4/3/2019

Buffer range + preparation to target pH

best to have equal amts of weak acid/conj. base

$$pH = pK_a + \log\frac{[base]}{[acid]}$$

$$\frac{[base]}{[acid]} = 1 \Rightarrow \log\left(\frac{b}{a}\right) = 0$$

$$\Rightarrow \boxed{pH = pK_a}$$

if [base] > [acid]

let's say it's 10× bigger ... $\frac{[base]}{[acid]} = 10$, $\log(10) = 1$

$$\Rightarrow \boxed{pH = pK_a + 1}$$

if [acid] is 10× larger than [base] ... $\frac{[base]}{[acid]} = 0.1$, $\log(0.1) = -1$

$$\boxed{pH = pK_a - 1}$$

"Effective buffer range": pH: $pK_a - 1 \leq pH \leq pK_a + 1$

ex: let's say we need a buffer of pH 4.25

$HClO_2$, $pK_a = 1.95$ $\quad$ $HCHO_2$, $pK_a = 3.74$ ✓ (only 0.51 high)

HNO_2, $pK_a = 3.34$ $\quad$ $HClO$, $pK_a = 7.54$

$pK_a = -\log(K_a)$

H-H: $pH = pK_a + \log b/a$, so $pH - pK_a = \log b/a$

$$\Rightarrow 4.25 - 3.74 = 0.51 = \log\frac{[b]}{[a]} \Rightarrow \frac{[b]}{[a]} = 10^{0.51} = 3.24$$

acid: $HCHO_2$, $\frac{[b]}{[a]} = 3.24$

$$\frac{[CHO_2^-]}{[HCHO_2]} = 3.24$$

ex: if $[HCHO_2] = 0.10M$, then $[CHO_2^-] = 3.24 \times [HCHO_2]$
$= 0.324M$

CH_2O^- ? formate

$NaCH_2O$ sodium formate

Titration curves + acid-base indicators

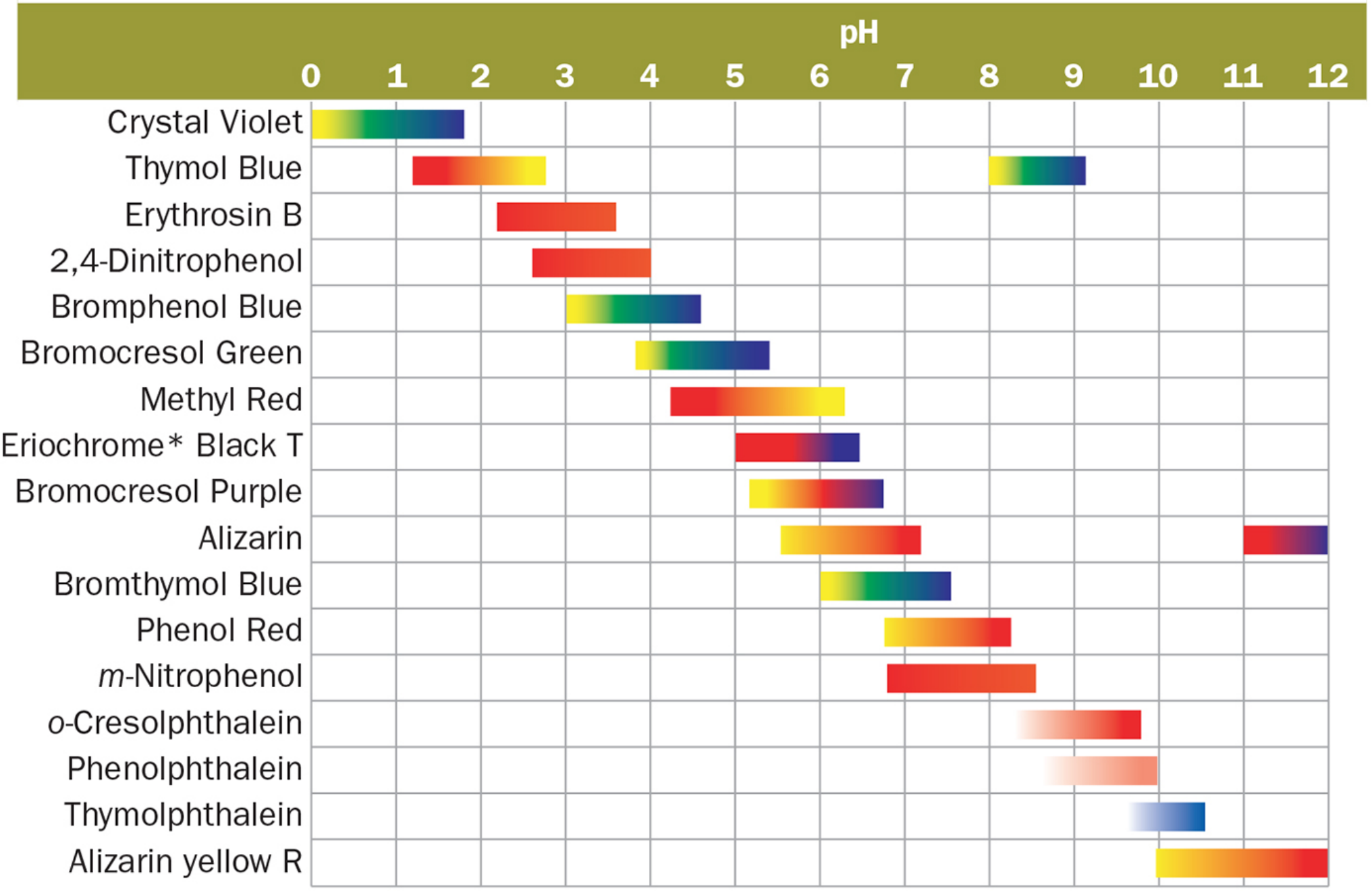
pH indicators are usually weak acids (such as phenolphthalein)

Ka rxn: $HIn(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + In^-(aq)$
color-1 color-2

Using H-H: $pH = pK_a + \log \frac{[In^-]}{[HIn]}$ (b / a)

- $pH = pK_a$, $[In^-]/[HIn] = 1$ equal amounts of 2-colored forms.
- when $pH = pK_a + 1$, $\frac{[In^-]}{[HIn]} = 10$, so 10x more In^- (color-2) than HIn (color-1) — COLOR-2
- when $pH = pK_a - 1$, $\frac{[In^-]}{[HIn]} = 0.1$, so 10x more HIn (color-1) than In^- (color-2) — COLOR-1

TABLE 17.1 Ranges of Color Changes for Several Acid–Base Indicators



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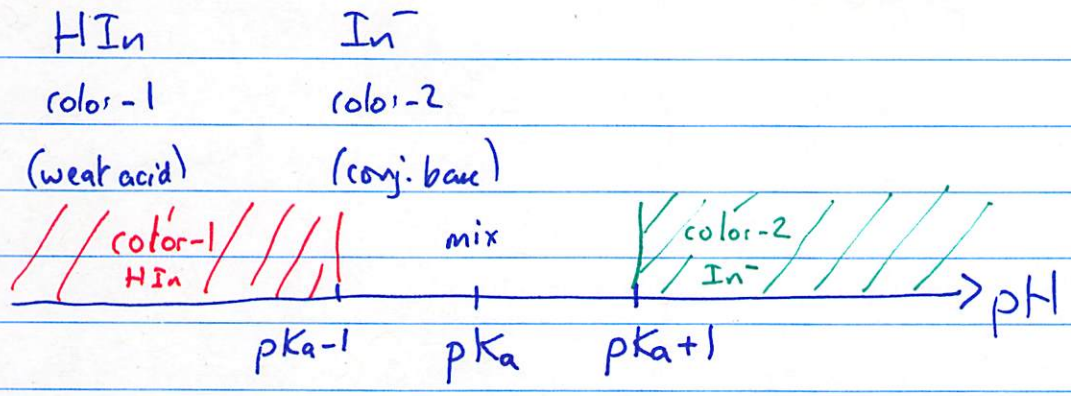
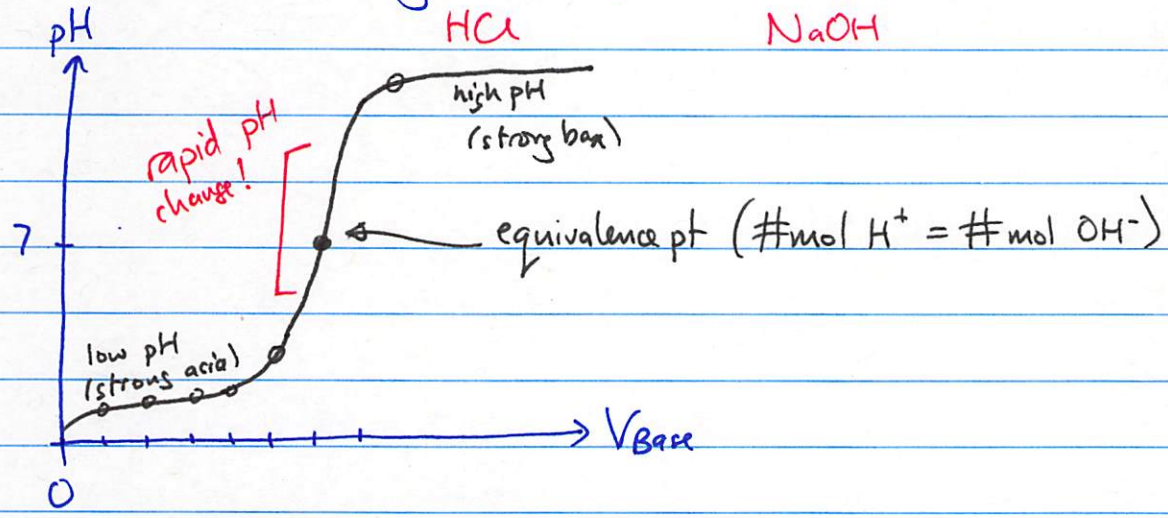


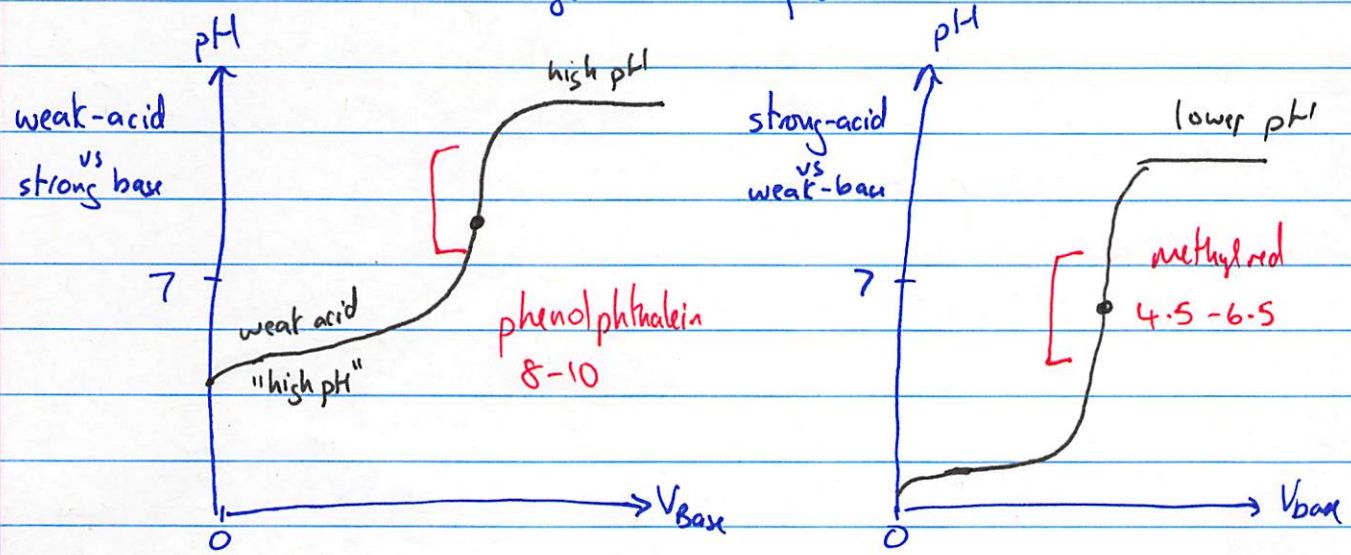
Why useful? During a titration, pH changes!

ex: titrate a strong acid w/ a strong base



A good Indicator: Bromothymol blue, pH: 6-8

color change: "end point"



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