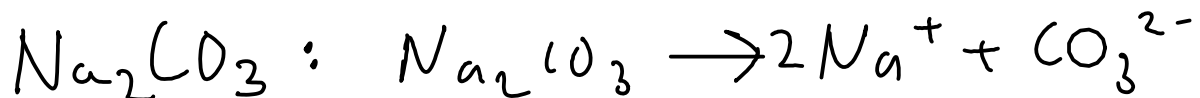


## SALTS

- 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?

$\text{Na}^+$ : neutral. Not a proton donor or a proton acceptor

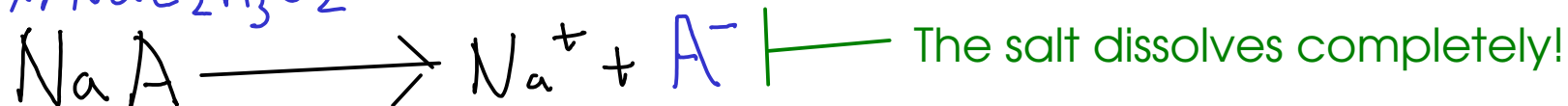
$\text{CO}_3^{2-}$ : BASIC, since it can accept protons to form the weak acid CARBONIC ACID in solution.



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  $\text{A}^-$  and HA are a conjugate pair, the ionization constants are related!

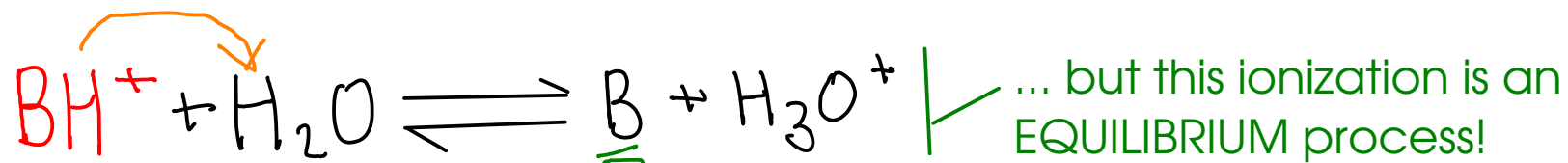
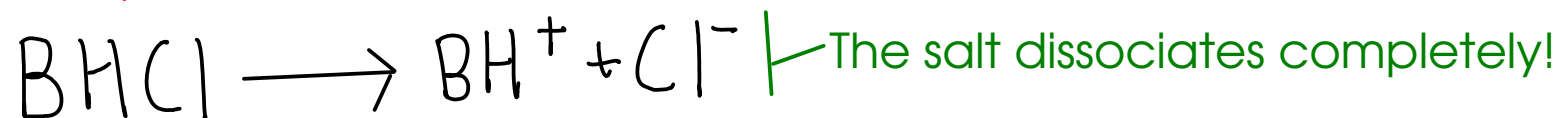
$$K_w = (K_{a,\text{HA}})(K_{b,\text{A}^-})$$

$1.0 \times 10^{-14}$

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

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:  $\text{NH}_4\text{Cl}$ 

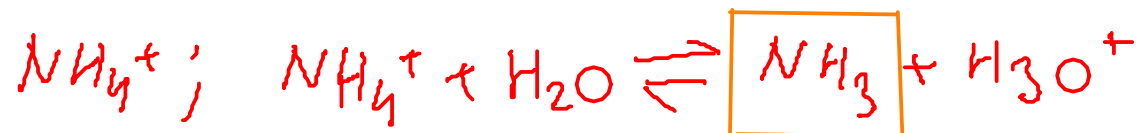
$$K_a = \frac{[\text{B}][\text{H}_3\text{O}^+]}{[\text{BH}^+]} \quad \left| \text{Acid ionization constant for BH}^+ \right.$$

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

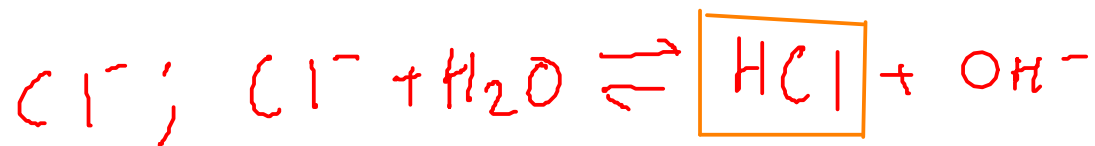
$1.0 \times 10^{-14}$

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  $\text{NH}_4\text{Cl}$  ... Find the pH of the solution

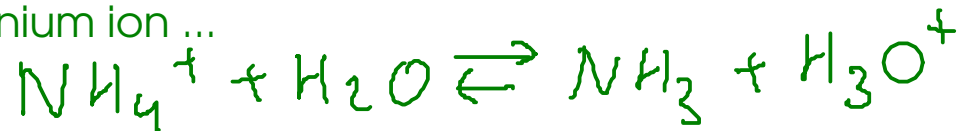


Ammonia is a WEAK base. This means that it exists in water as a molecule. Since ammonia is stable in water, AMMONIUM ION should function as an ACID.

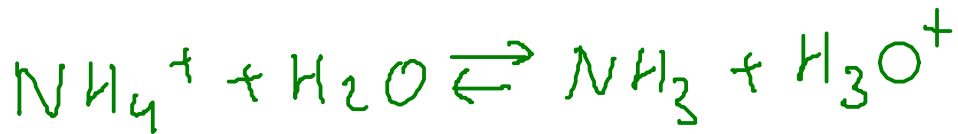


HCl is a STRONG ACID. It completely ionizes in water (in other words, it's unstable). Therefore, CHLORIDE ION is NEUTRAL.

So, to find the pH of the solution, we will have to solve the equilibrium of the ammonium ion ...



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



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

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

Value for  $K_a$ ? Chart on A-13 to A-14 does not give  $K_a$  for ammonium, but DOES give  $K_b$  for ammonia!

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

$$K_a, \text{NH}_4^+ \times K_b, \text{NH}_3 = 1.0 \times 10^{-14}$$

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

Species	[Initial]	$\Delta$	[Equilibrium]
$\text{NH}_3$	0	+X	X
$\text{H}_3\text{O}^+$	0	+X	X
$\text{NH}_4^+$	0.100	-X	0.100 - X

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

$$0.100 - x$$

Assume x is small

$$0.100 - x = 0.100$$

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

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

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

Compare!

pH = 1.00 for 0.100 M strong acid

pH = 2.17 for 0.100 M nitrous acid

pH = 7.00 for distilled water

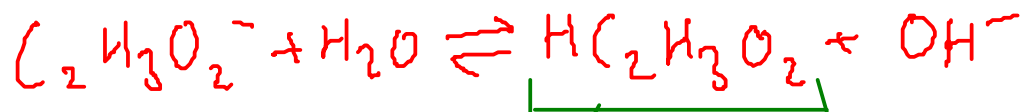
0.100 M  $\text{NaC}_2\text{H}_3\text{O}_2$ , Find pH



Check the ions formed by the salt to see if they have acidic or basic properties.

$\text{Na}^+$ : Not B-L acid, since there is no  $\text{H}^+$  to donate. Not likely to be B-L base due to the positive charge ... which would repel  $\text{H}^+$

$\text{C}_2\text{H}_3\text{O}_2^-$ : This species does have hydrogen atoms, but it's more likely to be a B-L base due to the negative charge. (Would attract  $\text{H}^+$ )

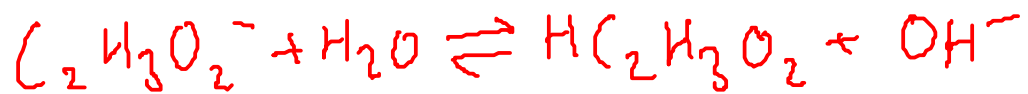


ACETIC ACID: This is a WEAK ACID, meaning that it's stable in water. Since the product of the reaction is stable in water, we expect acetate ion to function as a base.

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

Value for  $K_b$ ? As before, we don't find a  $K_b$  for acetate in our chart on A-14, so we look for a  $K_a$  for the conjugate - acetic acid.

$$K_{a, \text{HC}_2\text{H}_3\text{O}_2} = 1.7 \times 10^{-5}; \text{ so } K_b = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} = 5.89 \times 10^{-10}$$



$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5.89 \times 10^{-10}$$

Now, we set up and solve this B-L BASE equilibrium to find the pH

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

$$\frac{(x)(x)}{0.100 - x} = 5.89 \times 10^{-10}$$

$$\frac{x^2}{0.100 - x} = 5.89 \times 10^{-10}$$

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

$$\frac{x^2}{0.100} = 5.89 \times 10^{-10}$$

$$x = 7.669649688 \times 10^{-6}$$

$$[\text{OH}^-] = x, \text{ so}$$

$$\text{pOH} = 5.12$$

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

$$\boxed{\text{pH} = 8.88}$$

(Need to convert OH- to hydronium. We'll use pOH to do this ...)

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!



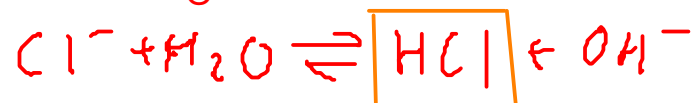
0.100 M NaCl, Find pH



Find out whether this salt is acidic or basic. Look at the ions the salt contains.

$\text{Na}^+$ : Not likely to be a B-L acid, as it has no hydrogen atoms to lose. Not likely to be a B-L base due to the positive charge repelling  $\text{H}^+$ . NEUTRAL.

$\text{Cl}^-$ : Not likely to be a B-L acid, as it has no hydrogen atoms to lose. Might be a base, though - let's check:



STRONG acid!

Since the conjugate acid (HCl) is STRONG, we don't expect chloride ion to be able to hold on to a proton (since strong acids completely ionize!). Chloride ion will be NEUTRAL.

The pH of this solution, then, will be controlled by the water equilibrium only. That means the pH of 0.100 M NaCl will be 7.00 (neutral).

Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate (FW=80.052 g/mol) solid into enough water to make 250. mL of solution.

---

First, find out whether the salt is acidic or basic. Look at the ions ...

$\text{NH}_4^+$ : Looks like it could be an acid. (H's and positive charge)



Ammonia ... WEAK base

We expect ammonium ion to be acidic, since its conjugate is a weak base (and therefore stable in water!)

$\text{NO}_3^-$ : Might be a base ... (no protons, but negative charge)



Nitric acid ... STRONG acid.

Nitrate ion will be neutral, as nitric acid ionizes completely.

We'll work with the ammonium ion equilibrium. This is an ACIDIC salt!



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

We don't find  $K_a$  for ammonium ion on pages A-13 or A-14. But we CAN find  $K_b$  for its conjugate, ammonia.

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

So,  $K_a \times K_b = 1.00 \times 10^{-14}$  ... for conjugate pairs

$$K_a (1.8 \times 10^{-5}) = 1.00 \times 10^{-14}$$

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



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

Species	Initial	$\Delta$	Equilibrium
$\text{NH}_3$	0	+x	x
$\text{H}_3\text{O}^+$	0	+x	x
$\text{NH}_4^+$	0.1499	-x	0.1499 - x

$$\frac{(x)(x)}{(0.1499 - x)} = 5.56 \times 10^{-10}$$

assume  $x \ll 0.1499$   
 $0.1499 - x \approx 0.1499$

$$\frac{x^2}{0.1499} = 5.56 \times 10^{-10}$$

$$x = 9.13 \times 10^{-6} = [\text{H}_3\text{O}^+] \quad \dots \text{from our equilibrium chart}$$

$$\text{So, } \boxed{\text{pH} = 5.04}$$

Find initial concentration of ammonium:

$$3.00 \text{ g NH}_4\text{NO}_3 \times \frac{\text{mol NH}_4\text{NO}_3}{80.052 \text{ g NH}_4\text{NO}_3} =$$

$$= 0.0374756408 \text{ mol NH}_4\text{NO}_3$$

$$M = \frac{0.0374756408 \text{ mol}}{0.250 \text{ L}} =$$

$$= 0.1499025633 \text{ M NH}_4\text{NO}_3$$

$$\approx 0.1499 \text{ M NH}_4^+$$

## 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!