TITRATIONS

- Titrations are used to determine the unknown concentration or molar mass of an acid or base
- A solution of unknown concentration is combined with a solution of known concentration (standard solution) until the number of moles of the hydrogen ions equal the number of moles of the hydroxide ions
- When you have stoichiometrically equivalent amounts of the 2 solutions, it is called the **equivalence point**
- To observe the equivalence point, we use an acid/base indicator (otherwise the neutralization reaction starts with 2 clear, colourless solutions and ends with a clear, colourless solution)
- When the indicator changes colour, the **end-point** is reached
- Obviously, it is important to choose an indicator whose endpoint = the equivalence point (or as close to it as possible)
- Notice that as you titrate, the pH changes slowly at first but changes very rapidly near the equivalence point therefore great care and patience must be taken when titrating (drop by drop near this point)

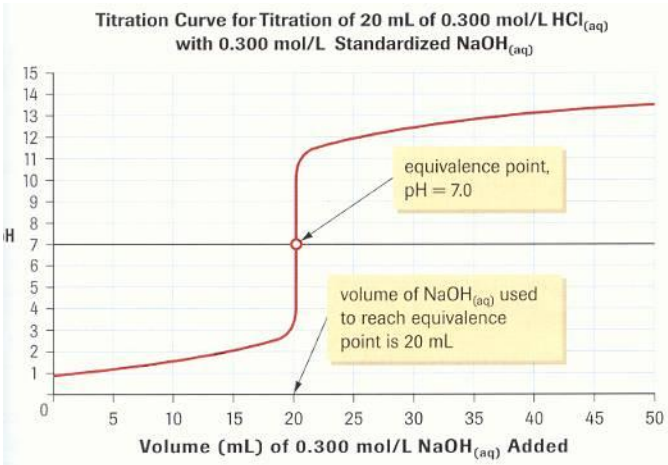

Figure 2
This curve is typical of curves depicting the titration of a strong acid with a strong base. Notice that the curve sweeps up and to the right as $NaOH_{(aq)}$ is added, beginning at a pH below 7 and ending at a pH above 7. The equivalence point is reached at pH 7.

- Acid/base indicators are weak acids that are also dyes
- The acid form has one colour and the base form another colour
- Be careful not to add too much indicator; it will require more base to neutralize your acid than what normally would be expected

$$\underset{\text{(one colour)}}{H\text{-}In} \leftrightarrow H^{+1}(aq) + \underset{\text{(another colour)}}{In^{-1}(aq)}$$

Titration of a Strong Acid and a Strong Base

- Equivalence point = pH 7.0
- The best indicator to use is bromothymol blue because its endpoint range is pH 7 (yellow) to pH 7.6 (blue)

Titration of a Weak Acid and a Strong Base

- Equivalence point = pH >7.0
- The best indicator to use is phenolphthalein because its endpoint range is pH 8.2 (clear) to pH 10.0 (pink)

Titration of a Strong Acid and a Weak Base

- Equivalence point = pH < 7.0
- The best indicator to use might be methyl orange pH 3.2 (red) to pH 4.4 (yellow)

Practice grade 11 titration questions

1. Determine the [NaOH] knowing that it required 25 mL of a 2.5 mol/L hydrochloric acid solution to neutralize 15 mL of the base. (4.17 M)

2. A 1.5 g sample of ascorbic acid (monoprotic) is dissolved in 100 mL of water and titrated with 0.25 M NaOH to the methyl orange endpoint. The volume of base used is 34.1 mL. Calculate the molar mass of vitamin C. (175.95 g/mol)

3. In a titration, you used 300 mL of a 0.85 M solution of calcium hydroxide to neutralize 500 mL of phosphoric acid. What is the phosphoric acid's concentration? (0.34 M)

4. In an experiment, you must perform 2 separate but related titrations to get your results. First, you must standardize your base. Upon doing so, you found it took 50 mL of a 0.5 M hydrochloric acid solution to neutralize 70 mL of sodium hydroxide. Now that you have found the concentration of your base, you must find the molar mass of an unknown, diprotic acid. In this titration, it took 35 mL of your base to neutralize 2.34 g of your acid. (374.5 g/mol)