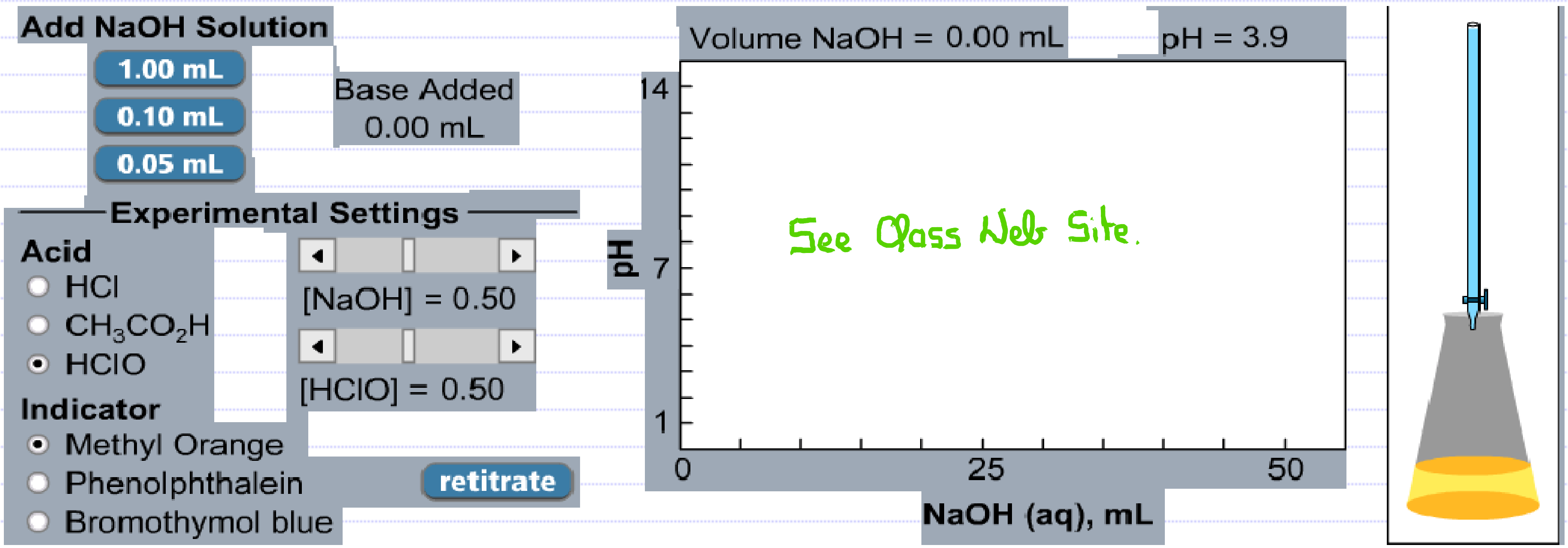
17.3 Acid-Base Titrations

Weak Acid/Strong Base

Choose: 0.5M HClO vs 0.5M NaOH



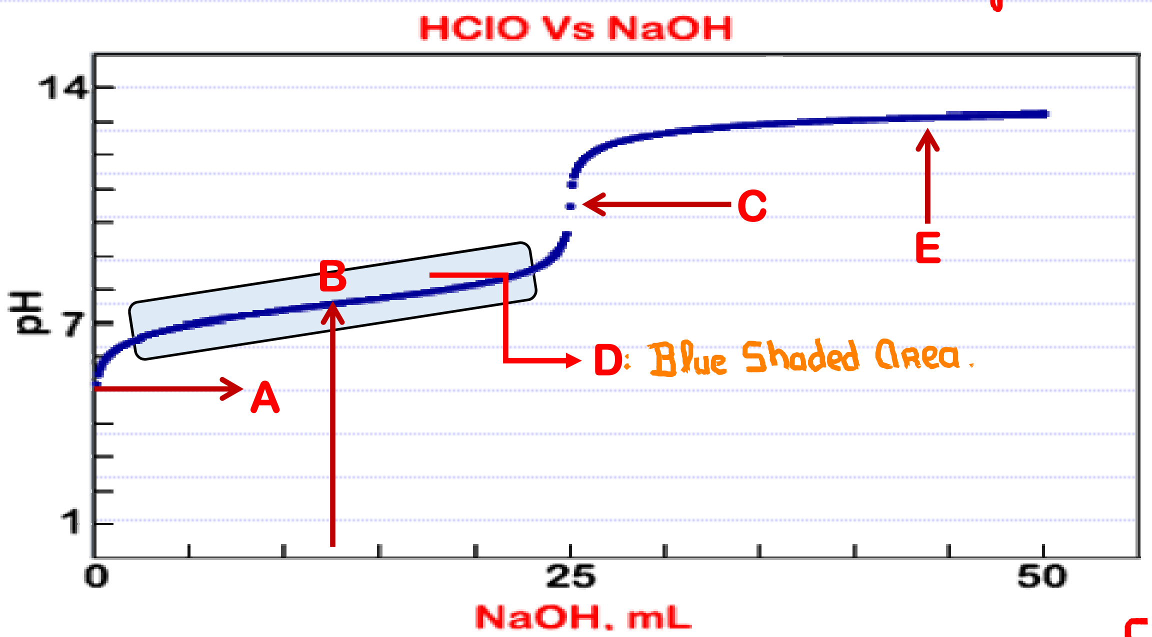
See Class Web Site.

17.3 Acid-Base Titrations

Weak Acid/Strong Base

$HClO_{(aq)} + NaOH_{(aq)} = H_2O_{(l)} + NaClO_{(aq)}$

K_a for $HClO = 3.5 \times 10^{-8}$ @ 25°C



E: $NaOH_{(aq)} = Na^+ + OH^-$

$pOH = -\log_{10}[OH^-]$

$pH = 14 - pOH$

A: $HClO_{(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+ + ClO^-$

B: This is the shaded Buffer Region

Special in that [Buffer Acid] = [Buffer Base]

$pH = pK_a$ of $HClO$

D: The Buffer Region.

$$pH = pK_a + \log_{10}\frac{[\text{Buffer Base}]}{[\text{Buffer Acid}]}$$

C: Equivalence point where the pH is determined by hydrolysis of the salt

$NaClO_{(aq)} = ClO^- + \underline{Na^+}$

Neutral Cation

$ClO^- + H_2O_{(l)} \Leftrightarrow HClO_{(aq)} + OH^-$

$$K_{ClO^-} = \frac{1 \times 10^{-14}}{3.5 \times 10^{-8}} =$$

Calculate pOH

$pH = 14 - pOH$

18.1 Solubility Equilibria and Ksp

The Solubility Product Constant

Compound	K_{sp} at 25 °C
$PbBr_2$	6.3×10^{-6}
$AgBr$	3.3×10^{-13}
$CaCO_3$	3.8×10^{-9}
$CuCO_3$	2.5×10^{-10}
$NiCO_3$	6.6×10^{-9}
Ag_2CO_3	8.1×10^{-12}
$PbCl_2$	1.7×10^{-5}
$AgCl$	1.8×10^{-10}
BaF_2	1.7×10^{-6}
CaF_2	3.9×10^{-11}
$Cu(OH)_2$	1.6×10^{-19}
$Fe(OH)_3$	6.3×10^{-38}
$Ni(OH)_2$	2.8×10^{-16}
$Zn(OH)_2$	4.5×10^{-17}
$Ca_3(PO_4)_2$	1.0×10^{-25}
$CaSO_4$	2.4×10^{-5}
$PbSO_4$	1.8×10^{-8}

$PbBr_2(s) \Leftrightarrow Pb^{2+} + 2Br^-$

Remember that pure liquids, solvents and solids do not appear in the equilibrium expression.

$K = [Pb^{2+}][Br^-]^2$

→ K_{sp} : Solubility Product Constant

Note that the salts listed are those that using Solubility Guide lines in Chem III would have all been deemed insoluble.

Looking at the Ksp values, these are all Reactant-favored equilibria.