P G4 B (pg 1 of 4) **Acid – Base Titration Calculations (WB + SA)** Name_____Per____

Consider a 0.060 M solution of ammonia, NH_3 to which you plan to add 0.040 M hydroiodic acid, HI.
The K_b of ammonia is 1.8×10^{-5}

1. Calculate the pH of 20.0 ml of a 0.060 M solution of ammonia, NH_3.

2. What volume of 0.040 M HI would be needed to neutralize this 20.0 ml of a 0.060 M solution of ammonia, NH_3?

3. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 5.0 ml of 0.040 M hydroiodic acid, HI?

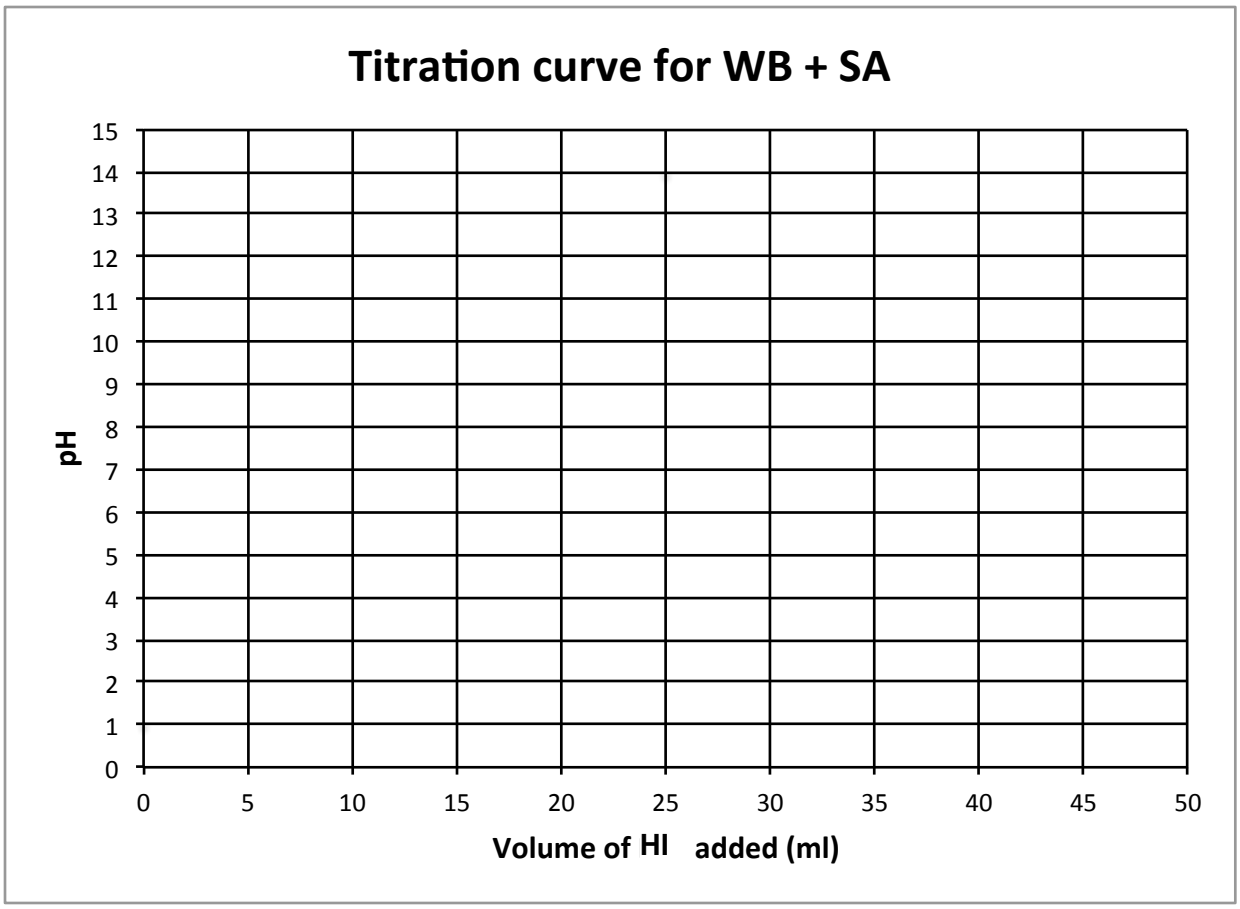
4. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 15.0 ml of 0.040 M hydroiodic acid, HI?

5. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 20.0 ml of 0.040 M hydroiodic acid, HI?

6. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 30.0 ml of 0.040 M hydroiodic acid, HI?

7. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 50.0 ml of 0.040 M hydroiodic acid, HI?

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 20 ml of 0.060 M ammonia with 0.040 M hydroiodic acid.



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

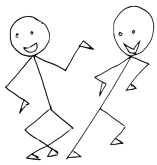
8. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 25.0 ml of 0.040 M hydroiodic acid, HI?

9. Calculate the pH of 20.0 ml of 0.060 M ammonia, NH_3 after it has been titrated with a total of 35.0 ml of 0.040 M hydroiodic acid, HI?

Be sure that if you need help, you refer to NS G.1 – Acid-Base Titrations.

First you need to recognize that this is a WB being titrated with a SA.

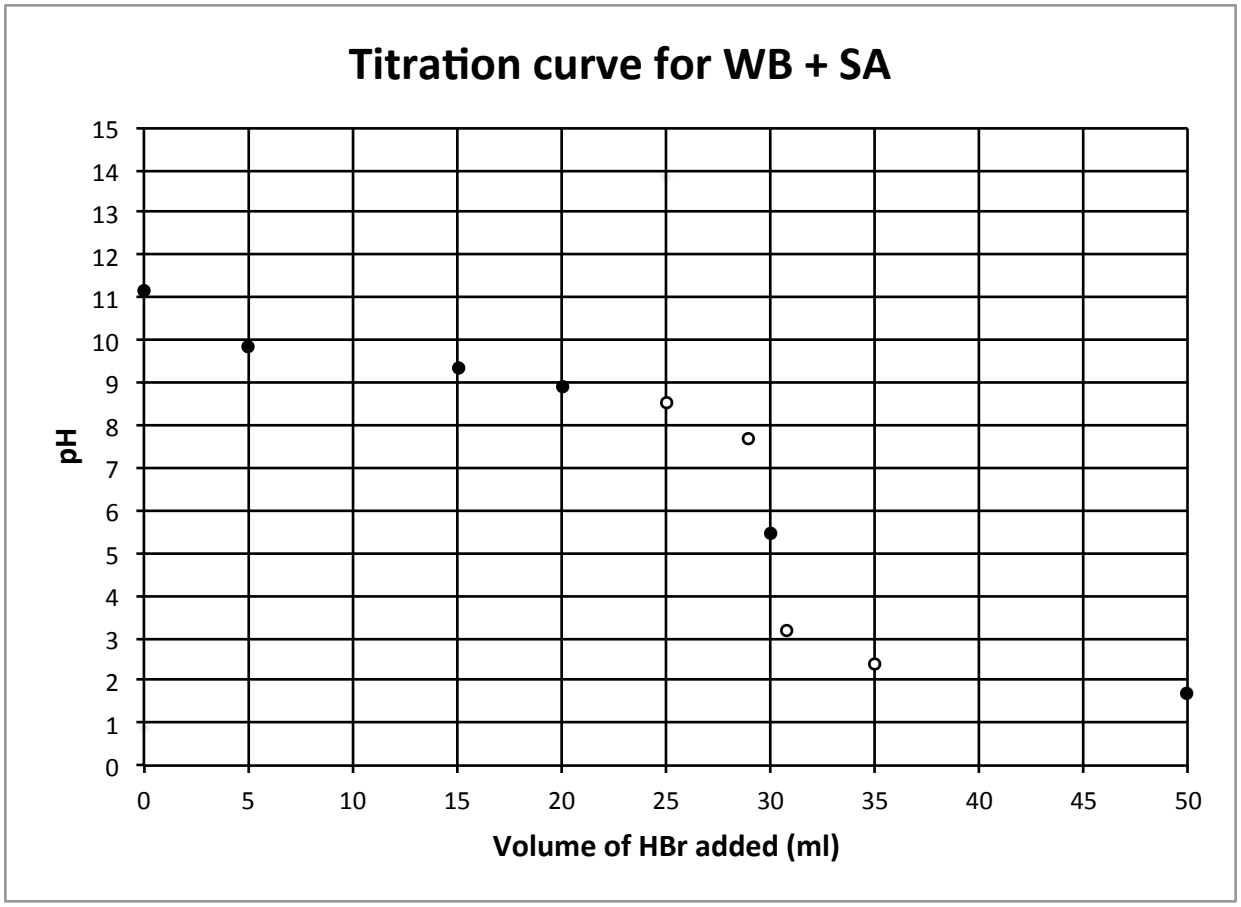
1. Solve for pH *before* any titration with a simple x² set-up. $K_b = \frac{\left[OH^-\right]\left[HB^{+\ +incomingSA}_{\ -incomingSB}\right]}{\left[B^{\ -incomingSA}_{\ +incomingSB}\right]}$
   - Remember that the WB really is 0.060 − x, however the − x is too small to worry about. The AP exam will not require that you solve a problem in which you would need to use the quadratic equation. Thus $1.8\times10^{-5} = \frac{x^2}{[0.060]}$ x = [OH⁻] = 1.04 × 10⁻³ M
   - then −log[OH⁻] to determine that pOH = 2.98 and then "14-it" for **pH = 11.02**
2. At the equivalence point, the moles of base = moles of acid, so use: $M_bV_b = M_aV_a$
   - so 0.06 M × 20 ml = 0.04 M × V thus the vol needed to reach the equivalence point is **30. ml of HI**
3. At this point you are partway through the titration, but not halfway, thus you are in the midst of the titration.
   - Calculate the millimoles of base, ammonia (0.06 M × 20 ml = 1.2 mmol)
   - and millimoles of incoming strong acid (0.04 M × 5 ml = 0.2 mmol)
   - $K_b = \frac{\left[OH^-\right]\left[HB^{+\ +incomingSA}_{\ -incomingSB}\right]}{\left[B^{\ -incomingSA}_{\ +incomingSB}\right]}$ $1.8\times10^{-5} = \frac{\left[OH^-\right]\left(\sim 0 + 0.2 mmol H^+\right)}{\left(1.2 mmol - 0.2 mmol H^+\right)}$
   - [OH⁻] = 9.0 × 10⁻⁵, then log for pOH = 4.05 thus "14 it" and **pH = 9.95**
4. Yippee!!! This is the halfway point of the titration, and at this point, the [HB] = [HB⁺] thus $K_b = [OH^-]\left(\frac{HB^+ =}{B =}\right)$
   - so the K_b = [OH⁻] = 1.8 × 10⁻⁵, thus pH = 4.74 **pH = 9.26**



5. Just like question # 3, you are still partway through the titration and in the midst of the titration. Have no fear....you can do this.
   - As calculated before the millimoles of base, ammonia (0.06 M × 20 ml = 1.2 mmol)
   - and millimoles of incoming strong acid (0.04 M × 20 ml = 0.8 mmol)
   - $K_b = \frac{\left[OH^-\right]\left[HB^{+\ +incomingSA}_{\ -incomingSB}\right]}{\left[B^{\ -incomingSA}_{\ +incomingSB}\right]}$ $1.8\times10^{-5} = \frac{\left[OH^-\right]\left(\sim 0 + 0.8 mmol H^+\right)}{\left(1.2 mmol - 0.8 mmol H^+\right)}$ [OH⁻] = 9.0 × 10⁻⁶, thus pOH = 5.05 and "14-it" so **pH = 8.95**
6. At this point you are at the equivalence point and all of the WB has been neutralized, and the equilibrium has been driven all the way to the right leaving only the conjugate weak acid, NH_4^+ aka WB^+ in the solution.
   - You have already calculated the millimol of WB, (0.06 M × 20 ml = 1.2 mmol), and since the equilibrium has been driven completely to the right, that number now is the amount of conjugate HB^+, NH_4^+ in solution.
   - Calculate the concentration of this conjugate weak acid 1.2 mmole by dividing by total volume (50. ml) $\frac{1.2 mmol}{50 ml} = 0.024 M$
   - You can then use the [HB⁺], an acid, in a simple "x²" WA calculation to determine the [H⁺] $K_a = \frac{[B][H^+]}{[HB^+]}$

     We'll need the K_a for the calculation at the equivalence point. Remember $K_a = \frac{K_w}{K_b}$ $K_a = \frac{1\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$
   - thus $5.6\times10^{-10} = \frac{x^2}{[0.024]}$ and x = [H⁺] = 3.7 × 10⁻⁶ M. Calculate pH with −log(3.7 × 10⁻⁶), thus **pH = 5.44**
7. At this point in the titration the amount of acid added exceeds the amount of base. From the calculation in question #2, we know that there are 20 ml extra acid beyond the equivalence point.
   - (0.04 M × 20 ml) results in an extra 0.80 mmol of excess strong acid.
   - Divide this extra 0.08 mmol of acid by the total volume of 70. ml; $\frac{0.08 mmol}{70 ml} = 0.0114 M$ of strong acid which contributes far more H⁺ than any H⁺ contributed by the conjugate weak acid (so you can ignore the weak acid's contribution).
   - Calculate −log[0.014] thus **pH = 1.94**

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 20 ml of 0.060 M ammonia with 0.04 M hydroiodic acid.



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

8. Just like question # 3 and #5, you are still partway through the titration and in the midst of the titration
   - As calculated before the millimoles of base, ammonia (0.06 M × 20ml = 1.2 mmole)
   - and millimoles of incoming strong acid (0.04 M × 25.0 ml = 1.0 mmole)
   - $K_b = \dfrac{\left[OH^-\right]\left[HB^{+}\ {}^{+incomingSA}_{-incomingSB}\right]}{\left[B\ {}^{-incomingSA}_{+incomingSB}\right]}$ $1.8\times10^{-5} = \dfrac{\left[OH^-\right]\left[\sim 0+1mmolH^+\right]}{\left[1.2mmol-1mmolH^+\right]}$ thus $[OH^-] = 3.6 \times 10^{-6}$, pOH = 5.44 thus **pH = 8.56**

9. Just like the calculation for #7, at this point in the titration the amount of acid added exceeds the amount of base. From the calculation in question #2, we know that there are 5 ml extra strong acid beyond the equivalence point.
   - (0.04 M × 5 ml) results in an extra 0.2 mmole of strong acid.
   - Divide this extra 0.2 mmole of strong acid by the total volume of 55 ml; $\dfrac{0.2mmol}{55ml} = 0.0036M$ of strong acid which contributes far more H^+ than any H^+ contributed by the conjugate weak acid (so you can ignore the weak acid's contribution).
   - Calculate −log(0.0036) thus **pH = 2.44**

 Acid – Base Titration Calculations (WB w/ SA) Name_____Per____

Consider a 0.080 M solution of trimethylamine, $(CH_3)_3N$ to which you plan to add 0.066 M hydrobromic acid, HBr.
The K_b of trimethylamine is 6.4×10^{-5}

1. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$.

2. What volume of 0.066 M HBr would be needed to neutralize this 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$? (This should be the first step you would take to solve any of the following problems, even if you weren't asked to do it.)

3. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with 5.0 ml of 0.066 M hydrobromic acid, HBr?

4. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 20.6 ml of 0.066 M hydrobromic acid, HBr?

5. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 30.0 ml of 0.066 M hydrobromic acid, HBr?

6. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 41.2 ml of 0.066 M hydrobromic acid, HBr?

7. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 45 ml of 0.066 M HBr?

Be sure that if you need help, you refer to NS G.1 – Acid-Base Titrations.

First you need to recognize that this is a WB being titrated with a SA.

1. Solve for pH *before* any titration with a simple x² set-up. $K_b = \frac{[OH^-][WA]}{[WB]}$
   - Remember that the WB really is 0.080 − x, however the − x is too small to worry about. The AP exam will not require that you solve a problem in which you would need to use the quadratic equation. Thus $6.4\times10^{-5} = \frac{x^2}{[0.080]}$ x = [OH⁻] = 2.3 × 10⁻³ M
   - then −log[OH⁻] to determine that pOH = 2.65 and then "14-it" for **pH = 11.35**

2. At the equivalence point, the moles of base = moles of acid, so use: $M_bV_b = M_aV_a$
   - so 0.08 M × 34 ml = 0.066 M × V thus the vol needed to reach the equivalence point is **41 (41.2) ml of HBr**

3. At this point you are partway through the titration, but not halfway, thus you are in the "land of buffers."
   - Calculate the millimol of base, trimethylamine (0.08 M × 34 ml = 2.72 mmol)
   - and millimol of incoming strong acid (0.066M × 5ml = 0.33 mmol)
   - $K_b = \frac{[OH^-]\left[HB^{+}\ {}^{+incomingSA}_{-incomingSB}\right]}{\left[B\ {}^{-incomingSA}_{+incomingSB}\right]}$ $6.4\times10^{-5} = \frac{[OH^-][\sim0+0.33mmolH^+]}{[2.72mmol-0.33mmolH^+]}$ [OH⁻] = 4.64 × 10⁻⁴, pOH = 3.33 then "14-it" for **pH = 10.67**

4. Yippee!!! This is the halfway point of the titration, and at this point, the [B] = [HB⁺] thus $K_b = [OH^-]\left(\frac{HB^+ =}{B =}\right)$

   - so the K_b = [OH⁻] = 6.4 × 10⁻⁵, thus pOH = 4.19 and **pH = 9.81**

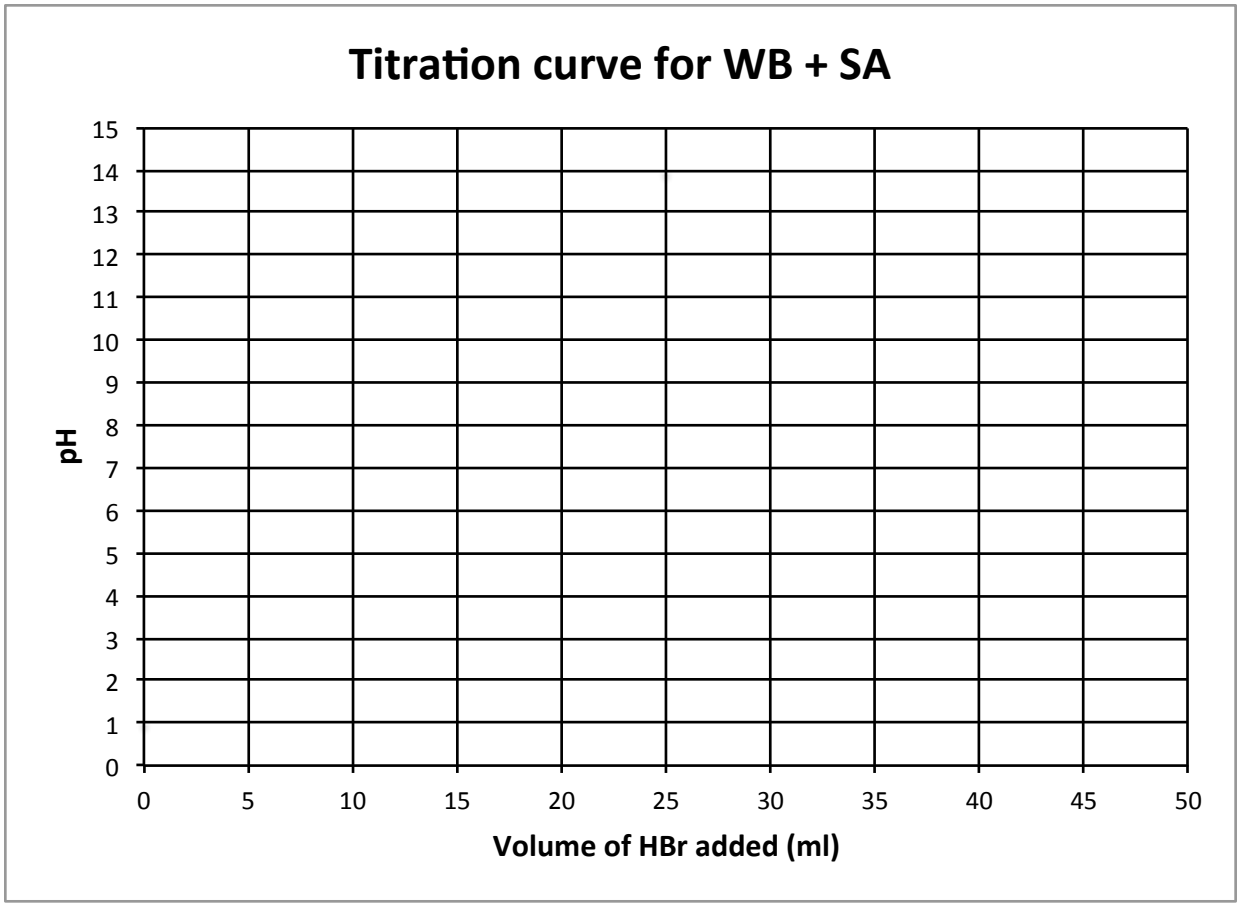
5. Just like question # 3, you are still partway through the titration and in the "land of buffers."
   - As calculated before the millimoles of base, trimethylamine (0.08 M × 34 ml = 2.72 mmol)
   - and millimoles of incoming acid (0.066M × 30ml = 1.98 mmol)
   - $K_b = \frac{[OH^-]\left[HB^{+}\ {}^{+incomingSA}_{-incomingSB}\right]}{\left[B\ {}^{-incomingSA}_{+incomingSB}\right]}$ $6.4\times10^{-5} = \frac{[OH^-][\sim0+1.98mmolH^+]}{[2.72mmol-1.98mmolH^+]}$ [OH⁻] = 2.39 × 10⁻⁵, pOH = 4.62 and pH = 9.38

6. At this point you are at the equivalence point and all of the WB has been neutralized, and the equilibrium has been driven all the way to the right leaving only the conjugate weak acid, $(CH_3)_3NH^+$ aka HB^+ in the solution.
   - You have already calculated the millimoles of B, (0.08M × 34ml = 2.72 mmol), and since the equilibrium has been driven completely to the right, that number now is the amount of conjugate weak acid, HB^+ ($(CH_3)_3NH^+$) in solution.
   - Calculate the concentration of this conjugate weak acid 2.72 mmol by dividing by total volume (75.2 ml) $\frac{2.72mmol}{75.2ml} = 0.036M$
   - You can then use the [HB⁺] in a simple "x²" WA calculation to determine the [H⁺] $K_a = \frac{[H^+][B]}{[HB^+]}$

     We'll need the K_a for the calculation at the equivalence point. Remember $K_a = \frac{K_w}{K_b}$ $K_a = \frac{1\times10^{-14}}{6.4\times10^{-5}} = 1.56\times10^{-10}$
   - thus $1.56\times10^{-10} = \frac{x^2}{[0.036]}$ and x = [H⁺] = 2.4 × 10⁻⁶ M. Calculate pH with −log(2.4 × 10⁻⁶) Thus **pH = 5.62**

7. At this point in the titration the amount of acid added exceeds the amount of base. From the calculation in question #2, we know that there are 3.8 ml extra acid beyond the equivalence point.
   - (0.066 M × 3.8 ml) results in an extra 0.25 mmol of strong acid.
   - Divide this extra 0.25 mmol of acid by the total volume of 79 ml; $\frac{0.25mmol}{79ml} = 0.0032M$ of strong acid which contributes far more H⁺ than any H⁺ contributed by the conjugate weak acid (so you can ignore the weak acid's contribution).
   - Calculate −log(0.0032) thus **pH = 2.50**

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 34 ml of 0.080 M trimethylamine with 0.066 M hydrobromic acid.



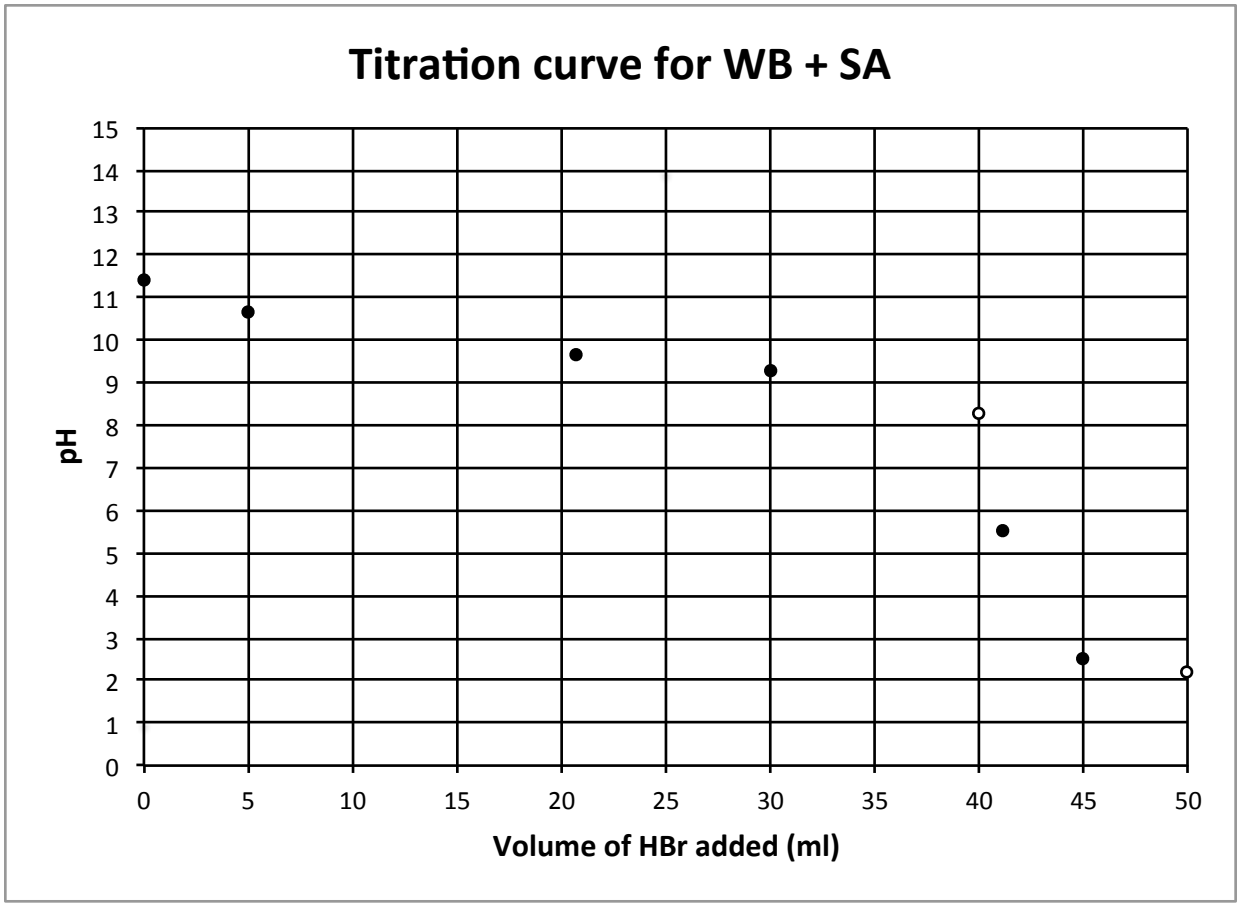
As you construct this graph, you will get a much more accurate sense of the curve if you make two more calculations and place two more points on the graph.

8. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 40.0 ml of 0.066 M hydrobromic acid, HBr?

9. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 42.0 ml of 0.066 M hydrobromic acid, HBr?

10. Calculate the pH of 34.0 ml of a 0.080 M solution of trimethylamine, $(CH_3)_3N$ after it has been titrated with a total of 50.0 ml of 0.066 M HBr?

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 34 ml of 0.080 M trimethylamine with 0.066 M hydrobromic acid.



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

8. Just like question # 3 and #5, you are still partway through the titration and in the "land of buffers." Have no fear....you can do this.
   - As calculated before the millimoles of base, trimethylamine (0.08M × 34ml = 2.72 mmole)
   - and millimoles of incoming acid (0.066M × 40.0 ml = 2.64 mmole)
   - $K_b = \dfrac{\left[OH^-\right]\left[HB^{+\ +incomingSA}_{\ -incomingSB}\right]}{\left[B^{\ -incomingSA}_{\ +incomingSB}\right]}$ $6.4\times10^{-5} = \dfrac{\left[OH^-\right]\left[\sim 0 + 2.64 mmol H^+\right]}{\left[2.72 mmol - 2.64 mmol H^+\right]}$ $[OH^-] = 1.94 \times 10^{-6}$, pOH = 5.71 and **pH = 8.29**

9. Just like the calculation for #7, at this point in the titration the amount of acid added exceeds the amount of base. From the calculation in question #2, we know that there are 8.8 ml extra acid beyond the equivalence point.
   - (0.066M × 8.8ml) results in an extra 0.58 mmole of strong acid.
   - Divide this extra 0.55 mmole of acid by the total volume of 84 ml; $\dfrac{0.58 mmol}{84 ml} = 0.0069 M$ of strong acid which contributes far more H^+ than any H^+ contributed by the conjugate weak acid (so you can ignore the weak acid's contribution).
   - Calculate −log(0.0069) thus **pH = 2.16**