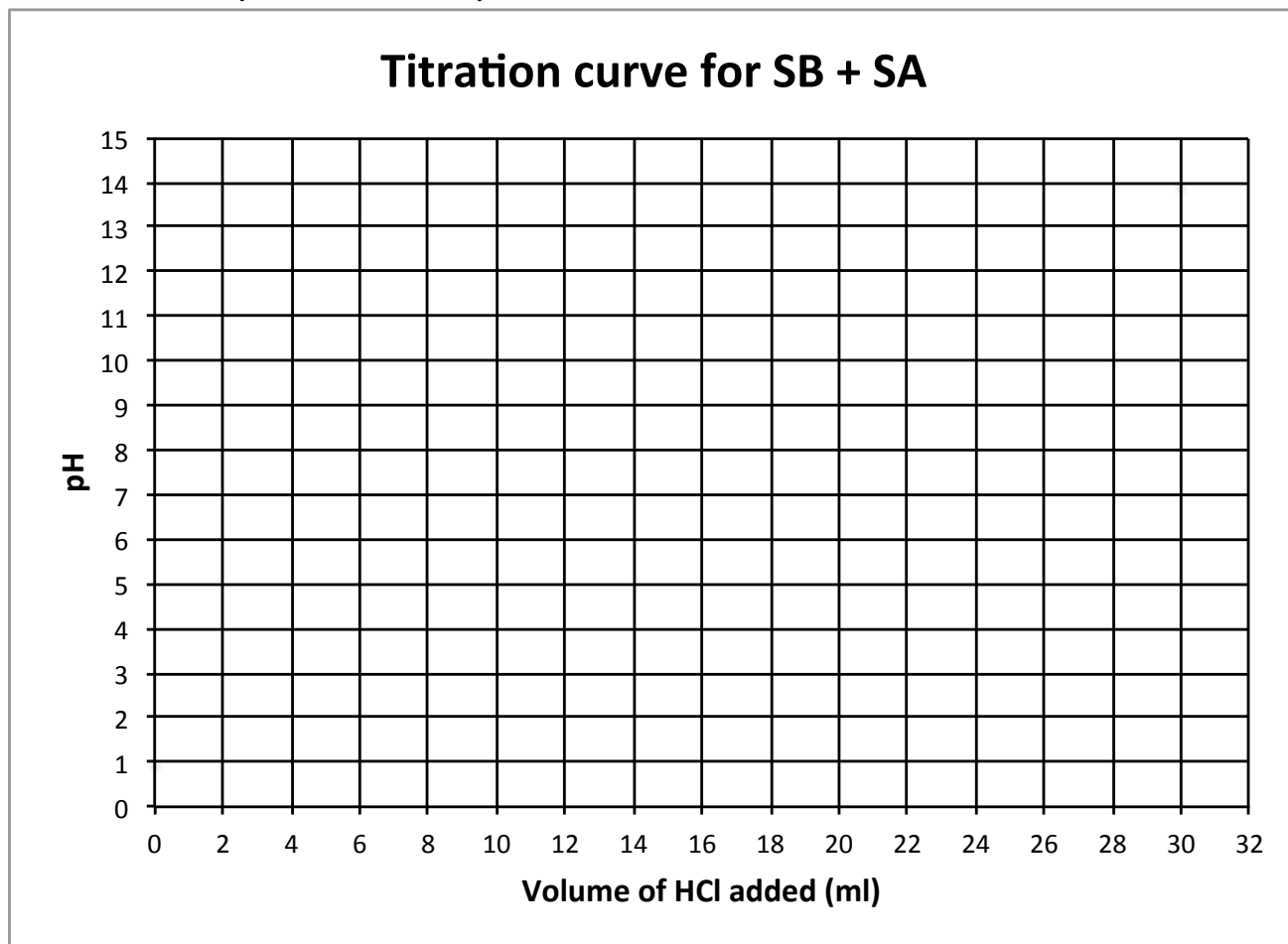


Consider the changes (pH, $[H^+]$, $[OH^-]$) of titrating 30.0 ml of 0.040 M potassium hydroxide with 0.050 M hydrochloric acid.

Then construct a titration curve on the graph on page 2.

1. Is hydrochloric acid a strong or weak acid? Write a reaction that represents its ionization in water.
2. Is potassium hydroxide a strong or weak base? Write a reaction that represents its ionization in water.
3. Determine the initial pH of the 30.0 ml of 0.040 M potassium hydroxide solution.
4. Determine the initial pH of the 0.050 M hydrochloric acid solution.
5. What volume of hydrochloric acid would you need to add to the 30.0 ml of 0.040 M potassium hydroxide if you wanted to completely neutralize this acid?
6. What would be the pH of this resulting solution from the previous question?
7. Calculate the pH of the resulting solution when 2.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.
8. Calculate the pH of the resulting solution when 12.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.
9. Calculate the pH of the resulting solution when 20.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.
10. Calculate the pH of the resulting solution when 24.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.
11. Calculate the pH of the resulting solution when 30.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 30 ml of 0.00 M sodium hydroxide with 0.050 M hydrochloric acid.



As you construct this graph, you will get a much more accurate sense of the titration curve if you make two more calculations and place two more points on the graph.

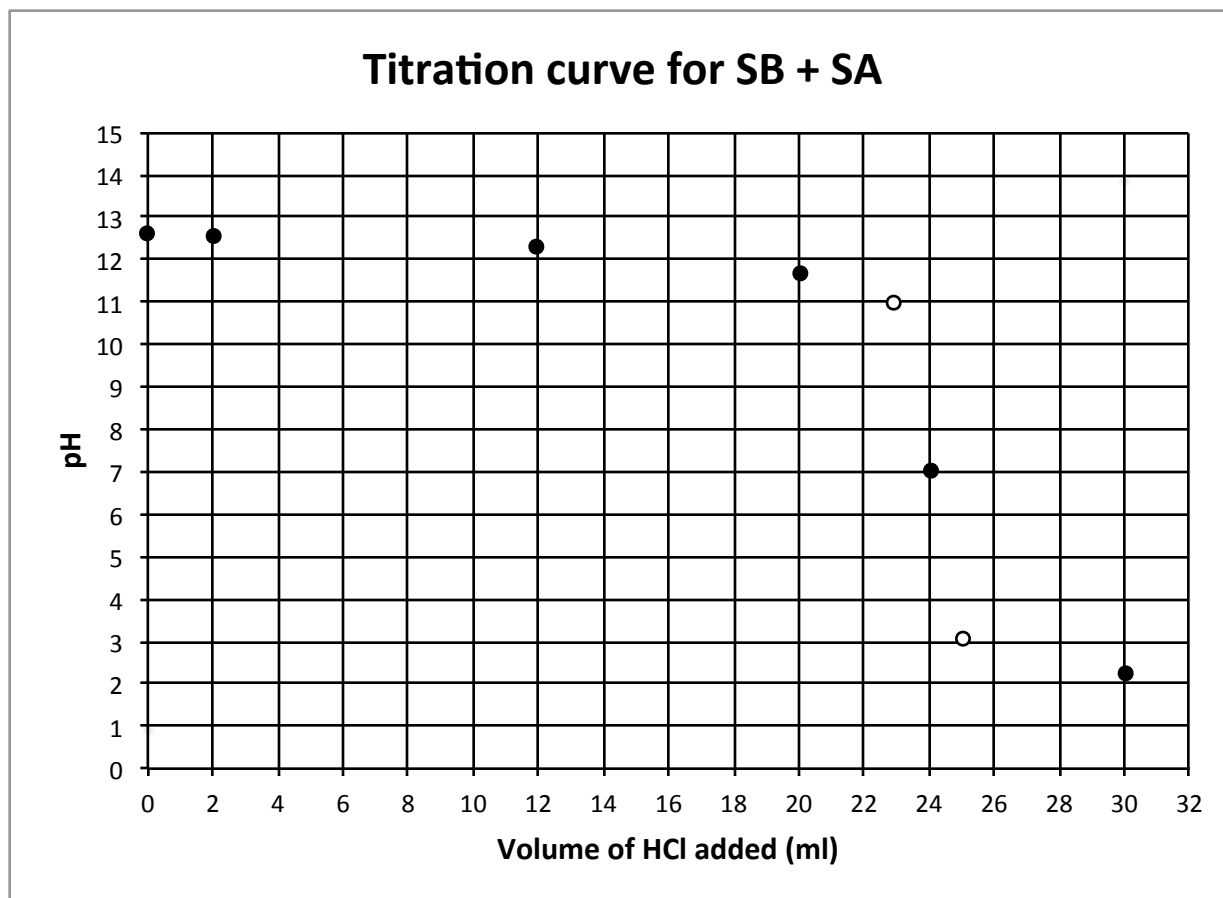
12. Calculate the pH of the resulting solution when 23.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.

13. Calculate the pH of the resulting solution when 25.0 ml of the 0.050 M hydrochloric acid is added to 30.0 ml of 0.040 M potassium hydroxide.

Consider the consequences (pH, $[H^+]$, $[OH^-]$) of adding 0.05 M hydrochloric acid to 30.0 ml of 0.040 M potassium hydroxide.

1. HCl is one of the “seven” **strong acids** that you need to memorize. $HCl \rightarrow H^+ + Cl^-$
2. Potassium hydroxide a **strong base**, as are all soluble metal hydroxides. $KOH \rightarrow K^+ + OH^-$
3. Don’t be distracted by the volume in this question. The pH of *any* volume of 0.04 M KOH will be the same. Thus, for 0.040 M KOH, the $[OH^-] = 0.04$ M, calculate $-\log(0.04)$ for $pOH = 1.40$ then “14-it” for **pH = 12.60**
4. As in the last question, don’t be distracted by the volume in this question. The pH of *any* volume of 0.05 M HCl will be the same. Thus for 0.05 M HCl, the $[H^+] = 0.05$ M, calculate $-\log(0.05)$ for **pH = 1.3**
5. At the equivalence point, moles of acid equals moles of base $M_a V_a = M_b V_b$
 - so $0.05\text{ M} \times V_a = 0.040\text{ M} \times 30.0\text{ ml}$
 - thus the volume of acid needed to reach the equivalence point is **24.0 ml of HCl**
6. The pH at the equivalence point of any SB titrated with strong acid always equals 7 because the two ions left in solution are pathetic conjugate acid, K^+ and the pathetic conjugate base, Cl^- .
7. The incoming acid neutralizes an equivalent portion of the base, and the pH is caused by remaining un-neutralized base.
 - Calculate millimoles of base: $0.040\text{ M} \times 30\text{ ml} = 1.2\text{ mmoles of } OH^-$
 - Calculate millimoles of acid: $0.050\text{ M} \times 5\text{ ml} = 0.1\text{ mmoles of } H^+$
 - Thus there is 1.1 millimole of OH^- remaining, divide this by total volume, 32 ml to get 0.0344 M OH^-
 - $-\log(0.0344)$ to get $pOH = 1.46$ and subtract from 14 to get **pH = 12.54**
8. As in question #7, the incoming acid neutralizes an equivalent portion of the base, and the pH is caused by remaining un-neutralized base.
 - Calculate millimoles of base: $0.040\text{ M} \times 30\text{ ml} = 1.2\text{ mmoles of } OH^-$
 - Calculate millimoles of acid: $0.05\text{ M} \times 12\text{ ml} = 0.6\text{ mmoles of } H^+$
 - Thus there is 0.6 millimole of OH^- remaining, divide this by total volume, 42 ml to get 0.0143 M of OH^-
 - $-\log(0.0143)$ to get $pOH = 1.85$ and subtract from 14 to get **pH = 12.15**
9. As in question #7 and #8, the incoming acid neutralizes an equivalent portion of the base, and the pH is caused by remaining un-neutralized base.
 - Calculate millimoles of base: $0.040\text{ M} \times 30\text{ ml} = 1.2\text{ mmoles of } OH^-$
 - Calculate millimoles of acid: $0.05\text{ M} \times 20\text{ ml} = 1\text{ mmoles of } H^+$
 - Thus there is 0.2 millimole of OH^- remaining, divide this by total volume, 50 ml to get 0.004 M of OH^-
 - $-\log(0.004)$ to get $pOH = 2.40$ and subtract from 14 to get **pH = 11.60**
10. This question was asked in #6, here’s a bit more explanation. At this point the amount of incoming acid equals the base, which will be completely neutralized. This would imply that there would be 0 mmole of H^+ and 0 millimole of OH^- , but of course in any aqueous solution, they are never 0.
 - The H^+ and OH^- would have turned into water with only K^+ and Cl^- ions (pathetic conjugate acid and pathetic conjugate base)
 - and in pure water (or water that only contains pathetic conjugate ions), the $[H^+]$ always equals $[OH^-]$ which is $1 \times 10^{-7}\text{ M}$
 - Thus $-\log(1 \times 10^{-7})$ to get the **pH = 7.0**
11. At this point there is 6 ml more of incoming acid than is required to neutralize the base, which will be completely neutralized by 24 ml of the incoming acid and the pH is caused by the 6 ml of excess acid present
 - Calculate millimoles of excess acid: $0.050\text{ M} \times 6\text{ ml} = 0.3\text{ mmoles of } H^+$
 - Divide the mmoles of excess acid by total volume, 60 ml to get 0.005 M of H^+
 - $-\log(0.005)$ to get **pH = 2.30**

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 30 ml of 0.00 M sodium hydroxide with 0.050 M hydrochloric acid.



- As in question #7, 8, and 9, the incoming acid neutralizes an equivalent portion of the base, and the pH is caused by remaining un-neutralized base.
 - Calculate millimoles of base: $0.040 \text{ M} \times 30 \text{ ml} = 1.2 \text{ mmoles of OH}^-$
 - Calculate millimoles of acid: $0.05 \text{ M} \times 23 \text{ ml} = 1.15 \text{ mmoles of H}^+$
 - Thus there is 0.05 millimole of OH^- remaining, divide this by total volume, 43 ml to get 0.0012 M of OH^-
 - $-\log(0.004)$ to get $\text{pOH} = 2.93$ and subtract from 14 to get **pH = 11.07**
- At this point, as in question # 11, there is 1 ml more of incoming acid than required to neutralize the base, which will be completely neutralized by the 24 ml of incoming acid and the pH is caused by the 1 ml of excess acid present.
 - Calculate millimoles of excess acid: $0.050 \text{ M} \times 1 \text{ ml} = 0.05 \text{ mmoles of H}^+$
 - Divide the mmoles of excess acid by total volume, 55 ml to get 0.00091 M of H^+
 - $-\log(0.005)$ to get **pH = 3.04**