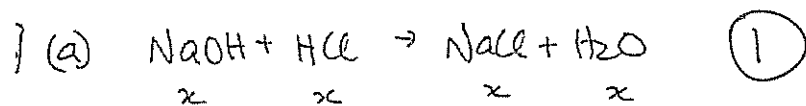


## Exam Practice 3: Titration of Acids and Bases

Aspirin is a weak monoprotic acid of molecular formula  $\text{HC}_9\text{H}_7\text{O}_4$ . In order to determine the aspirin content of a headache tablet, an analyst ground up the tablet and dissolved it in 25.00 mL of  $0.106 \text{ mol L}^{-1}$  sodium hydroxide solution. After complete reaction, the excess alkali was titrated with  $0.0958 \text{ mol L}^{-1}$  hydrochloric acid solution, 10.33 mL being required.

- (a) Calculate the mass of aspirin in the tablet.
- (b) The label on the packet stated that each tablet contains 300 mg of aspirin. Comment on this claim.

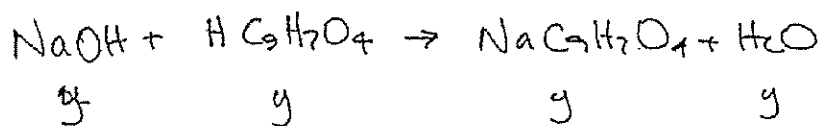
[7 Marks]



$$n_{\text{NaOH}} = n_{\text{HCl}} = cV = (0.0958)(0.1033) \text{ mol} = 9.896 \times 10^{-4} \text{ mol NaOH} \quad (1)$$

$$\text{Original } n_{\text{NaOH}} = cV = (0.106)(0.025) \text{ mol} = 2.65 \times 10^{-3} \text{ mol NaOH} \quad (1)$$

$$\therefore n_{\text{NaOH reacted with aspirin}} = (2.65 \times 10^{-3} - 9.896 \times 10^{-4}) \text{ mol NaOH} \\ = 1.66 \times 10^{-3} \text{ mol NaOH} \quad (1)$$



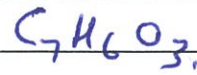
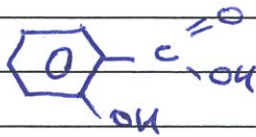
$$\therefore n_{\text{HC}_9\text{H}_7\text{O}_4} = 1.66 \times 10^{-3} \text{ mol aspirin} \quad (1)$$

$$\therefore M_{\text{aspirin}} = nM \\ = (1.66 \times 10^{-3})(8.064 + 108.09 + 64) \text{ g} \\ = 0.299 \text{ g aspirin} \quad (1)$$

(b) Pretty close - probably within experimental uncertainty. (1)

Back Titration -

Salicylic Acid



$$M_r = 138$$

liposol.

$$K_a = 1.1 \times 10^{-3}$$

known HCl 31.5 ml 0.05M at 0.1 ml  $l^{-1}$   
NaOH 0.1M  
mmol/tablet 300mg

$$\begin{aligned} n(\text{HCl}) &= CV \\ &= 0.05 \times (31.5 \times 10^{-3}) \\ &= 1.575 \times 10^{-3} \text{ mol} \end{aligned}$$



$$n(\text{HCl}) = n(\text{NaOH})_{\text{excess}}$$

$$\begin{aligned} n(\text{NaOH})_{\text{total}} &= CV \\ &= 0.1 \times (20 \times 10^{-3}) \\ &= 0.002 \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{NaOH})_{\text{reacted}} &= n(\text{NaOH})_{\text{total}} - n(\text{NaOH})_{\text{excess}} \\ &= 0.002 - (1.575 \times 10^{-3}) \\ &= 4.25 \times 10^{-4} \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{NaOH})_{\text{reacted}} \times \frac{200}{20} &= n(\text{NaOH})_{\text{total}} \\ &= 4.25 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} n(\text{NaOH})_{\text{reacted}} \times 2 &= n(C_7H_6O_3) \\ &= 2.125 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} M_r \text{ of } C_7H_6O_3 &= 138 \\ n &= \frac{m}{M_r} \quad 2.125 \times 10^{-3} = \frac{m}{138} \end{aligned}$$

$$m \text{ of } 0.293 \text{ g}$$

or 293 mg

$$\% \text{ purity} = \frac{0.293}{\text{mmol of tablet}} \times 100 = 100 \times \frac{0.293}{0.300} = 97.6\%$$

**Worksheet 8.2****A back titration**

NAME:

CLASS:

**INTRODUCTION**

Marble is a metamorphic rock that is almost pure calcium carbonate. The following experiment was conducted to determine the calcium carbonate content of a marble sample. An accurately weighed sample of crushed marble was added to a measured volume of recently standardised HCl solution. The solution was heated to drive off the evolved carbon dioxide. The remaining solution was titrated with a recently standardised NaOH solution, using a methyl red indicator. The results obtained are shown below.

Mass of marble sample: 1.740 g

Volume of HCl solution added: 40.00 mL

Concentration of the standardised HCl solution:  $1.020 \text{ mol L}^{-1}$ Concentration of the standardised NaOH solution:  $0.275 \text{ mol L}^{-1}$ 

Average titre of NaOH: 25.56 mL

No.	Question	Answer
1	Write an equation for the reaction of HCl with: a $\text{CaCO}_3(\text{s})$ b $\text{NaOH}(\text{aq})$	
2	Calculate: a the amount (in mol) of HCl added initially b the amount (in mol) of NaOH used in the titration c the amount (in mol) of unreacted HCl d the amount (in mol) of HCl reacting with the $\text{CaCO}_3$ e the amount (in mol) of $\text{CaCO}_3$ in the marble sample f the mass (in g) of $\text{CaCO}_3$ in the marble sample g the percentage by mass of $\text{CaCO}_3$ in the marble sample.	

**Worksheet 8.2****A back titration**

No.	Question	Answer
3	The sodium hydroxide solution used was recently standardised. Why is sodium hydroxide unsuitable as a primary standard?	
4	Why was it necessary to drive off the evolved carbon dioxide before performing the titration? How would the result be affected if this step was omitted?	
5	Why was it necessary to use a back titration for this analysis, rather than a direct titration of the marble with HCl solution?	
6	How would each of the following errors, if made during the analysis, alter the calculated value for the percentage $\text{CaCO}_3$ ? a The 40.0 mL pipette used to deliver the HCl was rinsed only with water prior to its use. b The burette was rinsed only with water prior to its use. c The volumetric flask was rinsed only with water prior to its use.	
7	An alternative method of analysis involves reacting the crushed marble with excess HCl and collecting the evolved carbon dioxide. In one such experiment, 95.0 mL of gas was collected at a pressure of 765 mmHg at 23°C when 0.411 g of marble was reacted. Determine the percentage by mass of $\text{CaCO}_3$ in the marble sample, based on this data.	

**Worksheet 8.2****A back titration**

<b>No.</b>	<b>Question</b>	<b>Answer</b>
8	Suggest reasons why the value obtained by this method is smaller than that obtained using the back titration.	

**Worksheet 8.2: Solutions****A back titration**

No.	Answer
1	a $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ b $\text{HCl}(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
2	a $n(\text{HCl}) \text{ initially} = c \times V = 1.020 \times 40.00 \times 10^{-3} = 4.080 \times 10^{-2} \text{ mol}$ b $n(\text{NaOH}) = c \times V = 0.275 \times 25.56 \times 10^{-3} = 7.029 \times 10^{-3} \text{ mol}$ c $n(\text{HCl}) \text{ unreacted} = n(\text{NaOH}) = 0.275 \times 25.56 \times 10^{-3} = 7.029 \times 10^{-3} \text{ mol}$ d $n(\text{HCl}) \text{ reacting} = n(\text{HCl}) \text{ initially} - n(\text{HCl}) \text{ unreacted}$ $= 4.080 \times 10^{-2} - 7.029 \times 10^{-3} = 3.377 \times 10^{-2} \text{ mol}$ e $n(\text{CaCO}_3) = \frac{1}{2} \times n(\text{HCl}) \text{ reacting} = \frac{1}{2} \times 3.377 \times 10^{-2} = 1.689 \text{ mol}$ f $m(\text{CaCO}_3) = n \times M = \frac{1}{2} \times 3.377 \times 10^{-2} \times 100.09 = 1.690 \text{ g}$ g $\% \text{ CaCO}_3 \text{ in marble} = \frac{m(\text{CaCO}_3)}{m(\text{marble})} \times \frac{100}{1} = \frac{1.690}{1.740} \times \frac{100}{1} = 97.1\%$
3	Solid sodium hydroxide absorbs water from the atmosphere. Solid samples may therefore be damp and therefore impure. Sodium hydroxide solutions react with carbon dioxide in the atmosphere, decreasing the concentration of the solution: $2\text{NaOH}(\text{aq}) + \text{CO}_2(\text{g}) \rightarrow \text{Na}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$
4	$\text{CO}_2$ is an acidic oxide. It reacts with NaOH. If the $\text{CO}_2$ was not removed, more NaOH would be required for the titration, leading to an increased value for the unreacted HCl, and hence a decreased value for the $\text{CaCO}_3$ percentage.
5	Calcium carbonate is not soluble. The carbonate ion is a weak base that gives an indistinct endpoint in a direct titration.
6	a This would dilute the acid. More would be required to react with the marble, giving an increased percentage for the carbonate. b This would dilute the NaOH. More would be required to react with the HCl, giving an increased value for the unreacted HCl. This, in turn, would give a decreased value for the reacting HCl, and so a decreased percentage for the carbonate. c This has no effect. Flasks should be rinsed with water.
7	$n(\text{CO}_2) = \frac{pV}{RT} = \frac{765 \times 101.3 \times 95.0 \times 10^{-3}}{760 \times 8.314 \times 296} = 0.003\,936 \text{ mol}$ $n(\text{CaCO}_3) = n(\text{CO}_2) = 0.003\,936 \text{ mol}$ $m(\text{CaCO}_3) = n \times M = 0.003\,936 \times 100.09 = 0.3940 \text{ g}$ $\% \text{ CaCO}_3 \text{ in marble} = \frac{m(\text{CaCO}_3)}{m(\text{marble})} \times \frac{100}{1} = \frac{0.3940}{0.411} \times \frac{100}{1} = 95.9\%$
8	Small amounts of gas may have been lost in the collection process. Some carbon dioxide may remain dissolved in the reaction solution.