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!

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

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

$$14 = pK_a + pK_b$$

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!

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

... 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

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

Acidic, basic, or neutral salt?

This is the WEAK BASE ammonia. Stable in water.

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

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

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!

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

$$K_{a,NH_4^+} = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

We can get Ka for the acid from Kb for the base!

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

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

So, $K_a = 5.56 \times 10^{-10}$ (for NH_4^+)

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
H_3O^+	0	+ x	x
NH_3	0	+ x	x
NH_4^+	0.100	− x	0.100 − x

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

$x \ll 0.100$

$0.100 - x \approx 0.100$

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

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

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

$pH = -\log_{10}(7.45 \times 10^{-6}) = 5.13$

Compare:
pH = 1.00 for 0.100 M strong acid
pH = 2.16 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 to see if they are acidic, basic, or neutral!

Na^+ : Cannot be a B-L acid (no H), also not likely to be B-L base, since it's positively charged.

$C_2H_3O_2^-$: Has protons, but also has a negative charge - so it may be more likely to receive protons!

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

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

$$C_2H_3O^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$

$$K_{b,C_2H_3O_2^-} = \frac{[HC_2H_3O_2][OH^-]}{[C_2H_3O_2^-]}$$

Kb for acetate ion isn't in the chart in the appendix, but the Ka for acetic acid (the conjugate acid of acetate ion) is available!

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

$K_b = 5.88 \times 10^{-10}$

$$C_2H_3O^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$

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

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
OH^-	0	+ x	x
$HC_2H_3O_2$	0	+ x	x
$C_2H_3O_2^-$	0.100	− x	0.100 − x

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

↓ x << 0.100, so 0.100 − x ≈ 0.100

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

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

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

$pOH = -\log_{10}(7.67 \times 10^{-6}) = 5.12$

Calculate poH, then convert to pH using 'pH + pOH = 14'

$pH = 14.00 - 5.12$

$pH = 8.88$

Compare:
pH = 7.00 for pure distilled water
pH = 13.00 for 0.100 M strong base
pH = 11.13 for 0.100 M ammonia

0.100 M NaCl, Find pH

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

Check these ions to see if they're acid, basic, or neutral:

Na^+: Cannot be a B-L acid (no H), also not likely to be B-L base, since it's positively charged.

Cl^-: Cannot be a B-L acid (no H), but can it act as a base?

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

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

So, chloride ion is ALSO a neutral ion!

Since neither sodium ion nor chloride ion affect the water equilibrium, the pH of this solution will be the same as the pH of pure water: 7.00