Chemistry Experiment 2: Titration:

Quantitative Analysis and Performing Calculations of a Standard Solution.

Aim:

Acid-base titration involving a neutralisation reaction that is performed on a standard solution in the laboratory to determine the molarity of sodium carbonate (Na_2CO_3) and to determine the desired end point of hydrochloric acid (HCl), which number of moles of HCl is 0.1012 M.

Upon successful completion of the titration experiment the following correct experimental procedures should follow:

- Observe occupational health and safety standards in a laboratory setting (i.e. glass eyewear protection).
- Put a quantitative thinking cap on in order to learn objectively how to follow lab rules and guidelines to set up an experiment.
- Read the lab manual before undertaking the experiment.
- Understand correctly how to set up laboratory equipment, including a burette.
- Observe, collect, retrieval of correct basic scientific notation to be used in the process of quantifying calculations.
- Perform a standard solution ready for titration.

Method:

Due to the nature of this experiment being conducted with the use of 'fake data' supplied by the lecturer, please refer to the method section outlining the whole titration procedure in the lab manual. Refer to Fudge (2014) pages 150-154 for further information.

Observations:

According to the general chemistry techniques by the Purdue University (n.d), the correct techniques that should be administered for titrations of standard solutions should reflect the two main titration handling methods: ‘scout titration’ and ‘actual titration’.

Before performing a standard titration, learning objectives that are key are:

- Stir but not splash the solution. This makes sure the mixture is mixed properly.
- Clean the flask with distilled water.

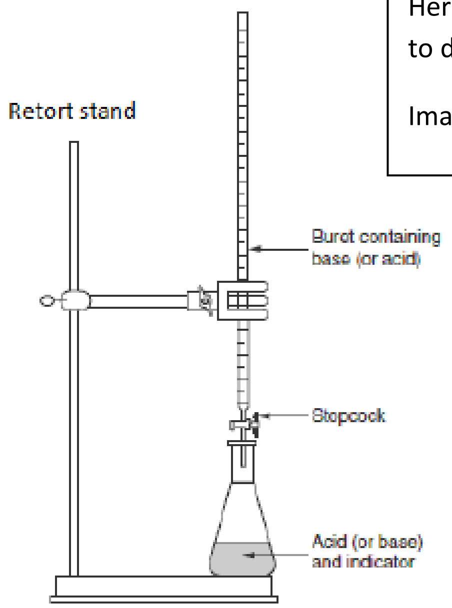
After performing the titration, it was observed that video tips and tricks used by NAIT institute for technology (2008) were employed during apparatus setup. (Note: it is important to remember to employ these tips while doing the titration:

- Rinse three times to make sure all the contaminants are not contained within the buret.
- Pour slowly so the burette doesn’t overflow.
- Limit the amount of bubbles so significant errors are not made
- Use a card with a dark line marker to help record accurate meniscus reading.
- Significant figures should be observed when calculating
- Put the burette onto a white background so any ph/colour indications can remain visible.

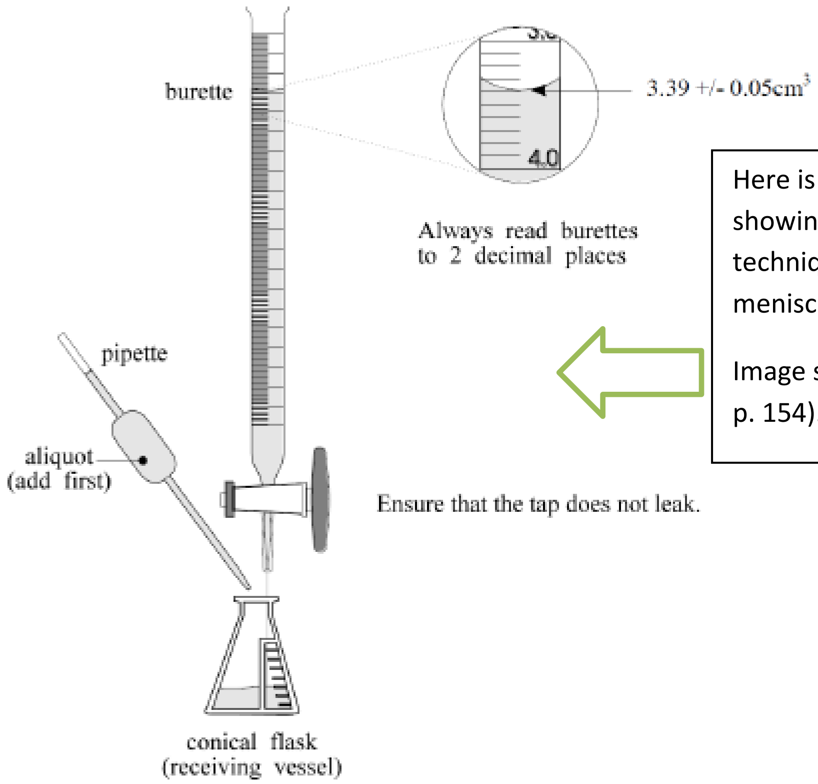
While conducting this titration with ph indicator Methyl Orange it is important to note that Methyl Orange’s (effective ph) ranges are 3.1-4.4(Nicole Dupuis 2008). Methyl Orange is a red colour in an acid and yellow colour in a base, and an orange appearance once the endpoint is reached (Kilbaha Multimedia Publishing 2007).

Here is an Illustration showing how to do a basic titration setup.

Image source: (Fudge 2014, p. 152).



General acid–base titration set-up



Here is an Illustration showing the correct technique for reading the meniscus.

Image source: (Fudge 2014, p. 154).

Chemistry Practical 2: Titration

For 2014 students who participated in the laboratory practical;

The actual concentration of the hydrochloric acid (HCl) was:

0.1034 M

This 0.1034 M (or mol/L) is what you are aiming for in your calculated concentration. This is why in the practical procedure it says that the concentration was approximately 0.1 M and that is what was stated on the bottle in the lab that students used in the titration experiment.

External students (or students who are using the fake data posted online and not their own) note:

Last year the concentration was 0.1012 M

So if the 'fake' data is being used, for example by an external student. Then you are looking for a result close to this value of 0.1012 M not the newer value above, as this is for the fresh solution that was made up by the technical staff this year and it differs in concentration slightly to last year.

Here are the results data sheets supplied by the lecturer.

For externals or those who missed it

'Fake' data to use for Chemistry Practical 3: Titration.

- Weight of sodium carbonate =

1.2905g

- Titration results =

	Initial burette reading /mL	Final burette reading /mL	Titre/mL
Rough determination	0.68	18.32	?
First determination	1.57	18.89	?
Second determination	0.96	18.24	?
Third determination	1.10	18.49	?
Average titre (do not include rough)			?

?

- You will need to calculate the titre values for each & then determine your average titre value.
- Then continue on with the rest of the calculation as normal.

Here are the results data sheets supplied by the lecturer.

Results:

Results

Preparation of the standard solution of Na_2CO_3

1. Record the mass of anhydrous sodium carbonate Na_2CO_3 used ____________________g

2. Calculate the concentration (molarity) of your sodium carbonate solution using the following equations in order.

(a) First calculate the number of moles of Na_2CO_3 used:

n(moles of Na_2CO_3)= **m**(mass of Na_2CO_3)/**MW**(Molecular weight of Na_2CO_3)

Therefore the number of moles of Na_2CO_3 used = ______________mol

(b) Then calculate the concentration of Na_2CO_3 in solution

c(Concentration of Na_2CO_3) = **n**(moles of Na_2CO_3)/**V**(Volume of flask)

Therefore the concentration of the standard solution of Na_2CO_3 is = ______________mol/L

Here are the results data sheets supplied by the lecturer.

Titration results

3. Titration results for Na_2CO_3/HCl neutralisation reaction:

	Initial burette reading/mL	Final burette reading/mL	Titre/mL
Rough determination			
First determination			
Second determination			
Third determination			
Average titre (do not include rough)			

Here are the results data sheets supplied by the lecturer.

Determining the concentration of HCl

4. Calculate the concentration (molarity) of the acid:

(a) Now we know that 20 ml of the standard solution in the conical flask was added. We can use this to calculate the number of moles used on average to neutralise the acid using the equation below.

n(Na_2CO_3) = **c**(of your standard Na_2CO_3solution calculated in part 2b) x **V**(volume in conical flask)

Therefore the number of moles of Na_2CO_3 used to neutralise the acid was = ___________mol

(b) The neutralisation point of the titration occurs when the acid is completely used up. Therefore, we must find out how many moles of Na_2CO_3 it took to neutralise the acid. To do this we must balance the neutralisation equation.

$$Na_2CO_3(aq) \quad + \quad HCl(aq) \quad \rightarrow \quad NaCl(aq) \quad + \quad H_2O(l) \quad + \quad CO_2(g)$$

(c) This allows us to calculate the ratio of the number of moles of Na_2CO_3 moles of HCl

Ratio of n(Na_2CO_3) : n(HCl) = ____________

(d) Now we can work out how many moles of HCl there were in the 20ml aliquot pipetted into the conical flask.

For example if the ratio of $n(Na_2CO_3)$: $n(HCl)$ is 1:3

then $n(HCl) = n(Na_2CO_3) \times 3$

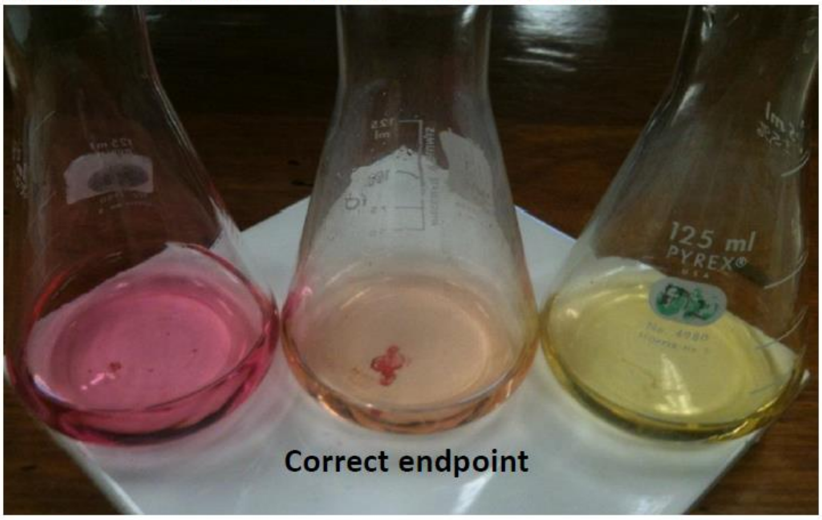
Therefore the number of moles of HCl in the conical flask was = ___________mol

(e) Now we know how many moles of acid (HCl) were present in the 20 ml solution. From the burrette you have an average titre reading, use this to calculate the concentration of acid using the equation below.

c(Concentration of HCl) = **n**(moles of HCl)/**V**(average titre reading)

Therefore the concentration of HCl = ___________mol/L

5. Does the answer you calculate match well (within 10%) of the actual concentration of HCl? If not why not, please think about possible sources of error.



Titration colour changes of Methyl Orange indicator.

Calculations:

Filled in results sheet showing basic titration calculations

Adapted from the science for tertiary learning study guide (Nasc18) 2014

Preparation of the standard solution Na_2CO_3

1. Record the mass of anhydrous carbonate Na_2C0_3 used 1.2905 g.
2. Calculate the concentration (molarity) of your sodium carbonate solution using the following equations in order.

a. First calculate the number of moles of Na_2C0_3 used:

N (moles of Na_2C0_3) = m (mass of Na_2C0_3)/ MW (Molecular weight of Na_2C0_3)

Calculations are as follows:

M= 1.2905 g/ MW=105.99g

Therefore the number of moles of Na_2C0_3 used equals 0.01217567695 mol L^{-1}.

(B) Then calculate the concentration of Na_2C0_3 in solution.

C (Concentration of Na_2C0_3) = n (moles of Na_2C0_3)/ V (Volume of flask)

Calculations are as follows:

0.01217567695 mol L^{-1}/0.250 ml

Therefore the concentration of the standard solution of Na_2C0_3 is = 0.0487027078 mol $L^{-1.}$

Titration Results

3. Titration results for Na_2C0_3/HCl neutralisation reaction:

Determination	Initial burette reading/ml	Final burette reading	Titre/ml
Rough determination	0.68	18.32	17.64
First determination	1.57	18.89	17.32
Second determination	0.96	18.24	17.28
Third determination	1.10	18.49	17.39

Total average titre=	*17.4075 ml*

Determining the concentration of HCl

4. Calculate the concentration (molarity) of the acid:

a) Now we know that 20ml of the standard solution in the conical flask was added. We can use this to calculate the number of moles used on average to neutralise the acid using the equation below:

N(Na_2C0_3)=c(of your standard Na_2C0_3 solution calculated in part 2b) x **V** (volume in conical flask)

0.487027078 ml x 0.020 = 0.00974054156

Therefore the number of moles of Na_2C0_3 used to neutralise the acid was $9.74054156 \times 10^{-4}$ mol.

(b) The neutralisation point of the titration occurs when the acid is completely used up. Therefore, we must find out how many moles of Na_2C0_3 it took to neutralise the acid. To do this we must balance the neutralisation equation.

Na_2C0_3 (aq) + 2HCl (aq) $\longrightarrow$ 2NaCl (aq) + H_2O (l) + CO_2 (g)

b. This allows us to calculate the ratio of the number of moles of Na_2C0_3 moles of HCl.

Ratio of n (Na_2C0_3): n (HCL) = 1: 2.

(d) Now we can work out how many moles of HCl there were in the 20ml aliquot pipetted into the conical flask.

Therefore the number of moles of HCl in the conical flask was $1.948108312 \times 10^{-3}$ mol.

(e) Now we know how many moles of acid (HCl) were present in the 20ml solution. From the burette you have an average titre reading, use this to calculate the concentration of acid using the equation below.

C (concentration of HCl) =n (moles of HCl ($1.948108312 \times 10^{-3}$)/ **V** (average titre reading (0.0174075).

Therefore the concentration of HCl= 0.111912 mol/L

Discussion:

During the endpoint of the actual experiment, the actual concentration of hydrochloric acid (HCl) solution is determined to be 0.111912 mol/L. In this practical this acid-base titration using Na_2CO_3 shows that error composition, also known as titrimetric methodic errors; in this case for example, 'difference between the observed end point and the stoichiometric equivalence point of a reaction' can be present in titrimetric methods of analytical chemistry (Kaur 2008, p. 47). The many forms of different types of errors that can also have been attributed in this experiment are observed in the following below:

- Personal Errors
- Operational Errors
- Instrumental and Reagent Errors
- Methodic Errors
- Additive and Proportional Errors
- Random or Indeterminate Errors
- Errors in Measurements (Kaur 2008, pp. 47-50)

The personal factor that an analyst is responsible for in the lab is making sure 'errors in reading a burette' is not made (Kaur 2008, p. 47). Kaur (2008, p. 49) says that one of the main indicator (causing factor) in gravimetric analysis that can add absolute error in the process of lab work titration is that '… impurity is a standard substance which leads to an incorrect value for the molarity of a standard solution'.

Just how accurate this experiment is and how close is the value concentration 0.111912 mol^{-1}/L is close to the accepted vale of 0.1012 mol/L is needed for further investigations by the analyst. Because the precision darts are very close to the target of 0.1012 mol^{-1}/L this confirms that this scientific experiment is within the 10% of HCl. One way of hoping for an HCl value that is calculated as being accurate and precise is to cross-examine all the data supplied.

True indicators of the propagation of errors could have come about while also counting the numbers of significant figures; this includes calculations involving common digit placeholders. As the analyst is a novice that performed these titration calculations, methodic errors are likely to have occurred.

Thus, this experiment produced a consistency of manually working out how to balance the acid-base neutralisation reaction. Stoichiometric co-efficients in the chemical equation were used to determine if the Methyl Orange indicator performed closely between the end point and the equivalence point (Fudge 2014, p. 152).

Conclusion:

Titrimetric methods were conducted to determine the end point concentration of HCl. The results conducted by the analyst indicate that the correct end point for HCl was reached at 0.111912 mol/L which is within the 10% of the actual concentration of HCl (0.1012 mol/L). Stoichiometric co-efficients were used to balance the neutralisation point of the titration. The number of moles of Na_2CO_3 and HCl were found to be on the ratio of 1:2. However, further testing and analysis to see whether random errors existed should be carried out.

Questions:

Two other indicators (besides methyl orange that was used for this titration) that can be used for acid-base titration include indicators phenolphthalein and cresol purple. Chemical analyst Levie (1996, p. 588) states that the ph range of phenolphthalein is between 8.2-10.0. The equivalence point (transition range) for phenolphthalein is calculated as V_b/V_a= -0.00 to + 2.02(Levie 1996, p. 588).

Another indicator is cresol purple which can also be used for acid-base type titrations. Cresol purple's corresponding ph ranges are listed in the ranges of 7.4-9.0. The transition range for cresol is V_b/V_a= 0.22 to 0.19(Levie 1996, p. 588).

Interestingly, ph indicators generally supply information that can be used by analysts to compute algorithms of titration curves (Levie 1996, p. 585).

It is important to completely dissolve the chemical (i.e. Na_2CO_3) before making the solution up to the meniscus mark in the volumetric flask so that

References

Fudge, A 2014, NASC18 Study guide: Science for tertiary learning, ver 1, University of South Australia, Adelaide.

Kaur, H 2008, *Analytical Chemistry*, Global Media, India.

Kilbaha Multimedia Publishing 2007, *Methyl Orange endpoint*, Kilbaha Multimedia Publishing, viewed 8 June 2014,<https://www.youtube.com/watch?v=3tFUDJTWc88>.

Levie, R 1996, '*General Expressions for Acid-Base Titrations of Arbitrary Mixtures*', Analytical Chemistry, pp. 585-590.

NAIT 2008, *titration technique using a buret*, NAIT, viewed 8 June 2014, <https://www.youtube.com/watch?v=9DKB82xLvNE7>.

Nicole Dupuis 2008, *Acid-Base Indicators*, Nicole Dupuis, viewed 8 June 2014, <https://www.youtube.com/watch?v=6Y4Y-_ME60#t=48>.

Purdue University n.d, *performing the titration*, Purdue University, viewed 8 June 2014, <http://chemed.chem.purdue.edu/genchem/lab/techniques/titration/perform.html>.