Solution Equilibria **Name**________________________________

Common Ion Effects and Buffers

Read from **Lesson 2 Part e: Common Ion Effects and Buffers** in the **Chemistry Tutorial** Section**, Chapter 16** of **The Physics Classroom**:

What Is a Common Ion?

- A common ion is shared between a weak acid/base and a salt. Example: HF in NaF solution → both contain F^-. The presence of a common ion suppresses further dissociation of the weak acid/base.

Common Ion Effect

- Adding a common ion shifts equilibrium toward reactants (Le Chatelier's Principle).
- Result: Lower $[H_3O^+]$ or $[OH^-]$, higher pH (less acidic) or lower pH (less basic).
  - Example: 0.50 M HF alone has a pH ≈ 1.74, but 0.50 M HF + 0.50 M NaF has a pH ≈ 3.18
- Why does pH go up (less acidic) when a common ion is present? Because the presence of F^- suppresses HF dissociation, so fewer H_3O^+ are produced.

Le Chatelier's Principle in Action

- Adding a product ion (e.g., F^-) shifts equilibrium left.
- Less dissociation → lower $[H_3O^+]$ or $[OH^-]$ → altered pH.

Henderson-Hasselbalch Equation

- Used to calculate pH of buffer systems: $\mathbf{pH = pK_a - log ([HA] / [A^-])}$
- Shows how changing $[A^-]$ or $[HA]$ affects pH.
- The role of $[A^-]$: As $[A^-]$ increases (common-ion concentration increases), dissociation of HA decreases, so $[H_3O^+]$ (x) decreases → higher pH (less acidic). **(Increasing $[A^-]$ → higher pH (more basic)**
- The role of [HA]: As [HA] increases (common-ion concentration decreases), dissociation of HA increases, so $[H_3O^+]$ (x) increases → lower pH (more acidic). (**Increasing [HA] → lower pH (more acidic))**

Buffer Systems

- A buffer is a solution of a weak acid/base and its conjugate.
- It resists pH changes when small amounts of acid/base are added.

How Buffers Work

- Mechanism (acid/conjugate base buffer):
- If a strong acid (H_3O^+) is added → the conjugate base in the buffer reacts with it, reducing the increase in $[H_3O^+]$.
- If a strong base (OH^-) is added → the weak acid in the buffer donates H^+ to neutralize some OH^-, reducing the decrease in $[H_3O^+]$.
- The net effect: pH changes only modestly, rather than dramatically.

Problem-Solving Steps

- Identify the weak acid + its conjugate base OR the weak base + its conjugate acid.
- Write the dissociation equation.
- Write the equilibrium constant expression (K_a or K_b).
- Set up an ICE table: Initial concentrations (including the common ion), Change, Equilibrium.
- Substitute equilibrium concentrations into the constant expression.
- Solve for **x** (change, or the $[H_3O^+]$ or $[OH^-]$). Check approximation (e.g., $x < 5\%$of initial).
- Compute pH (or pOH) from the relevant ion concentration.
- Interpret the effect of the common ion: compared to the system without common ion, dissociation is suppressed; pH shifts accordingly.

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Questions

1. Define in your own words the **common ion effect**. Then explain why adding a soluble salt of a conjugate base (e.g., **$KClO_2$)** to a weak acid solution (**$HClO_2$)** causes the pH to increase.

2. **For each pair of compounds listed below, identify the common ion.**

- If the common ion acts as a **spectator ion**, label it as *spectator ion*.
- If the common ion **undergoes hydrolysis in water**, write the balanced equation for its hydrolysis reaction.

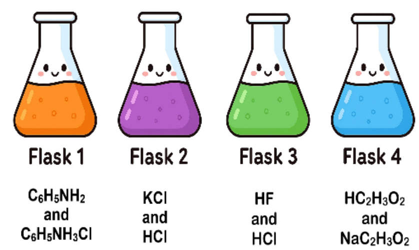
a. equal volumes of **1 M NaBr** and **1 M KBr**

b. equal volumes of **1 M H_2SO_3** and **1 M $KHSO_3$**

c. equal volumes of **1 M CsCl** and **1 M HCl**

d. equal volumes of **1 M H_3PO_4** and **1 M NaH_2PO_4**

3. Bella Buffer and Aaron Agin are reviewing their chemistry notes after a lesson on buffer solutions. Their teacher presented four examples of solutions, each in a separate flask, and asked the class to determine which ones are buffers. Aaron claims that all four solutions are buffers, but Bella disagrees. **How can Bella explain to Aaron why not all the solutions are buffers?** *Note: Each flask contains 100.0 mL of an equimolar mixture of the compounds listed.*



Flask 1: $C_6H_5NH_2$ and $C_6H_5NH_3Cl$
Flask 2: KCl and HCl
Flask 3: HF and HCl
Flask 4: $HC_2H_3O_2$ and $NaC_2H_3O_2$

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4. **A 0.100 M solution of propanoic acid ($HC_3H_5O_2$)** is prepared. *(The K_a for propanoic acid is 1.3×10^{-5})*
   a. Write the balanced chemical equation for the dissociation of propanoic acid in water. Then, write the equilibrium-constant expression for this reaction.

   b. Calculate the pH of the 0.100 M propanoic acid solution.

   c. A mass of 0.500 g of sodium propanoate (**$NaC_3H_5O_2$**) is added to 250.0 mL of the 0.100 M propanoic acid solution. How many **moles** of **$NaC_3H_5O_2$** were added? What is the **concentration** of **$NaC_3H_5O_2$?**

   Assuming the volume of the solution remains unchanged, calculate the following:
   d. The new concentration of the propanoate ion, $C_3H_5O_2^-$(aq), in the solution.

   e. The new concentration of the hydronium ion, H_3O^+ (aq), in the solution.

   f. The new pH of the solution.

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5. **A 0.250 M solution of pyridine (C_5H_5N)** is prepared. *(The K_b for pyridine is 1.5×10^{-9})*

   a. Write the balanced chemical equation for the dissociation of pyridine in water. Then, write the equilibrium-constant expression for this reaction.

   b. Calculate the pH of the 0.250 M solution of pyridine solution.

   c. A mass of 0.350 g of pyridinium chloride (**C_5H_6NCl**) is added to 500.0 mL of the 0.250 M solution of pyridine. How many **moles** of **C_5H_6NCl** were added? What is the **concentration** of **C_5H_6NCl?**

   Assuming the volume of the solution remains unchanged, calculate the following:

   d. The new concentration of the pyridinium ion, $C_5H_6N^+(aq)$, in the solution.

   e. The new concentration of the hydroxide ion, OH^- (aq), in the solution.

   f. The new pH of the solution.