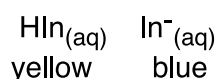


Titrations: This is where you figure out the concentration of something in a solution. It can be H^+ , OH^- , or some ion that you precipitate out of a solution by adding a known amount of reactant to it and somehow monitoring the extent of the rxn.

We'll get to titrations, but first you need a little background about how pH is monitored during a titration.

Indicators --> These are molecules that change color depending upon whether a hydrogen is attached to them or not.

4 XMPL: Toromethyl blue is an indicator that changes from yellow to blue



In stands for indicator. The solution color will depend upon the mix of yellow and blue species which is in turn dependent upon the solution pH.

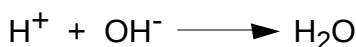
The indicator reaction is : $\text{HI}_{(\text{aq})} \rightleftharpoons \text{H}^{+}_{(\text{aq})} + \text{In}^{-}_{(\text{aq})}$

$$K_a = \frac{[\text{H}^{+}][\text{In}^{-}]}{[\text{HI}]}$$

depending upon pH the ratio $[\text{In}^{-}]/[\text{HI}]$ will change to satisfy K_a . When choosing an indicator try to match the pK_a of the indicator (you know the negative log of the K_a -- it gives you the point of color change) with the equivalence point of the acid-base titration. (You'll understand equivalence point soon enough. Just remember that you can't use the same indicator for every titration you have to select it accordingly.)

A. Strong Acid—Strong Base Titrations (easy)

For titrations where you use a strong acid and base the acid and base will completely dissociate and the net reaction is:



The titration is just about reacting H^{+} with OH^{-} so you just balance the moles of these:

$$(\text{N} \cdot \text{V})_{\text{acid}} = (\text{N} \cdot \text{V})_{\text{base}}$$

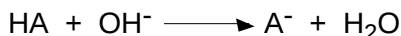
N is normality and you have to use it as the measure of concentration for titrations because it keeps track of the number of H^{+} 's or OH^{-} 's in the solution. Normality is just molarity times the number of H^{+} 's or OH^{-} 's in the molecule.
So for HCl, $\text{N} = \text{M}$ but for H_2SO_4 $\text{N} = 2\text{M}$.

The **equivalence point** is where all of the H^{+} or OH^{-} in the titrated solution has been used up. For a strong acid-strong base system the pH at the equivalence point is 7 (if you use up all the H^{+} or OH^{-} then you just have water equilibrium, right).

Near the equivalence point the concentrations of H^{+} and OH^{-} are at their smallest (remember that $[\text{H}^{+}][\text{OH}^{-}] = 10^{-14}$ so if the $\text{pH} = 7$, $[\text{H}^{+}] = [\text{OH}^{-}] = 10^{-7}$). This means that small additions of H^{+} or OH^{-} can cause large changes in $[\text{H}^{+}]$ or $[\text{OH}^{-}]$. That's why the titration curve is so steep at the equivalence point. One drop can cause a change in the pH that would correspond to a couple orders of magnitude change in $[\text{H}^{+}]$ or $[\text{OH}^{-}]$.

Weak Acid - Strong Base Titrations

These are like salts of weak acids problems. Essentially, you can think of this as adding enough strong base to react with all of the hydrogen on the weak acid, leaving behind the acid anion



Now you've reached the equivalence point (all the HA has been reacted) but the equilibrium of the system readjusts because of the A^- , forms the weak acid and OH^- .



Therefore the pH shifts above 7 at the endpoint according to the K_b of the acid anion.

$$K_b = \frac{[\text{HA}][\text{OH}^-]}{[\text{A}^-]}$$

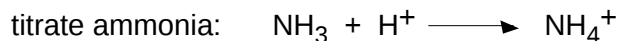
Use a table with the initial, change, and final and x stuff.

get $[\text{OH}^-]$ then pOH then pH.

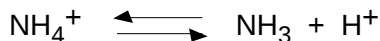
These titration curves will have a steep region near the equivalence point and also a plateau in the region where the titrated solution is behaving like a buffer.

Weak Base - Strong Acid Titrations

Here you add a strong acid (lots of H^+ 's) to react with all the base present (reach the equivalence point) and then the system readjusts because of the excess weak acid that's present:



At the equivalence point you've turned all of the ammonia into ammonium ion. Ammonium ion will then break apart and re-equilibrate according to the reaction:



$$K_a = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]}$$

Solve for $[\text{H}^+]$ and that will give you the pH at the equivalence point

Student Packet: 6.6: Titrations: AssessMe questions

- 1) 10.0 ml of a solution of HCl required 12.5 ml of 0.400 M Ba(OH)₂ for complete neutralization. How many moles of HCl were present in the sample?
- A) 5.00×10^{-3} B) 1.00×10^{-2}
C) 1.00 D) 2.00 E) I'm clueless
- 2) A 50.0 ml solution of 1.50 M NaOH is titrated with a 2.00 M HCl solution. What will the pH be after the addition of 35.0 ml of HCl?
- A) 1.23 B) 11.7 C) 12.8 D) 2.3 E) I'm clueless
- 3) What is the molarity of CH₃COOH in vinegar containing 4.0% CH₃COOH by mass and having a density of 1.02 g/ml?
- A) 0.50 M B) 0.68 M C) 0.75 M D) 1.36 M E) I'm clueless
- 4) The pH of a 0.10 M solution of a weak acid is 5.40. What is the K_a of the acid?
- A) 1.6×10^{-10} B) 3.2×10^{-10}
C) 1.6×10^{-8} D) 8.0×10^{-8} E) I'm clueless
- 5) What mass of silver chromate (Ag₂CrO₄) will dissolve in 1.0 liter of a solution that is 0.010 M in Ag⁺?
- $K_{sp} \text{ Ag}_2\text{CrO}_4 = 1.9 \times 10^{-12}$ and FM Ag₂CrO₄ = 332 g/mol.
- A) 1.9×10^{-12} g B) 1.9×10^{-8} g
C) 1.4×10^{-6} g D) 6.3×10^{-6} g E) I'm clueless
- 6) The solubility of CaF₂ in 1×10^{-3} M KF
- A) is equal to the solubility of CaF₂ in pure water.
B) is less than the solubility of CaF₂ in pure water.
C) is greater than the solubility of CaF₂ in pure water.
D) is not related to the solubility of CaF₂ in pure water.
E) I'm clueless

Student Packet: 6.6: Titrations: ShowMe questions

- 1) A solution of NaOH is titrated with H_2SO_4 . If 50 mL of the NaOH solution requires 30 mL of 0.1 M H_2SO_4 , what is the concentration of the NaOH solution?

- 2) A 2.73 g sample of an unknown weak monoprotic acid, HA, was dissolved in sufficient water to make 50 mL of solution and was titrated with a 0.250 M NaOH solution. After the addition of 12.0 mL of base, a pH of 5.26 was recorded. The equivalence point was reached after the addition of 59.2 mL of the 0.250 M NaOH.
 - a) Calculate the number of moles of acid in the original sample.

 - b) Calculate the molecular weight of the acid HA.

 - c) Calculate the number of moles of unreacted HA remaining in solution when the pH is 5.26.

 - d) Calculate $[\text{H}^+]$ at pH 5.26.

 - e) Calculate the value of K_a for HA.

- 3) What is the pH at the equivalence point of a titration of 50.00 mL of 0.100 M HOCl with 0.500 M NaOH? ($K_a = 3.7 \times 10^{-8}$)

Student Packet: 6.6: Titrations: TestMe questions

- 1) Calculate the pH of a solution prepared by mixing 1.0 mL of 0.50 M NaOH and 20.0 mL of 0.250 M $\text{C}_6\text{H}_5\text{COOH}$.
 $K_a = 6.7 \times 10^{-5}$
- A) 2.4 B) 3.2 C) 7.0 D) 13.7 E) 14.0
- 2) A sample of 80.0 mL of 0.200 M $\text{HC}_2\text{H}_3\text{O}_2$ is titrated with 0.150 M NaOH. What is the concentration of $\text{C}_2\text{H}_3\text{O}_2^-$ at the equivalence point?
- A) 0.0500 M
B) 0.0856 M
C) 0.171 M
D) 0.200 M
E) 0.236 M
- 3) Stomach acid is approximately 0.020 M HCl. What volume of stomach acid is neutralized by an antacid tablet that weighs 450 mg and contains 25.0% $\text{Mg}(\text{OH})_2$?
- A) 4.8 mL B) 9.6 mL C) 19 mL D) 23 mL E) 38 mL
- 4) In the titration of a weak acid, HA, with 0.200 M NaOH, the stoichiometric point is known to occur at a pH value of approximately 8. Which of the following indicator acids would be best to use to mark the endpoint of this titration?
- A) Indicator A, $K_a = 10^{-14}$
B) Indicator B, $K_a = 10^{-9}$
C) Indicator C, $K_a = 10^{-7}$
D) Indicator D, $K_a = 10^{-5}$
E) Indicator E, $K_a = 10^{-4}$
- 5) A student wants to standardize a solution of sodium hydroxide which was made by dissolving sodium hydroxide pellets in enough water to make 500 mL of solution. She uses 0.250 g of the monoprotic acid $\text{KHC}_8\text{H}_4\text{O}_4$ to perform the standardization. If the NaOH solution is supposed to be 0.200 M, what volume of NaOH solution should be required?
- A) 6.12 mL B) 12.2 mL C) 7.68 mL D) 3.06 mL E) 10.0 mL