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

\[ \text{Na}_2\text{CO}_3 : \quad \text{Na}_2\text{CO}_3 \rightarrow 2\text{Na}^+ + \text{CO}_3^{2-} \]

Do any of these ions have acidic or basic properties?

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

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

\[ \text{H}_2\text{CO}_3 + 2\text{H}_2\text{O} \rightleftharpoons 2\text{H}_3\text{O}^+ + \text{CO}_3^{2-} \]

ACID       BASE
SALT OF A WEAK ACID

\[ \text{ex: } \text{NaC}_2\text{H}_3\text{O}_2 \]

\[ \text{Na}^+ + A^- \quad \text{Na}^+ > N_a + A^- \quad \text{The salt dissolves completely!} \]

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

\[ A^- + H_2O \rightleftharpoons HA + OH^- \quad \text{... but the ionization of the salt's anion is an EQUILIBRIUM!} \]

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

\[ K_b = \frac{[HA][OH^-]}{[A^-]} \quad \text{This is the base ionization constant for } A^- \]

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

\[ K_w = (K_{a,H A})(K_{b,A^-}) \]

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!
The salt dissociates completely!

... but this ionization is an EQUILIBRIUM process!

Acid ionization constant for BH⁺:

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

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

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

Acidic, basic, or neutral salt?

- $\text{NH}_4^+: \text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons [\text{NH}_3] + \text{H}_3\text{O}^+$ \checkmark
- $\text{Cl}^-: \text{Cl}^- + \text{H}_2\text{O} \rightleftharpoons [\text{HCl}] + \text{OH}^-$ \xmark

This is the WEAK BASE ammonia. Stable in water.

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!

$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+$ This equilibrium affects the pH, so it is the equilibrium we'll need to solve to find pH!
Where do we get this $K_a$?

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We defined $\chi$ as the concentration of hydronium ion made by the equilibrium.

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

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

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

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

\( \text{C}_2\text{H}_3\text{O}_2^- \): Has protons, but might not be as likely to donate one as to receive one due to the negative charge.

\[ \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^- \]

\( \text{Acetic acid is a WEAK ACID and stable in water, so the acetate ion can function as a BASE.} \)

\[ \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^- \]

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

We don't know the \( K_b \) for acetate ion, but we DO know that it's related to \( K_a \) for acetic acid!

\[ K_a, \text{HC}_2\text{H}_3\text{O}_2 = 1.7 \times 10^{-5} \]

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

\( \text{So,} \ K_b, \text{C}_2\text{H}_3\text{O}_2^- = 5.88 \times 10^{-10} \)
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.

\[
\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2^- + \text{OH}^- \\
K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2^-][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = 5.88 \times 10^{-10}
\]

<table>
<thead>
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<th>[Initial]</th>
<th>(\Delta)</th>
<th>[Equilibrium]</th>
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<tbody>
<tr>
<td>(\text{HC}_2\text{H}_3\text{O}_2^-)</td>
<td>0</td>
<td>+ x</td>
<td>x</td>
</tr>
<tr>
<td>(\text{OH}^-)</td>
<td>0</td>
<td>+ x</td>
<td>x</td>
</tr>
<tr>
<td>(\text{C}_2\text{H}_3\text{O}_2^-)</td>
<td>0.100</td>
<td>- x</td>
<td>0.100 - x</td>
</tr>
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</table>

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.

\[
\frac{x^2}{0.100 - x} = 5.88 \times 10^{-10} \\
\downarrow \quad x << 0.100 \quad \Rightarrow \quad x = 7.67 \times 10^{-6} = [\text{OH}^-]
\]
\[ \chi = 7.67 \times 10^{-6} = [\text{OH}^-] \]

\[ \text{pOH} = -\log_{10}(7.67 \times 10^{-6}) \]

\[ \text{pOH} = 5.12 \]

Now, convert \( \text{pOH} \) to \( \text{pH} \) using the relationship from the water equilibrium:

\[ \text{pH} + \text{pOH} = 14.00 \]

So,

\[ \text{pH} = 14.00 - 5.12 \]

\[ \text{pH} = 8.88 \]

Compare:

- \( \text{pH} = 7.00 \) for pure distilled water
- \( \text{pH} = 13.00 \) for 0.100 M sodium hydroxide (strong base)
0.100 M NaCl, Find pH

\[ \text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^- \]

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

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

\[ \text{Cl}^- : \text{Cannot be B-L acid (no protons). What about a base?} \]

\[ \text{Cl}^- + \text{H}_2\text{O} \rightleftharpoons \text{HCl} + \text{OH}^- \]

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 H₃PO₄

... what's special about phosphoric acid?

1. \( H₃PO₄ + H₂O \rightleftharpoons H₂PO₄⁻ + H₃O⁺ \)
2. \( H₂PO₄⁻ + H₂O \rightleftharpoons HPO₄^{2⁻} + H₃O⁺ \)
3. \( HPO₄^{2⁻} + H₂O \rightleftharpoons PO₄^{3⁻} + H₃O⁺ \)

Phosphoric acid has THREE acidic protons!

| \( Kₐ₁ \) | \( 6.9 \times 10^{-3} \) |
| \( Kₐ₂ \) | \( 6.2 \times 10^{-8} \) |
| \( Kₐ₃ \) | \( 4.8 \times 10^{-13} \) |

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.
Solving the equilibrium of phosphoric acid’s first proton:

\[ \text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{PO}_4^- + \text{H}_3\text{O}^+ \]

\[ K_a = 6.9 \times 10^{-3} = \frac{[\text{H}_2\text{PO}_4^-][\text{H}_3\text{O}^+]}{[\text{H}_3\text{PO}_4]} \]

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<td>( \text{H}_2\text{PO}_4^- )</td>
<td>0</td>
<td>+ ( x )</td>
<td>( x )</td>
</tr>
<tr>
<td>( \text{H}_3\text{O}^+ )</td>
<td>0</td>
<td>+ ( x )</td>
<td>( x )</td>
</tr>
<tr>
<td>( \text{H}_3\text{PO}_4 )</td>
<td>0.10</td>
<td>- ( x )</td>
<td>0.10 - ( x )</td>
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\[ \frac{x^2}{0.10-x} = 6.9 \times 10^{-3} \]

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

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

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

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

\[ x = 0.0230434426 \text{ or } -0.0230434426 \]

\( [\text{H}_3\text{O}^+] \), so \( \text{pH} = 1.64 \)
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.

\[ \text{NH}_4 \text{NO}_3 : \quad \text{N} : 2 \times 14.01 \\
\quad \text{H} : 4 \times 1.008 \\
\quad \text{O} : 3 \times 16.00 \]

\[ 80.052 \text{ g NH}_4\text{NO}_3 = 1 \text{ mol NH}_4\text{NO}_3 \]

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

- Nitric acid is a strong acid, so nitrate ion must be neutral
- Ammonia is a weak base, so ammonium ion is a weak acid

So, we will solve this equilibrium:

\[ \text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}_3\text{O}^+ \]

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

\[ K_a = 1.8 \times 10^{-5} \]

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

\[ K_{a1} \text{NH}_4^+ = 5.56 \times 10^{-10} \]
To set this problem up further, we need to determine the initial ammonium ion concentration.

Solve the equilibrium:

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

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

\[
3.00 \text{ g NH}_4\text{NO}_3 \times \frac{\text{mol NH}_4\text{NO}_3}{60.052 \text{ g NH}_4\text{NO}_3} \times \frac{1}{0.250 \text{ L}} = 0.1499 \text{ M NH}_4\text{NO}_3
\]

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<td>0</td>
<td>+ (x)</td>
<td>(x)</td>
</tr>
<tr>
<td>(H_3O^+)</td>
<td>0</td>
<td>+ (x)</td>
<td>(x)</td>
</tr>
<tr>
<td>(NH_4^+)</td>
<td>0.1499</td>
<td>- (x)</td>
<td>0.1499 - (x)</td>
</tr>
</tbody>
</table>

Solve the equilibrium:

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

\[
x \ll 0.1499
\]

\[
x^2 = 5.56 \times 10^{-10}
\]

\[
x = 9.13 \times 10^{-6} = [H_3O^+]
\]

\[
\text{pH} = 5.04
\]