LESSON 123 CLASSWORK

How Colorful
Applying Le Châtelier's Principle

Name ____________________

Date ____________ Period ________

Purpose

To explore equilibrium with acid-base indicators.

Acid-Base Indicators

Acid-base indicators generally are weak acids or bases. The general reversible reaction for an indicator that is a weak acid can be written as

$$\mathrm{H}In(aq) \rightleftharpoons \mathrm{H}^+(aq) + In^-(aq)$$

Each indicator has a specific value for the equilibrium constant, K, and a certain color depending on the pH of the solution. Use the information in the table for three indicators to answer the questions below.

Indicator	H*In* color	*In*⁻ color	K	Color at Different pH values		
				pH=4	pH=7	pH=10
methyl orange	red	yellow	3.4×10^{-4}	orange	yellow	yellow
bromothymol blue	yellow	blue	5.0×10^{-8}	yellow	green	blue
phenolphthalein	colorless	pink	3.2×10^{-10}	colorless	colorless	pink

Part 1: Applying Le Châtelier's Principle to Acid-base Indicators

1. Explain why a solution of phenolphthalein with pH=7 is colorless.

2. Explain why a solution of methyl orange with pH=7 is yellow.

3. Explain why a solution of bromothymol blue with pH=7 is green.

4. Methyl orange has the largest value of K of the three indicators.
   a. What does a large value of *K* indicate?
   b. Is this consistent with the solution color at pH=7?

5. Phenolphthalein has the smallest value of K of the three indicators. Is this consistent with the solution color at pH=7?

6. Use Le Châtelier's principle to explain why the solution color for bromothymol blue changes from yellow to green to blue as the pH decreases.

7. What could you do to observe a red color with a solution of methyl orange?

8. What could you do to make a solution of phenolphthalein dark pink?

Part 1: Equilibrium Calculations

The equilibrium concentrations for three bromothymol blue solutions are provided in the table. In each case, the pH of the solution is controlled independently such that the addition of H*In* does not change the pH. Use the data in the table to answer the questions below.

pH	$[H^+]$	$[In^-]$	$[HIn]$	$K = \frac{[H^+][I^-]}{[HIn]}$
6	10^{-6} M	0.00073 M	0.00925 M	$K = \frac{(10^{-6})(0.00073)}{(0.00925)} = 7.9\times10^{-8}$
7	10^{-7} M	0.00440 M	0.00560 M	
8	10^{-8} M	0.00888 M	0.00112 M	

1. What happens to the concentration of In^- as the pH increases?

2. What happens to the concentration of H*In* as the pH increases?

3. Determine the equilibrium constant, *K*, for the solutions with pH=7 and pH=8. Enter the calculation into the table. Explain the resulting values for *K*.

4. **Making Sense** Explain how acid-base indicators are used to determine the acidity of a solution.