning it is neutral in solution. In fact all of the conjugates of strong acids are neutral in solution. It is a relatively short list to remember. Along with NO_3^- , the list also includes Cl^- , ClO_4^- , Br^- , I^- , and ClO_3^- . The conjugates of H_2SO_4 are not included because things get more complicated with polyprotic acids.

b. Acidic – lowers pH

If the salt is made up of a cation that is the conjugate of a weak base and an anion that is the conjugate of a strong acid then the salt is acidic. Remember that the conjugate of a weak base is an acid stronger than water.

For Example:



NH_4^+ is the conjugate of the weak acid, NH_3 . This means that if this salt was put in solution, the following reaction would take place:

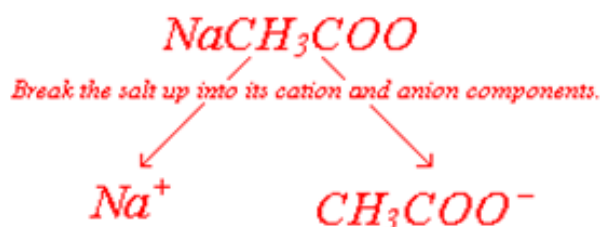


Cl^- is neutral in solution, as it is the conjugate of a strong acid. So it has no effect on the pH of the solution.

c. Basic – raises pH

If the salt is made up of an anion that is the conjugate of a weak acid and a cation that is the conjugate of a strong base the salt is basic. Remember that the conjugate of a weak acid is a base stronger than water.

For Example:



CH_3COO^- is the conjugate base of a weak acid. This means that if this salt was placed in solution, the following reaction would take place:



Na^+ is a group 1 cation, so it has no effect on the pH of the solution.

4. How do conjugate pairs equilibrium values relate?

Remember a conjugate pair differ by only one H^+ . The acid has one more H^+ than the base.

$$K_a \times K_b = K_w$$

This means that, for a conjugate pair, if you know the equilibrium constant for one, you can solve for the other.

5. What if both conjugates have strength?

You would then need to compare the K_a value of the cation with the K_b value of the anion.

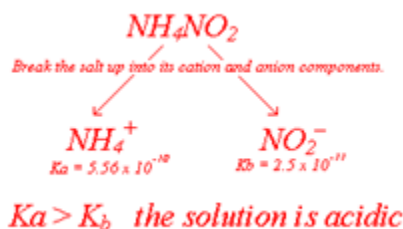
If:

$K_a = K_b$ the solution is neutral

$K_a > K_b$ the solution is acidic

$K_a < K_b$ the solution is basic.

For Example:



6. Calculate the pH of a 0.12M KNO_2 solution. ($K_a = 4.0 \times 10^{-4}$).

The first thing you want to do is determine the type of salt that you have



NO_2^- will react in aqueous solution.



$$K_b = \frac{[OH^-] [HNO_2]}{[NO_2^-]}$$

This problem is asking for pH which means, that I need to solve for the concentration of H^+ at equilibrium. If you notice there is no H^+ formed directly by this reaction. That does not mean that there isn't any H^+ in the solution though. This salt is placed in water. Water undergoes autoionization – meaning it produces H^+ . The amount of H^+ it produces is affected by the addition of the basic salt. Remember that

$$[H^+][OH^-] = K_w = 1.0 \times 10^{-14}$$

So if you determine the concentration of $[OH^-]$ you can solve for the concentration of $[H^+]$.

Alternatively, you could solve for the pOH and use the relationship

$$pH + pOH = 14$$

Both methods start with the ICE chart:

	$NO_2^-(aq)$	+	$H_2O(l)$	\rightleftharpoons	$HNO_2(aq)$	+	$OH^-(aq)$
<i>I</i>	0.12M		—		0		0
<i>C</i>	-x		—		+x		+x
<i>E</i>	0.12 - x		—		x		x

$$K_b = \frac{(x)(x)}{(0.12 - x)}$$

You have K_a not K_b . This means you need to solve for K_b using

$$K_a \times K_b = K_w$$

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{4.0 \times 10^{-4}} = 2.5 \times 10^{-11}$$

Now that you have K_b , you can plug it into the equation.

$$K_b = \frac{x^2}{(0.12 - x)} = 2.5 \times 10^{-11}$$

Because $K_b \ll 1$ you can make the assumption that $0.12 - x = 0.12$

$$\frac{x^2}{0.12} = 2.5 \times 10^{-11}$$

$$x = 1.73 \times 10^{-6}$$

$$\frac{1.73 \times 10^{-6}}{0.12} \times 100 = 1.44 \times 10^{-5} \%$$

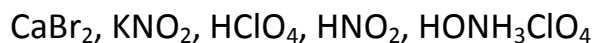
more than meets the 5% rule.

Now you know the equilibrium concentration of OH^- . At this point you can either solve for the $[H^+]$ or solve for the pOH. I'm going to solve for the pOH.

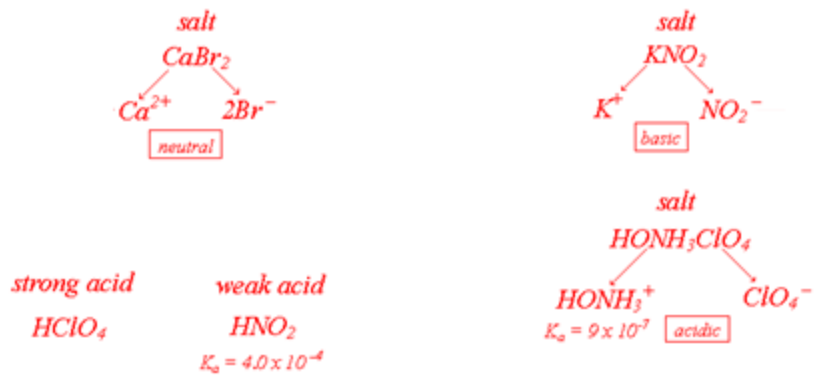
$$pOH = -\log [OH^-] = -\log (1.73 \times 10^{-6}) = 5.76$$

$$pH = 14 - pOH = 14 - 5.76 = \boxed{8.24}$$

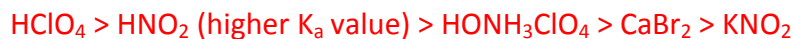
7. Arrange the following in order of acidic to most basic



Your first step is to identify what type of substance each compound is – meaning a salt, acid or base. If it is a salt break it up into its component pieces and determine their relative acid/base strengths of each.



Now we need to order these substances from most acidic to least acidic.
Strongest acidic → Strongest base



8. NaN_3 is added to water. Calculate the pH of a 0.010 M sol'n.
($K_a = 1.9 \times 10^{-5}$)

Before you can determine what the reaction equation is – you have to figure out if you have an acidic or basic salt.



$$K_b = \frac{[\text{HN}_3][^-\text{OH}]}{[\text{N}_3^-]}$$

Make the ICE chart...

	$\text{N}_3^-(\text{aq})$	+	$\text{H}_2\text{O}(\text{l})$	\rightleftharpoons	$\text{HN}_3(\text{aq})$	+	$^-\text{OH}(\text{aq})$
<i>I</i>	0.010M		—		0		0
<i>C</i>	-x		—		+x		+x
<i>E</i>	0.010 - x		—		x		x

Once again we plug into the equilibrium expression to determine x . Remember that you will be plugging into the expression for K_b – because this is a base reaction. In the problem they gave you the K_a for the conjugate acid HN_3 – so you will first need to solve for K_b .

$$K_b = \frac{K_w}{K_a} = \frac{1 \times 10^{-14}}{1.9 \times 10^{-5}} = 5.26 \times 10^{-10}$$

Now you can plug in and solve for x .

$$5.26 \times 10^{-10} = \frac{x^2}{0.010 - x}$$

As K_b is quite a small number, you are safe to assume that the change to the starting concentration is negligible.

$$5.26 \times 10^{-10} = \frac{x^2}{0.010}$$
$$x = 2.30 \times 10^{-6}$$

Using the 5% rule make sure your assumption was valid

$$\frac{2.30 \times 10^{-6}}{0.010} \times 100 = 0.02\%$$

This answer is well within the parameters of being valid.

Once again the x value that you have calculated for relates to the $[\text{OH}^-]$. So you will initially solve for pOH and then use that value to determine pH .

$$\text{pOH} = -\log [\text{OH}^-] = -\log (2.30 \times 10^{-6}) = 5.63$$
$$\text{pH} = 14 - \text{pOH} = 14 - 5.63 = \boxed{8.37}$$