Acid Base Indicators Handout

name: ______________________________________

Acid Base indicators are usually weak acids. A weak acid is an acid that does not dissociate well in water. When you put a million molecules of $HC_2H_3O_2$ into water, it forms acetic acid. But only about 5% of the molecules will ionize, while the vast majority (nearly 95%) just dissolve because they are polar. With so few ions in solution, the strength of the acid is weak. One mole of these molecules does NOT produce anywhere near 6.02.x 10^{23} H^{+1} ions in solution.

Since acid base indicator molecules are weak acids, so they too dissociate poorly. What's special about these molecules is that their molecules have a color we can see, and the ions that they dissociate into have a different color. We can see with our eyes how far the dynamic equilibrium shifts with the addition of acid or base ions. Forward shifts produce one color, reverse shifts produce a different color.

A simple example from our Table M is phenolphthalein. $HC_{20}H_{13}O_4$ is the formula for this weak acid. In water it forms a dynamic equilibrium, with most of the molecule dissolves but remains "whole", it does not ionize. The molecule is colorless. The few ions that do form, are hot pink in color, but there are so few that they are invisible to your eye, unless there is a big forward shift producing more pink anions.

Putting phenolphthalein into an acid will add H^{+1} ions, resulting in a reverse shift towards more colorless moleculs and less pink ions. Adding more base, which are OH^{-1} ions makes something happen to think hard about right now. The hydroxide ions combine with H^{+1} cations and form water. Adding hydroxides is the same as removing H^{+1} cations, resulting in a forward shift of more pink anions. At pH between 8-9 there is a near "balance" between the acid and base level, the numbers of pink ions and colorless molecules is almost even, giving the solution a light pink color. In chemical symbols:

$$\underset{\text{Colorless molecules}}{HC_{20}H_{13}O_{4(AQ)}} \overset{\text{water}}{\rightleftharpoons} H^{+1}{}_{(AQ)} + \underset{\text{Pink Anions}}{C_{20}H_{13}O_4{}^{-1}{}_{(AQ)}}$$

← ADD ACID H^{+1}

ADD BASE OH^{-1} →

Adding base ions is the same as removing H^{+1} cations; $OH^{-1} + H^{+1} \rightarrow HOH$ *water*

On the second page are the indicators from Table M. Each changes color at a different pH range. Each of their molecules and anions are different colors, which makes them "indicate" a pH for the solution they are in.

Using COLOR pencils, show where the pH color changes occur for each indicator. Use the correct colors for each indicator. Write the pH "range" where this change occurs as well. All pH lower than this "range" are ONE COLOR and all pH above this "range" are the other color. INSIDE the RANGE is where the colors change and sometimes mix: yellow and blue make green, red and blue make purple, colorless and pink make light pink.

Indicators can indicate to you a pH range, but you can't use them to tell you specifically what the pH of the solution actually is. If a solution is PINK with phenolphthalein YOU know that the pH is above 9. It could be 9, or 10, 12.74, or any decimal in between 9 and 14! Pink indicates to you only above pH 9, not what the pH actually is.

	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
methyl orange															
	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
bromthymol blue															
	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
phenolphthalein															
	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
litmus															
	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
bromcresol green															
	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
thymol blue															