Acid – Base Titration

Prelab: In grade 11 we used the titration equation: M_A x #H x V_A = M_B x #OH x V_B. This equation allows us to determine what combination of acid and base results in neutralization. Essentially, it employs the idea that neutralization occurs when the number of moles of H^+ is equal to the number of moles of OH^-.
I.e. M_A x #H gives the concentration of H^+ in mol/L. Then we multiply by V_A (in L) to get moles of H^+.
E.g. What volume of 1 M $Al(OH)_3$ would be required to neutralize 50 mL of 2 M H_2SO_4?

M_A = 2 mol/L, #H = 2, V_A = 0.050 L, M_B = 1 mol/L, #OH = 3, V_B = ?
(2 mol/L)(2)(0.050 L) = (1 mol/L)(3)(V_B), thus V_B = 0.067 L or 67 mL

Question: What volume of 6 M H_2SO_4 would be required to neutralize 500 mL of 0.5 M NaOH?

Procedure:

1. Your instructor will demonstrate the procedure for using and rinsing pipettes and burettes.
2. Gather together a burette, a squeeze bottle containing distilled water, a 100 mL beaker, a 50 mL beaker. Rinse all equipment well with tap water. Dry the beakers with paper towel.
3. Place 0.2 M HCl in the 50 mL beaker. Rinse the burette with acid (don’t forget to rinse the tip). Fill the burette to the 0 mL mark with HCl. (Ensure that there are no air bubbles in the tip of the burette).
4. Using the pipette at the front of the room, transfer 25 mL of 0.2 M NaOH into the 100 mL beaker.
5. Add 5 drops of phenolphthalein to the NaOH.
6. Rinse and calibrate a pH meter (calibrate it at pH 7 and pH 4).
7. Measure the pH of NaOH (record below). The tip of the pH meter should be submerged for the entire lab.
8. Carefully add the volume of HCl indicated in the chart (running totals are listed; at ‘15’ add 5 mL more).
9. Mix the solution gently with the tip of the pH meter. Measure and record the pH.
10. Continue adding acid and measuring pH as indicated in the chart below.
11. Rinse equipment well with tap water. Extra solutions can be flushed down the drain. Wipe off your lab bench.

a	HCl added (mL)	0	10	15	20	23	24	25	26	27	30	35	40	50
b	Measured pH													

c	Total volume (L)	0.025	0.035											
d	HCl added (mol)	0	0.0020											
e	Net NaOH (mol)	0.0050	0.0030											
f	Net HCl (mol)													
g	$[OH^-]$	0.2												
h	pOH	0.70												
i	$[H^+]$													
j	Predicted pH	13.3												

Questions (reference 15.10):

1. Complete the chart (rows c – i).
2. Graph ‘Predicted pH’ (y-axis) vs. mL ‘HCl added’ (x-axis).
3. Define titration, endpoint, and equivalence point.
4. Based on the shape of your graph, at what volume of HCl added would it be easiest to measure a different pH than what was expected? Explain.
5. Sketch graphs for these situations (for each, write one sentence describing how the graph has changed):
   a) H_2SO_4 is used instead of HCl,
   b) NaOH and HCl are switched (NaOH in burette, 25mL HCl in beaker)
   c) NH_3 (a weak base) is used in place of NaOH (see pg. 642 – 643)
   d) HF (a weak acid) is used in place of HCl.
6. Read “Selecting the Best Acid–Base Indicator” on page 643 – 644. Give one reason why bromothymol blue would be a better choice than phenolphthalein. Give one reason why it is a worse choice. Suggest 2 other acid-base indicators that would work for today’s lab (see pg. 606).

Calculations needed for chart

c. = original 0.025 L of NaOH + volume of HCl added (a)

d. = volume HCl (a) x concentration HCl (0.2 M)

e. = mol NaOH (0.005 mol) - mol HCl (d)

f. = mol HCl (d) - mol NaOH (0.005 mol)

g. = mol NaOH (e) ÷ volume (c)

h. = -log $[OH^-]$ (g)

i. = mol HCl (f) ÷ volume (c)

j. = 14 - pOH (h) or -log $[H^+]$ (i)