EXERCISES:

1. Why are different acid-base indicators recommended for different acid-base titrations?

2. Explain the difference between end point and equivalence point.

3. What is an air lock, how does one form, and how is it avoided?

4. Why should the amount of acid-base indicator used be kept to a minimum?

5. Both a volumetric flask and a burette are volume measuring devices, but you should rinse a volumetric flask with distilled water and you should rinse a burette with the solution that is to be run out of the burette. Explain.

6. Explain what is meant by the phrase "weigh out accurately approximately 2 grams".

The following are essay-style questions and require about 2 - 3 pages of written response. Diagrams and equations should be used to illustrate your response, and sub-headings used to denote different sections of the essay.

7. You are required to prepare 250.0 mL of a standard solution of sodium carbonate of concentration approximately 0.0500 mol L^{-1}. Describe how you would do this, emphasising the procedures you would employ to reduce errors and maintain safety.

8. Describe how you would find the concentration of ethanoic (acetic) acid in a bottle of commercial vinegar, given about 500 mL of standard sodium hydroxide solution (whose concentration is known to be 0.105 mol L^{-1}). Your response should describe the procedures you would use, the glassware you would require, and what steps you would take to ensure safe and error-free work.

The following calculation items should be set out in detail, and final answers given to **three significant figures.**

9. To determine the acidity of a wine, an analytical chemist titrated 20.0 mL of the wine with 0.0431 mol L^{-1} sodium hydroxide solution. An average titre of 19.88 mL of alkali was needed. Assuming that all of the acidity of the wine was due to the presence of tartaric acid ($H_2C_4H_4O_6$), a weak diprotic acid,

(a) What indicator should the analyst have used? Explain your answer.

(b) Write a balanced equation for the reaction between the acid and the base used in this titration.

(c) What was the concentration of tartaric acid in the wine, in moles per litre?

(d) Express the concentration of tartaric acid in the wine as a percentage by mass.

10. A chemist was asked to find the percentage purity of a commercial swimming pool additive, sodium carbonate decahydrate ($Na_2CO_3.10H_2O$). The chemist dissolved 1.286 g of the sample in distilled water, and titrated the resulting solution with 0.226 mol L^{-1} hydrochloric acid solution. 16.82 mL of acid was used to completely react with the sodium carbonate in the solution.

(a) What indicator should the analyst have used? Explain your answer.

(b) Write a balanced equation for the reaction between the acid and the base used in this titration.

(c) What was the amount, in moles, of sodium carbonate in the sample?

(d) What was the mass, in grammes, of sodium carbonate in the sample?

(e) What was the percentage by mass of sodium carbonate in the sample?

(f) What was the percentage purity of the sample?

11. A 25.00 mL sample of battery acid (dilute sulfuric acid) was made up to 500.0 mL in a volumetric flask, and 25.00 mL aliquots were titrated with 0.206 mol L^{-1} potassium hydroxide solution. An average titre of 38.8 mL was required.

(a) Calculate the molarity of the original (undiluted) battery acid.

(b) If 50.0 mL of the undiluted battery acid were spilled, what mass of anhydrous sodium bicarbonate would be required to neutralise the spillage?

12. Some lawn sand was known to be made up of a mixture of ammonium sulfate and sand.

2.00 g of this mixture was boiled with 50.0 mL of sodium hydroxide solution. This reacted with the ammonium sulfate, producing ammonia gas and sodium sulfate solution. The resulting solution contained some unreacted sodium hydroxide, which was titrated with 0.100 mol L^{-1} sulfuric acid solution, 30.0 mL of which was required to reach end point.

In a separate experiment, 20.0 mL of this acid solution was needed to neutralise 25.0 mL of the original sodium hydroxide solution. Calculate

(a) the concentration of the original sodium hydroxide solution;

(b) the amount of sodium hydroxide that was added to the sand-ammonium sulfate mixture;

(c) the amount of sodium hydroxide that reacted with the ammonium sulfate;

(d) the mass of ammonium sulfate in the 2.00 g of mixture; and hence

(e) the percentage of ammonium sulfate in the sand-ammonium sulfate mixture.