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: $NaC_2H_3O_2$

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

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 contant for A^-

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

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

$K_w = 1.0 \times 10^{-14}$

$$14 = pK_a + pK_b$$

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!

$$BH^+ + H_2O \rightleftharpoons B + H_3O^+$$ ... but this ionization is an EQUILIBRIUM process!

$$K_a = \frac{[B][H_3O^+]}{[BH^+]}$$ Acid ionization constant for BH^+

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

K_w: 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

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

Acidic, basic, or neutral salt?

This is the WEAK BASE ammonia. Stable in water.

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

Cl^-: $Cl^- + H_2O \rightleftharpoons HCl + OH^-$ X

This is a STRONG ACID, which does not exist as a stable molecule in water.

The conjugate of a strong acid or base is NEUTRAL - does not affect pH!

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ This equilibrium affects the pH, so it is the equilibrium we'll need to solve to find pH!

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

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

Where do we get this Ka?

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

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

	initial	Δ	equilibrium
$[NH_3]$	0	+x	x
$[H_3O^+]$	0	+x	x
$[NH_4^+]$	0.100	−x	0.100 − x

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

$x << 0.100$

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

$$x = 7.46 \times 10^{-6}$$

We defined "x" as the concentration of hydronium ion made by the equilibrium!

$$x = 7.46 \times 10^{-6} = [H_3O^+]$$

$$pH = -\log_{10}(7.46 \times 10^{-6}) = \boxed{5.13 = pH}$$

Compare:
pH = 1.00 for 0.100 M strong acid
pH = 2.16 for 0.100 M nitrous acid
pH = 7.00 for pure distilled water

0.100 M $NaC_2H_3O_2$, Find pH

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

Check these ions to see if they are acidic, basic, or neutral

Na^+: Cannot be B-L acid (no protons), also not likely to be B-L base; positively charged.

$C_2H_3O_2^-$: Has protons, but might not be as likely to donate one as to receive one due to the negative charge.

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

Acetic acid is a WEAK ACID and stable in water, so the acetate ion can function as a BASE.

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

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

We don't know the Kb for acetate ion, but we DO know that it's related to Ka for acetic acid!

$K_{a,HC_2H_3O_2} = 1.7 \times 10^{-5}$ $K_a \times K_b = 1.00 \times 10^{-14}$

True for any conjugate pair

So, $K_{b,C_2H_3O_2^-} = 5.88 \times 10^{-10}$

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

$$K_b = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]} = 5.88 \times 10^{-10}$$

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

We defined 'x' as the concentration of HYDROXIDE ION produced by the equilibrium. This is related to, but not quite the same thing, as HYDRONIUM ION concentration.

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

$$\downarrow \; x \ll 0.100$$

$$\frac{x^2}{0.100} = 5.88 \times 10^{-10} \; ; \; x = 7.67 \times 10^{-6} = [OH^-]$$

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

$pOH = -\log_{10}(7.67 \times 10^{-6})$

$pOH = 5.12$

Now, convert pOH to pH using the relationship from the water equilibrium

$pH + pOH = 14.00$

So, $pH = 14.00 - 5.12$

$pH = 8.88$

Compare:

pH = 7.00 for pure distilled water

pH = 13.00 for 0.100 M sodium hydroxide (strong base)

0.100 M NaCl, Find pH

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

Check these ions to see if they are acidic, basic, or neutral

Na^+: Cannot be B-L acid (no protons), also not likely to be B-L base; positively charged.

Cl^-: Cannot be B-L acid (no protons). What about a base?

$$Cl^- + H_2O \rightleftharpoons HCl + OH^-$$

Hydrochloric acid is a STRONG ACID, meaning that it doesn't stay together in water. The reaction we wrote above is not likely to occur. Chloride ion, therefore, is a neutral ion.

Since neither ion affects the water equilibrium, the pH of this solution should be 7.00 (same as distilled water)

POLYPROTIC ACIDS

Find pH of 0.10 M H_3PO_4

... what's special about phosphoric acid?

① $H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$

② $H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$

③ $HPO_4^{2-} + H_2O \rightleftharpoons PO_4^{3-} + H_3O^+$

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 dissocation 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.

Solving the equilibrium of phosphoric acid's first proton:

$$H_3PO_4 + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$$

$$K_a = 6.9 \times 10^{-3} = \frac{[H_2PO_4^-][H_3O^+]}{[H_3PO_4]}$$

Species	[Initial]	Δ	[Equilibrium]
$H_2PO_4^-$	0	+ X	X
H_3O^+	0	+ X	X
H_3PO_4	0.10	− X	0.10 − X

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

This time, we should solve the quadratic. The x<<0.10 might not be safe in this problem. (Ka isn't 1000x smaller than 0.10)

$$x^2 = (6.9 \times 10^{-4}) - (6.9 \times 10^{-3})x$$

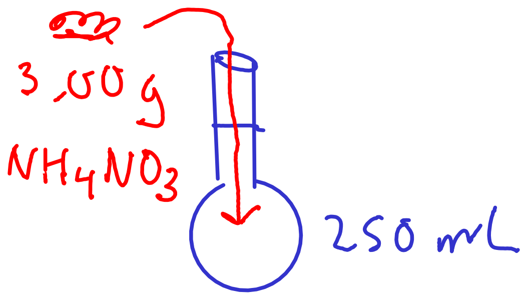
$$x^2 + (6.9 \times 10^{-3})x - (6.9 \times 10^{-4}) = 0$$

$a = 1$, $b = 6.9 \times 10^{-3}$, $c = -6.9 \times 10^{-4}$

$X = 0.023034426$ OR ~~-0.029934426~~

$= [H_3O^+]$, so **pH = 1.64**

Find the pH of a solution prepared by dissolving 3.00 g of ammonium nitrate solid into enough water to make 250. mL of solution.



NH_4NO_3:

N : 2 × 14.01

H : 4 × 1.008

O : 3 × 16.00

80.052 g NH_4NO_3 = 1 mol NH_4NO_3

Find out what kind of salt we have: acidic, basic, or neutral:

$$NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$$

NO_3^-: $NO_3^- + H_2O \rightleftharpoons HNO_3 + OH^-$ Nitric acid is a strong acid, so nitrate ion must be neutral

NH_4^+: $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$ Ammonia is a weak base, so ammonium ion is a weak acid

So, we will solve this equilibrium:

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

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

K_a:

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

$K_a \times K_b = 1.00 \times 10^{-14}$

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

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

To set this problem up further, we need to determine the initial ammonium ion concentration.

$$3.00\,g\,NH_4NO_3 \times \frac{mol\,NH_4NO_3}{80.052\,g\,NH_4NO_3} \times \frac{}{0.250\,L} = 0.1499\,\underline{M}\,NH_4NO_3$$

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

Solve the equilibrium:

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

$\downarrow$ $x << 0.1499$

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

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

$$\boxed{pH = 5.04}$$