

TITRATION ACTIVITY

Name KEY Date _____ Period _____

In this activity we will explore titration, pH curves and acid-base indicators. We will examine two types of titrations: strong acid-strong base titration (relatively simple) and strong acid-weak base/weak acid-strong base titrations (a lot more involved).

Before we get started, there are some terms and language you will need to know and understand:

titrant - one in the buret with known concentration

buret - calibrated volumetric instrument for dispensing a liquid

indicator - substance which changes color in different $[H_3O^+]$

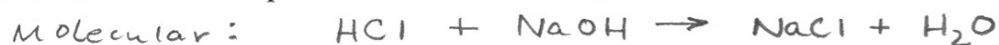
titration curve - a plot of pH vs volume of acid (or base) added

endpoint - point at which the indicator changes color

equivalence point - point where moles H_3O^+ = moles OH^-

Let's first examine strong acid-strong base titrations.

Write the molecular equation and the net ionic equation for the reaction of HCl with NaOH.



The reaction between a strong acid and strong base goes to completion! (Single arrow)

Strong acid-strong base titration calculations are relatively simple because one of three possibilities result after the reaction is complete:

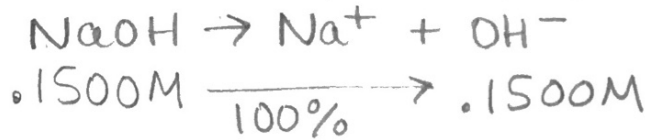
- Excess strong acid remains
- Excess strong base remains
- Equivalence point is reached

Let's try the following problem:

Calculate the pH after the following total volumes of 0.2500 M hydrochloric acid have been added to 50.00 mL of 0.1500M sodium hydroxide. Again it would be useful to calculate the initial moles or millimoles of acid and base reacting (from the molarity and volume), do the math to see which runs out, and then calculate the molarity of the excess and the resulting pH.

Note we are adding certain volumes of 0.2500M acid from a buret to an Erlenmeyer flask containing 50.00 mL of 0.1500M sodium hydroxide.

a) 0.00 mL of acid added (this means just plain strong base):



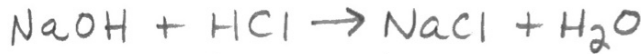
$$\text{pOH} = -\log .1500 = .8239$$

$$\text{pH} = 13.1461$$

b) 4.00 mL of acid added:

$$50.00\text{mL} \times .1500\text{M} \quad 4.00\text{mL} \times .2500\text{M}$$

Reaction:



I	7.500 mmol	1.00 mmol
C	$\frac{-1.00}{\quad}$	$\frac{-1.00}{\quad}$
E	6.50	0

I am going to work in mmoles, but you can work in moles if you prefer. I think mmoles are easier and I just like to say "millimole" ☺

millimoles = milliliters x molarity

Then just multiply mmoles by total ml in flask and you are back to molarity!

Do we have excess? Yes \rightarrow NaOH is excess
Calculate molar concentration of excess:

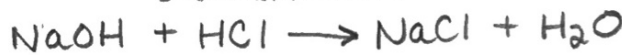
$$\frac{6.50 \text{ mmol}}{54.00 \text{ mL}} = .120 \text{ M NaOH}$$

Calculate pH: $\text{pOH} = -\log .120 = .921$
 $\text{pH} = 13.079$

c) 30.00 mL of acid added:

$$30.00\text{mL} \times .2500\text{M}$$

Reaction:



I	7.500	7.500
C	$\frac{-7.500}{\quad}$	$\frac{-7.500}{\quad}$
E	0	0

Do we have excess? NO

What do we call the point in a titration where moles acid = moles base?

equivalence point

In this titration of strong acid and strong base, what species is left at this point?

neutral salt and water
↓
NaCl

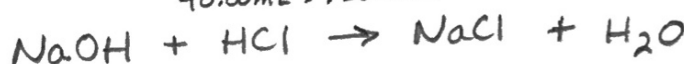
What is the pH?

$$\text{pH} = 7.00 \text{ (pH of pure H}_2\text{O)}$$

d) 40.00 mL of acid added:

$$40.00\text{mL} \times .2500\text{M}$$

Reaction:



I	7.500	10.00
C	$\frac{-7.500}{\quad}$	$\frac{-7.500}{\quad}$
E	0	2.50

Do we have excess? *yes* → HCl

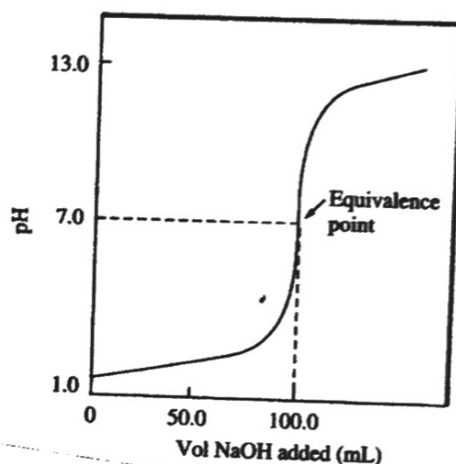
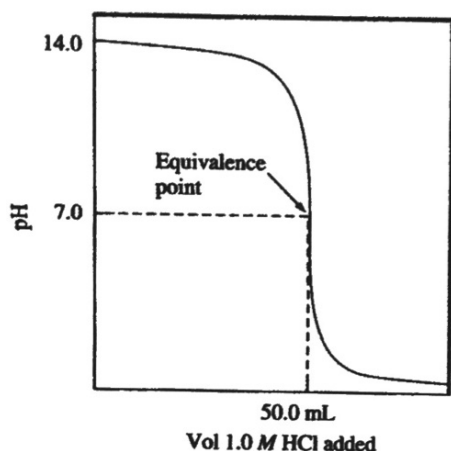
Calculate molar concentration of excess:

$$\frac{2.50 \text{ mmol}}{90.00 \text{ mL}} = .0278 \text{ M HCl}$$

Calculate pH: $\text{pH} = -\log .0278 = 1.556$

In the titration of a strong acid and a strong base the pH at equivalence = 7.00 because the only major species that remains is water (and the pH of water = 7.00 at 25.0°C)!

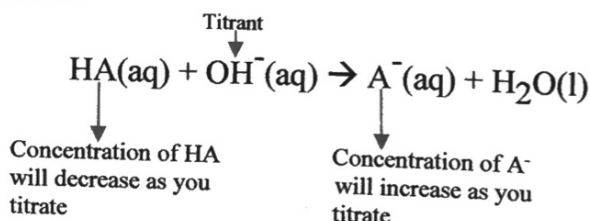
If we graphed the pH vs volume of acid added for the previous titration (starting with strong base), we would get the following curve shown below on the left:



Make special note that when strong acid is added the pH does not dramatically change until almost at the equivalence, and we observe an almost vertical line change in pH. Very near the equivalence point, 2-3 drops of acid results in a dramatic change in pH. We get the same type of curve if we started with strong acid and added strong base, as shown in the graph on the right.

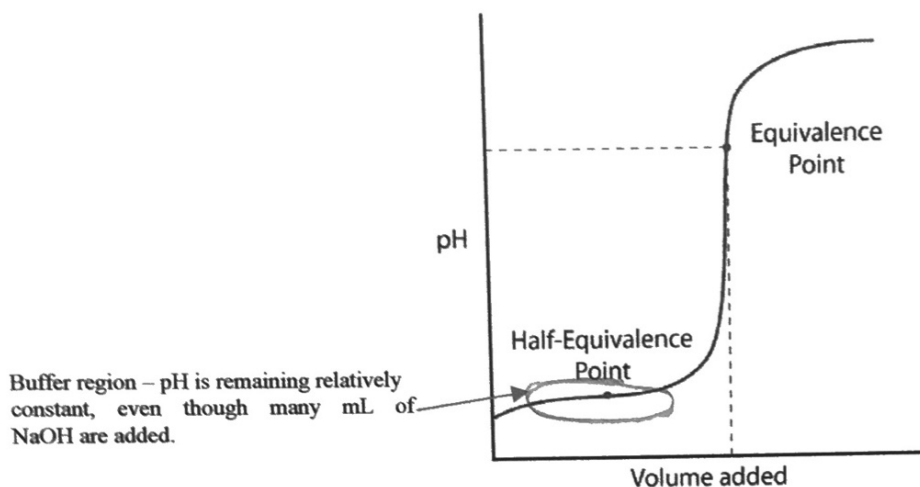
Now let's consider a little more challenging titration, that of a **WEAK ACID with a STRONG BASE** (weak base-strong acid titrations involve the same ideas).

When doing calculations for this type of titration it is useful to relate our thinking back to our buffered solution problems.

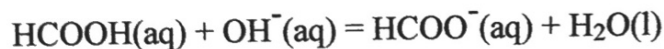


You will be generating a buffer (weak acid (HA) and conjugate base (A⁻)) as a result of the addition of OH⁻ ions to the solution. You will have a buffered solution until you exceed the buffer capacity of the solution, and then the pH will rise dramatically.

A graph of this type of titration – weak acid titrated with strong base – looks like this:

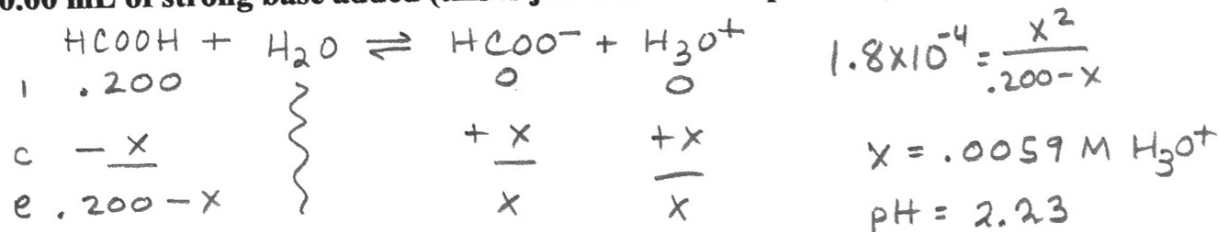


Let's use the titration of formic acid, HCOOH ($K_a = 1.8 \times 10^{-4}$) with NaOH as an illustrative example.



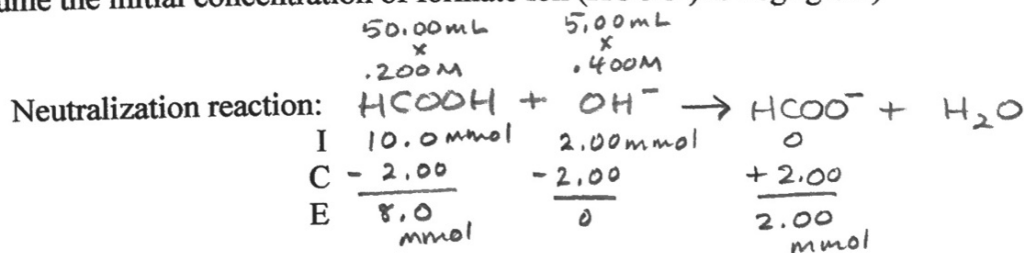
We fill a buret with 0.400M sodium hydroxide and we will slowly add (drop by drop) the following volumes of this base to an Erlenmeyer flask containing 50.00 mL of 0.200M formic acid. Calculate the pH after the following total volumes of the NaOH is added to the acid in the flask:

a) 0.00 mL of strong base added (this is just a weak acid problem):

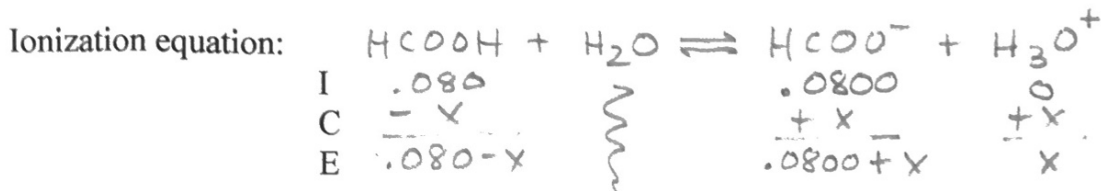


b) 5.00 mL of strong base added:

(note: you can assume the initial concentration of formate ion (HCOO⁻) is negligible)



I am going to work in mmoles, but you can work in moles if you prefer. I think mmoles are easier – and I just like to say “millimole” ☺
millimoles = milliliters x molarity
 Then just multiply mmoles by total ml in flask and you are back to molarity!



Use K_a to calculate new $[\text{H}_3\text{O}^+]$ and pH:

$$1.8 \times 10^{-4} = \frac{(.080+x)(x)}{.0800+x}$$

$$x = 1.79 \times 10^{-4} \text{ M } \text{H}_3\text{O}^+$$

$$\text{pH} = 3.75$$

Alternately, you can use Henderson-Hasselbalch to calculate pH. (Yay Henderson-Hasselbalch!)

Do that here:

$$\text{pH} = \text{p}K_a + \log \frac{B}{A}$$

$$= 3.74 + \log \frac{5.0}{5.00}$$

← just use mmoles since in same volume!

$$= 3.74$$

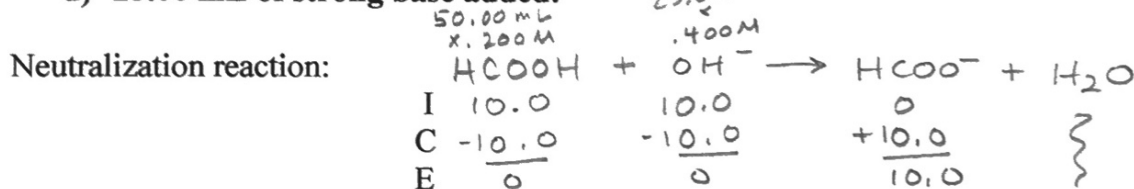
From the K_a expression you noticed that $[\text{HCOOH}] = [\text{HCOO}^-]$ and the $K_a = [\text{H}_3\text{O}^+]$. If

$K_a = [\text{H}_3\text{O}^+]$, then **$\text{p}K_a = \text{pH}$** (This is always true at the half-equivalence point! Very useful for identification of an unknown acid!)

$$K_a = \frac{[\text{HCOO}^-][\text{H}_3\text{O}^+]}{[\text{HCOOH}]} \Rightarrow K_a = [\text{H}_3\text{O}^+]$$

$$\text{p}K_a = \text{pH}$$

d) 25.00 mL of strong base added:



Compare the moles of acid titrated with the moles of base added at this point in the titration. What do you notice? They are equal.

What do we call this point in the titration? Equivalence point
How many moles of acid and moles of base are left? zero

What species is left in solution at this point? HCOO^- (SCB)

What does the remaining species do in water (donate or accept a proton)?

proton acceptor

Write the reaction of this species (should be a weak base) in water: hydrolysis reaction

