Solution Equilibria **Name**_____

Determining the pH of Salts

Read from **Lesson 2 Part d:** Determination of the pH of Salts **Acid-Base Equilibria** in the **Chemistry Tutorial** Section**, Chapter 16** of **The Physics Classroom**:

Introduction

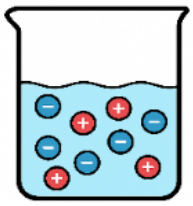
- Hydrolysis is the reaction of salt ions with water, producing either H_3O^+ (acidic) or OH^- (basic).
- Only ions that are conjugates of weak acids or bases hydrolyze.
- A cation may donate a proton to water → H_3O^+ formed → solution becomes acidic.
- Example: NH_4^+ (from NH_4Cl) is the conjugate acid of NH_3,

$$NH_4^+(aq) + H_2O(l) \rightleftarrows NH_3(aq) + H_3O^+(aq)$$

- An anion may accept a proton from water → OH^- formed → solution becomes basic.
  - Example: F^- (from NaF) is the conjugate base of HF, $F^-(aq) + H_2O(l) \rightleftarrows HF(aq) + OH^-(aq)$
- Some ions **do not** hydrolyze because they are derived from strong acids or strong bases, and act as spectator ions in water.

Predicting which ions hydrolyze

Ion Type	Hydrolyzes?	pH Effect
Cation of strong base (e.g., Na^+)	No	Neutral
Anion of strong acid (e.g., Cl^-)	No	Neutral
Conjugate acid of weak base (e.g., NH_4^+)	Yes	Acidic
Conjugate base of weak acid (e.g., F^-)	Yes	Basic



Determining the equilibrium constant for hydrolysis

For an ion that acts as an acid (i.e., a conjugate acid of a weak base), the hydrolysis reaction has a ($\mathbf{K_a}$). That can be calculated with $\mathbf{K_a = K_w/K_b}$

For an ion that acts as a base (i.e., a conjugate base of a weak acid), the hydrolysis reaction uses ($\mathbf{K_b}$). That can be calculated via with $\mathbf{K_b = K_w/K_a}$

Steps to Determine pH of a Salt Solution

- **Identify the ions**: Given the formula or name of the salt, determine the two ions it contains.
- **Assess hydrolysis**: Decide which ion (if any) will undergo hydrolysis. If neither ion hydrolyzes, the solution is neutral with a pH of 7.0.
- **Write the hydrolysis equation**: For the ion that hydrolyzes, write a balanced chemical equation. One of the products will be either H_3O^+ or OH^-.
- **Calculate K_a or K_b**: Use known values (such as Kw and the conjugate acid/base relationship) to calculate the appropriate equilibrium constant for the hydrolysis reaction.
- **Write the equilibrium expression**: Set up the equilibrium constant equation that relates the K_a or K_b to the concentrations of the species involved.
- **Set up an ICE table**: Use an ICE (Initial, Change, Equilibrium) table to express the equilibrium concentrations in terms of a variable, typically (x). Apply the Five-Percent Rule if needed to simplify.
- **Solve for (x)**: Substitute the expressions from the ICE table into the equilibrium equation and solve algebraically for (x). Check that the Five-Percent Rule remains valid.
- **Determine pH**: Use the value of (x) to find the concentration of ($[H_3O^+]$) or ($[OH^-]$), then calculate the pH or pOH accordingly.

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Example: 5.00 g of **potassium acetate** is dissolved enough water to make a **250.0 mL** solution. What is the **pH** of this solution of potassium acetate?

First, find the concentration of $KC_2H_3O_2$.

$$\frac{5.00\text{ g } KC_2H_3O_2}{0.250\text{ L solution}} * \frac{1\text{ mol } KC_2H_3O_2}{98.14\text{ g } KC_2H_3O_2} = \mathbf{0.204\ M\ KC_2H_3O_2}$$

Next, determine which of the two ions in the salt will hydrolyze. $\mathbf{K^+}$ is a spectator ion; it is the cation of a strong base (KOH). $\mathbf{C_2H_3O_2^-}$ is the conjugate base of a weak acid. It will hydrolyze by accepting a proton from water.

The balanced equation for the hydrolysis reaction is: $\mathbf{C_2H_3O_2^-(aq) + H_2O(l) \rightleftarrows HC_2H_3O_2\ (aq) + OH^-(aq)}$

The K_b for this reaction is $\mathbf{K_b = K_w/K_a = 1.0 \times 10^{-14} / 1.8 \times 10^{-5} = 5.6 \times 10^{-10}}$

Now, complete the ICE table:

	$C_2H_3O_2^-$ (aq)	+	H_2O(l)	⇄	$HC_2H_3O_2$ (aq)	+	OH^-
Initial	**0.204**		**Doesn't matter**		**0**		**0**
Change	**-x**		**-x**		**+ x**		**+ x**
Equilibrium	**0.204 -x**		**Doesn't matter**		**x**		**x**

$$\mathbf{K_b = \frac{[HC_2H_3O_2] * [OH^-]}{[C_2H_3O_2^-]} = 5.6 \times 10^{-10} = \frac{x * x}{0.204 - x}} \Rightarrow \mathbf{x^2 = (5.6 \times 10^{-10} * 0.204)} \Rightarrow \mathbf{x = 1.1 \times 10^{-5}\ M}$$

$\mathbf{x = [OH^-] = 1.1 \times 10^{-5}\ M}$

$\mathbf{pOH = -\log [OH^-] = -\log (1.1 \times 10^{-5} M) = 4.97}$

$\mathbf{pH = 14 - pOH = 14 - 4.97 = 9.03}$

Questions

Part 1: Acid, Basic or Neutral Salts.

For each salt below, decide how it affects the pH of water at 25 °C. Circle one: **Acidic**, **Basic**, or **Neutral**

Salt	Type of Solution:		
1. $NaNO_2$	**Acidic**	**Basic**	**Neutral**
2. $(CH_3)_2NH_2Cl$	**Acidic**	**Basic**	**Neutral**
3. $BaBr_2$	**Acidic**	**Basic**	**Neutral**
4. KCN	**Acidic**	**Basic**	**Neutral**
5. $Ca(ClO_4)_2$	**Acidic**	**Basic**	**Neutral**
6. NH_4ClO_3	**Acidic**	**Basic**	**Neutral**
7. NaOBr	**Acidic**	**Basic**	**Neutral**
8. $RbNO_3$	**Acidic**	**Basic**	**Neutral**

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Part 2 – Hydrolysis and pH Calculations

For each salt listed in Part 1:

a. Write the hydrolysis reaction that explains its pH behavior. If the salt is neutral, just write: "Neutral salt"

b. Calculate the pH when 5.00 g of the salt is dissolved in 250.0 mL of water. *Use the potassium acetate example as a guide.*

1. $NaNO_2$

a.

b.

2. $(CH_3)_2NH_2Cl$

a.

b.

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3. $BaBr_2$

a.

b.

4. KCN

a.

b.

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5. **$Ca(ClO_4)_2$**
   a.

   b.

6. **NH_4ClO_3**
   a.

   b.

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7. NaOBr

a.

b.

8. $RbNO_3$

a.

b.

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