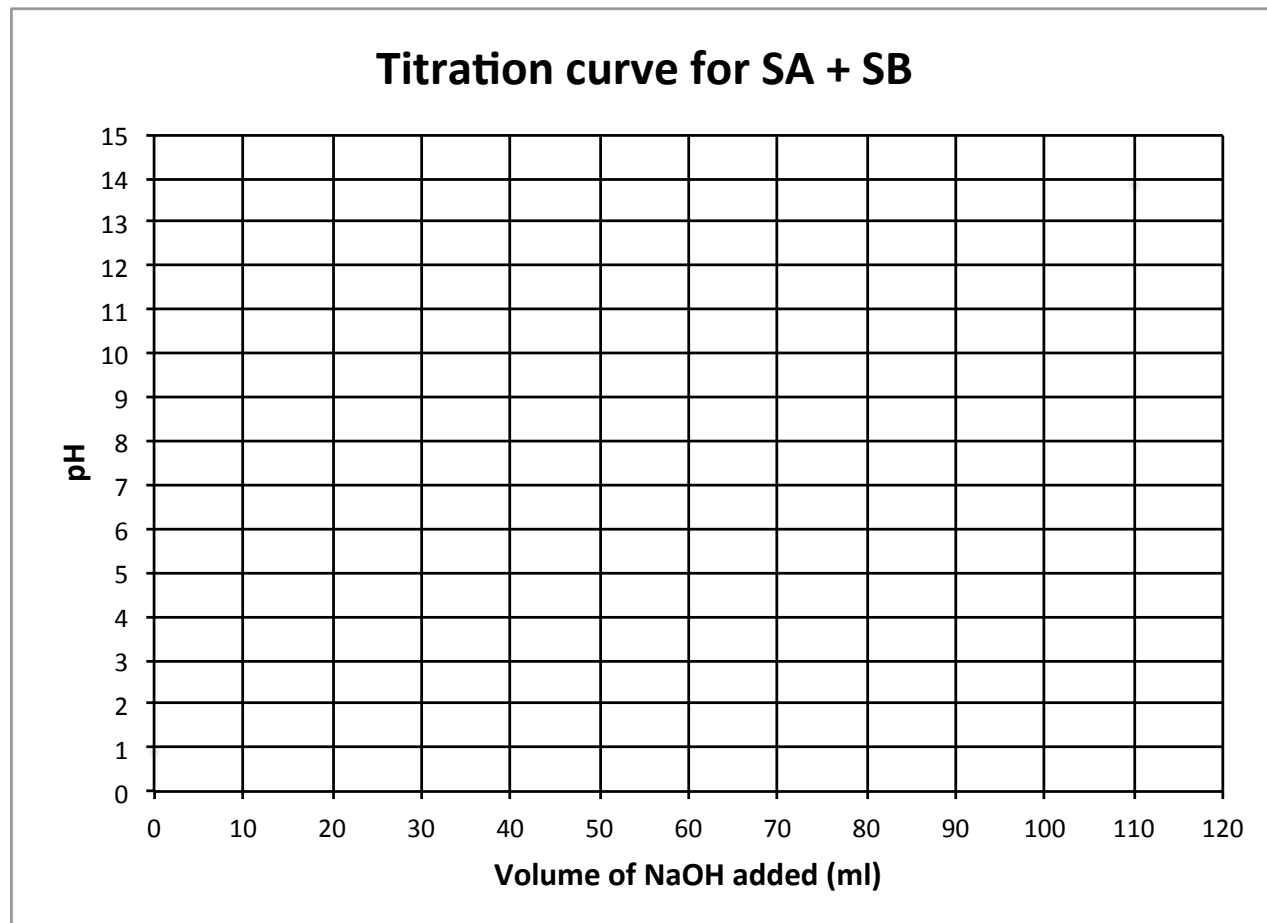


Consider the changes (pH, $[H^+]$, $[OH^-]$) when titrating 50.0 ml of 0.040 M nitric acid with 0.025 M sodium hydroxide. Then construct a titration curve on the graph on page 2.

1. Is nitric acid a strong or weak acid? Write a reaction that represents its ionization in water.
2. Is sodium hydroxide a strong or weak base? Write a reaction that represents its ionization in water.
3. Determine the initial pH of the 50.0 ml of 0.040 M nitric acid solution.
4. Determine the initial pH of the 0.025 M sodium hydroxide solution.
5. What volume of sodium hydroxide would you need to add to the 50.0 ml of 0.040 M nitric acid 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 5.0 ml of the 0.025 M sodium hydroxide is added to 50.0 ml of 0.040 M nitric acid.
8. Calculate the pH of the resulting solution when 40.0 ml of the 0.025 M sodium hydroxide is added to 50.0 ml of 0.040 M nitric acid.
9. Calculate the pH of the resulting solution when 75.0 ml of the 0.025 M sodium hydroxide is added to 50.0 ml of 0.040 M nitric acid.
10. Calculate the pH of the resulting solution when 90.0 ml of the 0.025 M sodium hydroxide is added to 50.0 ml of 0.040 M nitric acid.
11. Calculate the pH of the resulting solution when 110.0 ml of the 0.025 M sodium hydroxide is added to 50.0 ml of 0.040 M nitric acid.

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 50 ml of 0.40 M nitric acid with 0.025 M NaOH.



- Nitric is one of the seven **strong acids**. $\text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^-$ or $\text{HNO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{NO}_3^-$
- NaOH is a **strong base**. $\text{NaOH} \rightarrow \text{Na}^+ + \text{OH}^-$
- Since we have a SA which dissociates completely, the $[\text{H}^+]$ is the $[\text{HNO}_3]$, thus $\text{pH} = -\log[\text{H}^+]$
 - $-\log(0.040)$ thus $\text{pH} = 1.40$ (remember for pH 1.40 has only 2 sig figs, the 1.40, the 1 in front is magnitude, not actually part of the 2 sig figs in 0.040 M).
- Since we have a SB which dissociates completely, the $[\text{OH}^-]$ is the $[\text{NaOH}]$, thus $\text{pOH} = -\log[\text{OH}^-]$
 - $-\log(0.025)$ $\text{pOH} = 1.60$, the “14-it” to get $\text{pH} = 12.40$
- This is called the equivalence point. Since we have a “monobasic” SB reacting with a monoprotic SA the acid and base react in a 1:1 ratio. We need to calculate the moment when moles $\text{H}^+ = \text{moles OH}^-$.
 - We can use the $M_a V_a = M_b V_b$. You can do your calculations in ml, which would mean you would be calculating millimoles instead of moles, but since the volume is in this equation twice, it will simply report the volume in ml (You can use liters for your volume of acid, it will simply mean your volume of base would come out in liters, 0.080 L).
 - Thus $(0.04)(50) = (0.025)(V_b)$, and $V_b = 80 \text{ ml}$
- Always at the equivalence point of a SA with a SB, the **pH will be equal to 7**. This is because the only species in solution are the pathetic conjugates of the SA and SB.
- Tackling this question is really just a limiting reactant problem. Since you know that it takes 80 ml of base to neutralize, any volume of base up until 80 ml will mean that there is excess acid, and the pH will be a function of the $[\text{H}^+]$.
 - Calculate mmol of acid using MV, $(0.04\text{M})(50\text{ml}) = 2 \text{ mmol}$ and mmol base is $(0.025\text{M})(5\text{ml}) = 0.125 \text{ mmol OH}^-$
 - A simple subtraction leaves $1.875 \text{ mmol H}^+ \cdot \frac{1.875 \text{ mmol H}^+}{55 \text{ ml Total Vol}} = 0.034 \text{ M H}^+$. Calculate pH, $-\log(0.034)$ thus **pH = 1.47**
- This problem is the same process as question #7. Calculate the incoming mmol base; $(0.025\text{M})(40\text{ml}) = 1 \text{ mmol OH}^-$
 - A simple subtraction from the previously calculated 2 mmol acid leaves 1 mmol H^+ . Use this 1 mmol of H^+ to calculate the concentration of that excess acid that has not yet been neutralized $\frac{1.0 \text{ mmol H}^+}{90 \text{ ml Total Vol}} = 0.011 \text{ M H}^+$.
 - Then finish by calculating pH. $-\log(0.011)$ thus **pH = 1.95**
- Same process as problem #7 and #8. Calculate the incoming mmol base is $(0.025\text{M})(75\text{ml}) = 1.875 \text{ mmol OH}^-$
 - Subtraction from the previously calculated 2 mmol acid leaves 0.125 mmol H^+ Use this 0.125 mmol of H^+ to calculate the concentration of that excess acid that has not yet been neutralized $\frac{0.125 \text{ mmol H}^+}{125 \text{ ml Total Vol}} = 0.001 \text{ M H}^+$.
 - Then finish by calculating pH; $-\log(0.001)$ thus **pH = 3.00**
- Now that we have gone beyond the equivalence point, there is 10 ml excess base, since we calculated in #5 that 80 ml of base is required for the neutralization, thus the pH will be completely dependent on the moles of excess base in the total volume of the solution.
 - mmol excess base is $(0.025\text{M})(10\text{ml}) = 0.25 \text{ mmol OH}^-$ Use this mmol of OH^- to calculate the concentration of that excess base that is now in the solution. $\frac{0.25 \text{ mmol OH}^-}{140 \text{ ml Total Vol}} = 0.00178 \text{ M OH}^-$.
 - Calculate pOH, $-\log(0.00178)$ thus $\text{pOH} = 2.75$, and you can “14-it” to get the **pH = 11.25**
- This is the same process as question #10, we have simply gone further beyond the equivalence point, a total of 30 ml beyond
 - mmol excess base is $(0.025\text{M})(30\text{ml}) = 0.75 \text{ mmol OH}^-$ Use this mmol of OH^- to calculate the concentration of that excess base that is now in the solution. $\frac{0.75 \text{ mmol OH}^-}{160 \text{ ml Total Vol}} = 0.00469 \text{ M OH}^-$.
 - Calculate pOH, $-\log(0.00469) = 2.33$, thus **pH = 11.67**

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 50 ml of 0.40 M nitric acid with 0.025 M NaOH.

