

Salty Salts

To know if a salt will affect pH, determine: → Will the salt ions will hydrolyze (or split) water?

Conjugates of Strong Acids/Bases: do NOT hydrolyze water, and thus do NOT affect pH

Conjugates of Weak Acids/Bases: do hydrolyze water, and thus do affect pH!

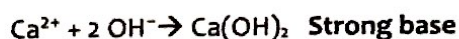
How to Determine the pH of a Salt

Example

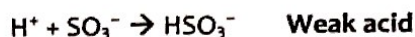
1. Dissociate your salt.



Make the cation into a base: is it strong or weak?



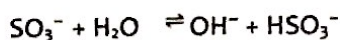
Make the anion into an acid: is it strong or weak?



2. Strong wins!

SB + WA, so this salt is basic!

3. If either is weak, write the hydrolysis reaction:



Conjugate base of WA: $\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{HA}$

Conjugate acid of WB: $\text{BH}^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{B}$

4. Use your hydrolysis equation to calculate the pH using the **Weak Acids/Bases** method.

Be careful. Did the problem give you K_a , or K_b instead? Do you need to convert based on your hydrolysis reaction? Remember: $K_w = K_a \times K_b = 1.0 \times 10^{-14}$

Practice: Identify the salt solutions below as acidic, basic, or neutral and justify your answer.

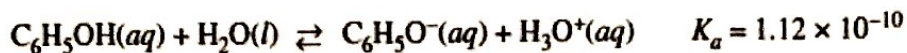
Salt	Parents	Acidic, basic, or neutral?	Justify your answer.
KCl	Parent acid: HCl = strong	neutral	Neither the conjugate base of a strong acid (Cl^-) nor the conjugate acid of a strong base (K^+) will hydrolyze water, so pH is unaffected.
Ions? K^+, Cl^-	Parent base: KOH = strong		
NH_4Cl	Parent acid: HCl = strong	acidic	The conjugate base of a strong acid (Cl^-) will not hydrolyze H_2O + will not affect pH, but the conjugate acid of a weak base will hydrolyze H_2O + produce a sol'n w/ $\text{pH} < 7$: $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3$
Ions? $\text{NH}_4^+, \text{Cl}^-$	Parent base: $\text{NH}_3 =$ weak		

pH vs pK_a: Which form dominates?

Given the generic weak acid reaction: $\text{HA}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{A}^-(\text{aq}) + \text{H}_3\text{O}^+$

- $\text{pH} \leq \text{pK}_a$ the acid form (HA) predominates (plenty of H^+ present in solution)
- $\text{pH} > \text{pK}_a$ the conjugate base form (A⁻) predominates (not enough H^+ present in solution)

Let's Practice! (2016 #4, 4 points)



Phenol is a weak acid that partially dissociates in water according to the equation above.

- What is the pH of a 0.75 M $\text{C}_6\text{H}_5\text{OH}(\text{aq})$ solution?
- For a certain reaction involving $\text{C}_6\text{H}_5\text{OH}(\text{aq})$ to proceed at a significant rate, the phenol must be primarily in its deprotonated form, $\text{C}_6\text{H}_5\text{O}^-(\text{aq})$. In order to ensure that the $\text{C}_6\text{H}_5\text{OH}(\text{aq})$ is deprotonated, the reaction must be conducted in a buffered solution. On the number scale below, circle each pH for which more than 50 percent of the phenol molecules are in the deprotonated form ($\text{C}_6\text{H}_5\text{O}^-(\text{aq})$). Justify your answer.

1 2 3 4 5 6 7 8 9 10 11 12 13 14

$$(a) \quad K_a = \frac{[\text{C}_6\text{H}_5\text{O}^-][\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_5\text{OH}]} = \frac{x^2}{0.75 - x} = 1.12 \times 10^{-10}$$

Assume $x \ll 0.75$

$$x = \sqrt{(1.12 \times 10^{-10})(0.75)} = 9.2 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log(9.2 \times 10^{-6}) = \boxed{5.04}$$

(b) for deprotonated form to dominate $\Rightarrow \text{pH} > \text{pK}_a$

$$\text{pK}_a = -\log(1.12 \times 10^{-10}) = 9.951$$