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

Na_2CO_3: $Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$

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.

$$H_2CO_3 + 2H_2O \rightleftharpoons 2H_3O^+ + CO_3^{-2}$$

ACID BASE

SALT OF A WEAK ACID

ex: $NaC_2H_3O_2$

$$NaA \longrightarrow Na^+ + A^-$$

The salt dissolves completely!

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

$$A^- + H_2O \rightleftharpoons HA + OH^-$$

... but the ionization of the salt's anion is an EQUILIBRIUM!

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

$$K_b = \frac{[HA][OH^-]}{[A^-]}$$

This is the base ionization constant for A^-

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

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

1.0×10^{-14}

$$14 = pK_a + pK_b$$

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

SALT OF A WEAK BASE

ex: NH_4Cl

$$BHCl \longrightarrow BH^+ + Cl^-$$

The salt dissociates completely!

$$BH^+ + H_2O \rightleftharpoons \underline{B} + H_3O^+$$

... but this ionization is an EQUILIBRIUM process!

$$K_a = \frac{[B][H_3O^+]}{[BH^+]}$$

Acid ionization constant for BH^+

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

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

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$

NH_4^+; $NH_4^+ + H_2O \rightleftharpoons \boxed{NH_3} + H_3O^+$

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

Cl^-; $Cl^- + H_2O \rightleftharpoons \boxed{HCl} + OH^-$

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 ...

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

$$NH_4^+ + H_2O \rightleftarrows NH_3 + H_3O^+$$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

Value for Ka? Chart on A-13 to A-14 does not give Ka for ammonium, but DOES give Kb for ammonia!

$$K_{b,NH_3} = 1.8 \times 10^{-5}$$

$$K_{a,NH_4^+} \times K_{b,NH_3} = 1.0 \times 10^{-14}$$

$$K_{a,NH_4^+} = 5.56 \times 10^{-10}$$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56 \times 10^{-10}$$

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

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

Assume x is small

0.100 -x = 0.100

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

$$x = 7.45 \times 10^{-6} = [H_3O^+]$$

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

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 $NaC_2H_3O_2$, Find pH

$$NaC_2H_3O_2 \rightarrow Na^+ + C_2H_3O_2^-$$

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

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

$C_2H_3O_2^-$: This species does have hydrogen atoms, but it's more likely to be a B-L base due to the negative charge. (Would attract H+)

$$C_2H_3O_2^- + H_2O \rightleftharpoons \underbrace{HC_2H_3O_2}_{} + OH^-$$

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.

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

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

$$K_{a,HC_2H_3O_2} = 1.7 \times 10^{-5}; \text{ so } K_b = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} = 5.89 \times 10^{-10}$$

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} = 5.89 \times 10^{-10}$$

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

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

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

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

$x << 0.100$, so $0.100 - x \approx 0.100$

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

$x = 7.67464888 \times 10^{-6}$

$[OH^-] = x$, so

$pOH = 5.12$

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

$pH = 14.00 - 5.12$

$\boxed{pH = 8.88}$

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

$$NaCl \rightarrow Na^{+} + Cl^{-}$$

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

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 H+. NEUTRAL.

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:

$$Cl^{-} + H_2O \rightleftharpoons \boxed{HCl} + OH^{-}$$

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).

163 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 ...

NH_4^+ : Looks like it could be an acid. (H's and positive charge)

$$NH_4^+ + H_2O \rightleftharpoons \boxed{NH_3} + H_3O^+$$

Ammonia ... WEAK base

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

NO_3^- : Might be a base ... (no protons, but negative charge)

$$NO_3^- + H_2O \rightleftharpoons \boxed{HNO_3} + OH^-$$

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!

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = ?$$

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

$K_{b,NH_3} = 1.8\times10^{-5}$

So, $K_a \times K_b = 1.00\times10^{-14}$ ... for conjugate pairs

$$K_a(1.8\times10^{-5}) = 1.00\times10^{-14}$$

$$K_a = 5.56\times10^{-10}$$

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

$$K_a = \frac{[NH_3][H_3O^+]}{[NH_4^+]} = 5.56\times10^{-10}$$

Species	Initial	Δ	Equilibrium
NH_3	0	$+x$	x
H_3O^+	0	$+x$	x
NH_4^+	0.1499	$-x$	$0.1499-x$

$$\frac{(x)(x)}{(0.1499-x)} = 5.56\times10^{-10}$$

assume $x \ll 0.1499$

$0.1499 - x \approx 0.1499$

$$\frac{x^2}{0.1499} = 5.56\times10^{-10}$$

$$x = 9.13\times10^{-6} = [H_3O^+]$$

... from our equilibrium chart

So, $\boxed{pH = 5.04}$

Find initial concentration of ammonium:

$$3.00\,g\,NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.052\,g\,NH_4NO_3} =$$

$$= 0.0374756408\ mol\ NH_4NO_3$$

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

$$= 0.149902563\ M\ NH_4NO_3$$

$$\approx 0.1499\ M\ NH_4^+$$

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 Chateleir's Principle applied to equilibria involving ions

ex: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$; $K_b = 1.8 \times 10^{-5}$

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?

$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$

The ammonium chloride provides the ammonium ion. According to Le Chateleir's principle, this would shift the ammonia equilibrium to the LEFT!

What would happen to the pH? Let's find out!