Which of the following combinations could make a buffer solution?

a. $\text{HClO}_4, \text{KClO}_4$

b. $\text{HMnO}_4, \text{LiMnO}_4$

c. $\text{C}_6\text{H}_5\text{N}, \text{C}_6\text{H}_5\text{NH}^+\text{Br}^-$

d. $\text{HI}, \text{NaI}$
17. Calculate the pH of each of the following solutions.
   a. 0.100 \( M \) propanoic acid (\( \text{HC}_3\text{H}_5\text{O}_2 \), \( K_a = 1.3 \times 10^{-5} \))
   b. 0.100 \( M \) sodium propanoate (\( \text{NaC}_3\text{H}_5\text{O}_2 \))
   c. pure \( \text{H}_2\text{O} \)
   d. a mixture containing 0.100 \( M \) \( \text{HC}_3\text{H}_5\text{O}_2 \) and 0.100 \( M \) \( \text{NaC}_3\text{H}_5\text{O}_2 \)

19. Compare the percent dissociation of the acid in Exercise 17a with the percent dissociation of the acid in Exercise 17d. Explain the large difference in percent dissociation of the acid.
21. Calculate the pH after 0.020 mol HCl is added to 1.00 L of each of the four solutions in Exercise 17.

23. Calculate the pH after 0.020 mol NaOH is added to 1.00 L of each of the four solutions in Exercise 17.

25. Which of the solutions in Exercise 17 shows the least change in pH upon the addition of acid or base? Explain.
15. A certain buffer is made by dissolving NaHCO₃ and Na₂CO₃ in some water. Write equations to show how this buffer neutralizes added H⁺ and OH⁻.
31. Calculate the pH of each of the following buffered solutions.
   a. 0.10 \text{ M} \text{ acetic acid}/0.25 \text{ M} \text{ sodium acetate}
   b. 0.25 \text{ M} \text{ acetic acid}/0.10 \text{ M} \text{ sodium acetate}
   c. 0.080 \text{ M} \text{ acetic acid}/0.20 \text{ M} \text{ sodium acetate}
   d. 0.20 \text{ M} \text{ acetic acid}/0.080 \text{ M} \text{ sodium acetate}
33. Calculate the pH of a buffer solution prepared by dissolving 21.5 g benzoic acid (HC\textsubscript{7}H\textsubscript{5}O\textsubscript{2}) and 37.7 g sodium benzoate in 200.0 mL of solution.
37. Calculate the mass of sodium acetate that must be added to 500.0 mL of 0.200 M acetic acid to form a pH = 5.00 buffer solution.
16. A buffer is prepared by dissolving HONH$_2$ and HONH$_3$NO$_3$ in some water. Write equations to show how this buffer neutralizes added H$^+$ and OH$^-$. 
32. Calculate the pH of each of the following buffered solutions.
   a. 0.50 \textit{M} \text{C}_2\text{H}_5\text{NH}_2/0.25 \textit{M} \text{C}_2\text{H}_5\text{NH}_3\text{Cl}
   b. 0.25 \textit{M} \text{C}_2\text{H}_5\text{NH}_2/0.50 \textit{M} \text{C}_2\text{H}_5\text{NH}_3\text{Cl}
   c. 0.50 \textit{M} \text{C}_2\text{H}_5\text{NH}_2/0.50 \textit{M} \text{C}_2\text{H}_5\text{NH}_3\text{Cl}
34. A buffered solution is made by adding 50.0 g NH₄Cl to 1.00 L of a 0.75 \textit{M} solution of NH₃. Calculate the pH of the final solution. (Assume no volume change.)
35. Calculate the pH after 0.010 mol gaseous HCl is added to 250.0 mL of each of the following buffered solutions.
   a. 0.050 \text{ M} \text{ NH}_3/0.15 \text{ M} \text{ NH}_4\text{Cl}
   b. 0.50 \text{ M} \text{ NH}_3/1.50 \text{ M} \text{ NH}_4\text{Cl}

Do the two original buffered solutions differ in their pH or their capacity? What advantage is there in having a buffer with a greater capacity?
39. Consider a solution that contains both C₅H₅N and C₅H₅NHNO₃. Calculate the ratio \([C₅H₅N]/[C₅H₅NH⁺]\) if the solution has the following pH values.

a. pH = 4.50    c. pH = 5.23
b. pH = 5.00    d. pH = 5.50
41. Consider the acids in Table 14.2. Which acid would be the best choice for preparing a pH = 7.00 buffer? Explain how to make 1.0 L of this buffer.
43. Calculate the pH of a solution that is $0.40 \ M \ H_2NNH_2$ and $0.80 \ M \ H_2NNH_3NO_3$. In order for this buffer to have $\text{pH} = pK_a$, would you add HCl or NaOH? What quantity (moles) of which reagent would you add to 1.0 L of the original buffer so that the resulting solution has $\text{pH} = pK_a$?
45. Which of the following mixtures would result in buffered solutions when 1.0 L of each of the two solutions are mixed?

a. 0.1 M KOH and 0.1 M CH₃NH₃Cl
b. 0.1 M KOH and 0.2 M CH₃NH₂
c. 0.2 M KOH and 0.1 M CH₃NH₃Cl
d. 0.1 M KOH and 0.2 M CH₃NH₃Cl
47. What quantity (moles) of NaOH must be added to 1.0 L of 2.0 \( M \) HC\(_2\)H\(_3\)O\(_2\) to produce a solution buffered at each pH?

\[ \text{a. } \text{pH} = pK_a \quad \text{b. } \text{pH} = 4.00 \quad \text{c. } \text{pH} = 5.00 \]
Titrations and pH Curves

A titration is a technique used to determine the amount of acid or base present in a solution.

This process involves delivering a solution of known concentration from a burette (the titrant) into a solution of unknown concentration (the analyte) until the substance being analyzed is consumed at the equivalence (stoichiometric) point.

This can be signaled by a color change in an indicator or by using an electronic pH meter:

When a pH meter is used, a titration (pH) curve is generated which is a plot of pH against titrant volume.
What is the pH at the equivalence point when 50.0 mL of 0.0200 M acetic acid is titrated with 0.100 M sodium hydroxide?
What is the pH at the equivalence point when 50.0 mL of 0.0100 M magnesium hydroxide is titrated with 0.0500 M chlorous acid?
What is the pH at the equivalence point when 60.0 mL of 0.0200 M hydrochloric acid is titrated with 0.100 M methylamine?
What is the pH at the equivalence point when 45.0 mL of 0.0300 M ethylamine is titrated with 0.0500 M hydrobromic acid?
Bromthymol blue, an indicator with a $K_a$ value of $1.0 \times 10^{-7}$, is yellow in its HIn form and blue in its In$^-$ form. Suppose we put a few drops of this indicator in a strongly acidic solution. If the solution is then titrated with NaOH, at what pH will the indicator color change first be visible?
63. Two drops of indicator HIn ($K_a = 1.0 \times 10^{-9}$), where HIn is yellow and In$^-$ is blue, are placed in 100.0 mL of 0.10 $M$ HCl.

a. What color is the solution initially?

b. The solution is titrated with 0.10 $M$ NaOH. At what pH will the color change (yellow to greenish yellow) occur?

c. What color will the solution be after 200.0 mL NaOH has been added?
Methyl orange, an indicator with a $K_a$ value of $1.8 \times 10^{-4}$ is pink in its HIn form and yellow in its In$^-$ form. Suppose we put a few drops of this indicator in a strongly basic solution. If the solution is then titrated with HCl, at what pH will the color change first become visible?
71. Estimate the pH of a solution in which bromcresol green is blue and thymol blue is yellow. (See Fig. 15.8.)
73. A solution has a pH of 7.0. What would be the color of the solution if each of the following indicators were added? (See Fig. 15.8.)
   a. thymol blue  
   b. bromthymol blue  
   c. methyl red  
   d. crystal violet
When a 0.100 M hydrocyanic acid solution is titrated with a 0.100 M NaOH solution. What would be a good indicator to use for this titration?
If 100.0 mL of 0.10 M ammonia, NH₃, $K_b = 1.8 \times 10^{-5}$ is to be titrated with a 0.10 M HCl solution. What would be a good indicator to use for this titration?
Potassium hydrogen phthalate, known as KHP (molar mass = 204.22 g/mol), can be obtained in high purity and is used to determine the concentration of solutions of strong bases by the reaction

\[
\text{HP}^-(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{P}^2-(aq)
\]

If a typical titration experiment begins with approximately 0.5 g KHP and has a final volume of about 100 mL, what is an appropriate indicator to use? The pKₐ for HP⁻ is 5.51.