Unit 8 – Buffers - Titrations – Main ideas

What is a buffer? A solution that resists a change in pH by adding acid or base.

How can you make a buffer? There are two main ways:

First, take a weak acid and add the salt of the weak acid (that salt is the conjugate base). There is a ratio of acid to conjugate base. The number of moles of acid and base determine how much of an outside pH modifier it can take before excess reactant takes over.

Second, take a weak acid and titrate with a strong base, but not to the equivalence point. That base will neutralize some of that weak acid and make a salt, thus creating a buffer.

Working with millimoles and milliliters (mmol and mL) will simplify the calculations. Not having to move back and forth with mL and L, and not having tiny molarities is a time saver for the AP exam and tests in your class.

Note this example is different than the problem on the next page. It shows good info about the different points in the titration.

Weak Acid + Strong Base Experiment

Volume of 0.20 M NaOH Added (mL)
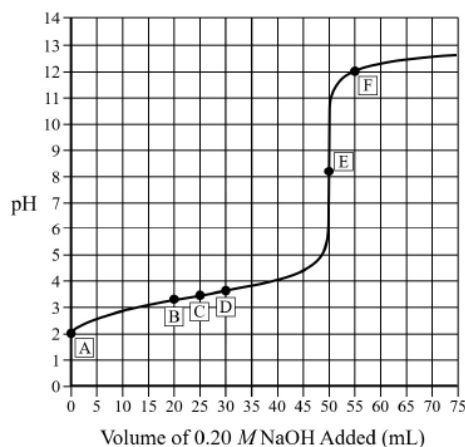
| Point	Description	pH
A	Initial pH of the weak acid, HA	2.05
B	Before the half-equivalence point; $[HA] > [A^-]$ and $pH < pK_a$	3.22
C	At the half-equivalence point; $[HA] = [A^-]$ and $pH = pK_a$	3.40
D	Past the half-equivalence point; $[HA] < [A^-]$ and $pH > pK_a$	3.57
E	At the equivalence point; Switch from K_a to K_b because of the presence of the conjugate base, A^-, which reacts with water as follows: $A^- + H_2O \rightleftarrows HA + OH^-$	8.20
F	Past the equivalence point; There is an excess amount of OH^- in solution.	11.98

Problem # 1 (labeled problem #6 on the video, see next page…)

This is a Michael Farabaugh Video, Unit 8 Summative, time stamp 1:01-1:06. Watch this video segment and be prepared to do a similar analysis with different reactants and different concentrations.

https://www.youtube.com/watch?v=VJ2pkmnUaUs

Take special note of points A, C and E. They are commonly tested topics on AP.

6. A student titrates 50.0 mL of 0.64 M HF(aq) with 2.0 M NaOH(aq), using a probe to monitor the pH of the solution. The data are plotted, producing the following titration curve.

Volume of 2.0 M NaOH Added (mL)
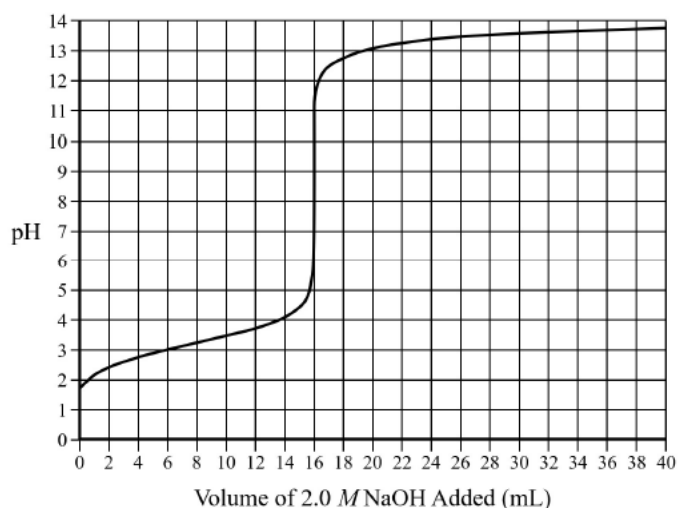
Point A has a 50.0 mL of 0.64 M HF, with a Ka of 7.2×10^{-4}. So just do a Ka calculation with ICE table to find the acid pH:

Point A) Ka=x^2/0.64 = 7.2×10^{-4} and x = 0.0215 and pH = 1.7 (easy peasy)

Point B) 4 mL base added

	HF +	OH- ->	F-
Before (mmol)	32.0	8.0	0
change	-8	-8.0	+8.0
after	24.0 in 50+4 = 54 mL	0.0	8.0
concentration	[HA]= 24.0 / 54 = 0.444 M		[A-]=8.0 / 54 = 0.148

Using a Ka expression instead of the HH equation:

Ka = [H+][F-]/[HF], 7.2×10^{-4} = [H+][0.148] / [0.444] so H+ = 0.00216 , pH = 2.7

Point C) 8 mL added, half titration point:

|	HF +	OH- ->	F-
Before (mmol)	32.0	16.0	0
change	-16	-16.0	+16.0
after	16.0 in 50+8 = 58 mL	0.0	16.0
concentration	[HA]= 16.0 / 58 = 0.276 M		[A-]=16.0 / 58 = 0.276

Using a Ka expression instead of the HH equation:

Ka = [H+][F-]/[HF], 7.2×10^{-4} = [H+][0.276] / [0.276] so H+ = 7.2×10^{-4} , pH = 3.1 (pi)

Shortcut: pH = pKa 😊

Point D) 12 mL added

	HF +	OH- ->	F-
Before (mmol)	32.0	24.0	0
change	-24	-24.0	+24.0
after	8.0 in 50+12 = 62 mL	0.0	24.0
concentration	[HA]= 8.0 / 62 = 0.129 M		[A-]=24.0 / 62 = 0.387

Using a Ka expression instead of the HH equation:

Ka = [H+][F-]/[HF], 7.2×10^{-4} = [H+][0.387] / [0.129] so H+ = 2.4×10^{-4} , pH = 3.6

Point E) 16 mL added – equivalence point 😊

You have 32.0 mmol of F-, do a Kb expression…

Since Ka * Kb = Kw, you can find the base equilibrium of fluoride hydrolysis:

F- + H_2O ↔ HF + OH-   and Kb = [HF][OH-] / [F-]
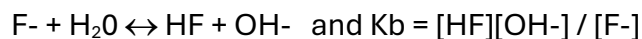
And [F-] = 32.0 mmol / 66 mL = 0.4848 at equivalence point

Kb = 10^{-14} / 7.2×10^{-4} = 1.39×10^{-11} and Kb = x^2 / 0.4848 , so x = [OH-] = 2.6×10^{-6}

And pOH = log (2.6×10^{-6}) = 5.6   and pH = 14-5.6 = 8.4 (whew!!)

Point F) 20 mL added, excess base, (easy peasy)

20mL * 2.0 M = 40 mmol base, which is 8.0 mmol excess OH,

[OH-] = 8.0mmol / 70 mL = 0.114, pOH = 0.9, pH = 13.1

Now, you do one 😊

Problem # 2

In this problem, there is a different molarity of acid and base.

Consider the titration of 100.0 mL of 0.200 M acetic acid (Ka = 1.8×10^{-5}). Adding different amounts of 0.100 M KOH will show different regions of the titration curve:

A) 0.0 mL
B) 50.0 mL
C) 100.0 mL
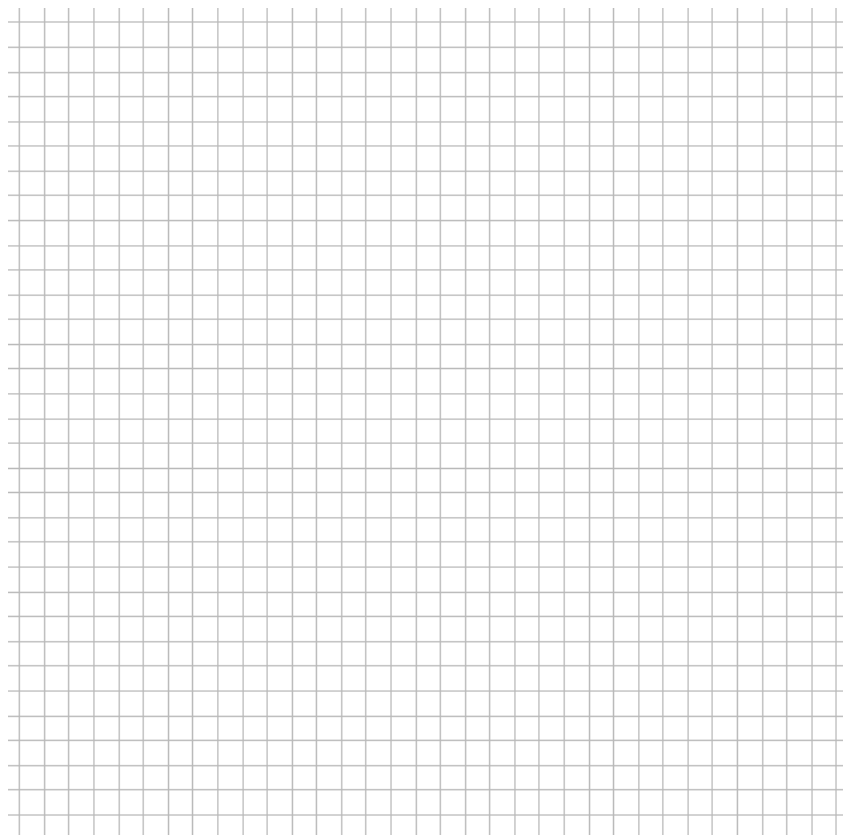D) 150.0 mL
E) 200.0 mL
F) 250.0 mL

1. Show the calculations and show the titration curve at each point, A-F. (room below)

2. What is a buffer? (ANSWER HERE)

3. Discuss the buffer capacity and buffer ratio in points B,C,D.

        B.

        C.

        D.

1. Show the titration curve at each point, A-F.

2. Show the calculations at points A-F

A.

B.

C.

D.

E.

F.