

Extended Essay
Chemistry

SOLUBILITY OF CALCIUM CARBONATE INVESTIGATION
BASED ON COMPLEXOMETRIC TITRATION OF CALCIUM
AMOUNT DETERMINATION

ABSTRACT

Calcium carbonate is a main component of fur which is usually deposited on bathroom equipment, boilers and inside household kettles. To remove fur it is recommended to use specific liquids designed for fur dissolving which usually contain different types of acids.

The aim of the essay is to find out the answer to the question: is calcium carbonate soluble in water and in acids – strong, weak and of medium strength? Additionally, it shows which of these acids is the most effective in dissolving calcium carbonate.

The investigation is based on the method of calcium amount determination using complexometric titration. The first part is devoted to determination of solubility of calcium carbonate in water and based on the solubility product. The second one is investigation on the effect of electrolytes of different strengths on the solubility of calcium carbonate. The chosen electrolytes are: hydrochloric acid (strong electrolyte), methanoic acid (electrolyte of medium strength) and ethanoic acid (weak electrolyte). To find out which of those electrolytes is the most effective in dissolving calcium carbonate 1 molar solutions of each acid were used but the amount of acids was manipulated.

The essay proves all acids used in the investigation which dissolve calcium carbonate but methanoic acid is less effective than ethanoic acid and hydrochloric acid. In the case of methanoic acid the greater amount of the acid must be used to dissolve calcium carbonate completely than in the case of the other two acids used in the investigation.

Though calcium carbonate seems to dissolve in water due to its solubility product it was impossible to confirm it as the method used was not sensitive enough to determine the amount of calcium in the prepared solution.

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I. INTRODUCTION

1.1 Description

In my essay I will investigate the solubility of calcium carbonate basing on complexometric titration.

The choice of such an investigation was not the matter of coincidence. I got interested on the solubility of calcium carbonate when I read that it is main compound responsible for formation of fur. As fur is hard to remove I had always a lot of problems with it especially during cleaning bathroom or the inside of the kettle. Usually I had to use special liquids designed to dissolve fur as using water only did not have any effect. The one thing that surprised me was the fact that different detergents contained different types of acids that could dissolve fur. That is why I decided to check if calcium carbonate is soluble in water and acids that belong to groups of strong electrolytes, weak electrolytes and electrolytes of medium strength. I knew that if it is soluble it will form ions in the solution. Then I will check the amount of calcium ions in the solution and on such basis I will be able to find out the amount of calcium carbonate dissolved.

To begin my investigation I had to gather information about calcium carbonate itself, about fur formation, about electrolytes and about the most recommended method to determine the amount of calcium in the solution which is presented in the theoretical part of my essay.

In my experiment I would like to find out the most effective method of dissolving calcium carbonate. Firstly would like to check if calcium carbonate dissolves in water - I will base on solubility product. Then I will investigate an effect of electrolytes of different strength on the solubility of calcium carbonate. For the experiment I am going to use three acids:

- a. hydrochloric acid – as a strong electrolyte
- b. methanoic acid – as an electrolyte of medium strength
- c. ethanoic acid – as a weak electrolyte

I will check how the amount of those acids influences on the solubility of calcium carbonate. To find it out I will use complexometric titration - I will determine the amount of calcium ions in the solution and then compare it with the theoretical one – basing on the amount of calcium carbonate I took for preparation of the solution. If the results will be the same or very similar it will mean that the type of electrolyte and the amount of it is the proper one.

1.2 Research question:

Is calcium carbonate (CaCO_3) soluble in water and in acids: strong, weak and of medium strength ?

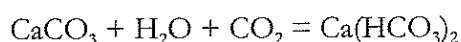
1.3 Hypothesis:

If calcium carbonate is soluble in acids – strong, weak and of medium strength – then it will be possible to determine the content of calcium amount in the solution using complexometric titration.

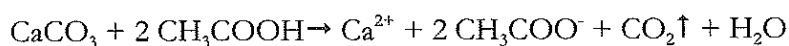
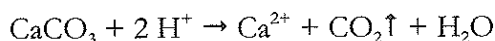
II. THEORETICAL BACKGROUND

1.1 Calcium Carbonate¹

CaCO₃ is a white crystalline powder hardly soluble in water. In 100g of water at 18°C 1.3x 10⁻³ to 3.5x 10⁻³ g of calcium carbonate can be dissolved. 100 times more CaCO₃ can be dissolved in water saturated with CO₂ as calcium hydrogen carbonate is then formed.



CaCO₃ is soluble in mineral acids and in ethanoic acid with release of CO₂



Application

Calcium carbonates have great application in technical industry as:

- a. a building material
- b. a material to obtain burnt lime
- c. a flux in steelworks

1.2 Hardness of water²

Calcium carbonate which is present in rocks and soil is hardly soluble in water. When subjected to the action of carbon dioxide calcium carbonate in the presence of water forms calcium hydrogen carbonate - Ca(HCO₃)₂. Calcium hydrogen carbonate is an acidic salt which is soluble in water and therefore it causes an increase in calcium ions concentration in water.

Calcium and magnesium salts dissolved in water causes so called hardness of water. In everyday life it can be noticed while washing or cleaning using a soap. – instead of lather a sticky sediment is formed which is deposited on hands and materials. Such phenomenon is caused by

¹ Description based on: A. Smockiewiczowa, *Materiały do wykładów z chemii nieorganicznej*. Wydawnictwo II, Poznań 1987

² Description based on: Z. Sołtys, *Chemia nieorganiczna*. PWSZ, Warszawa 1986

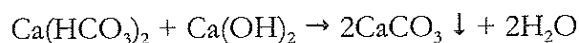
reaction between solution of sodium soap and calcium or magnesium ions which leads to formation of insoluble calcium or magnesium soaps.

There are two types of hardness of water: permanent and temporary

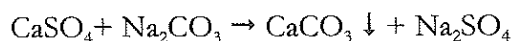
- temporary hardness of water is caused by the presence of calcium hydrogen carbonate in water. It can be removed by heating water to 70°C:



Or by adding calcium hydroxide



- permanent hardness of water is caused by dissolved magnesium and calcium sulphates (VI) and chlorides. Permanent hardness can be diminished by precipitating calcium and magnesium salts from the solution by adding sodium carbonate, borax ($\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{H}_2\text{O}$) and other softening agents:



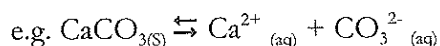
The most efficient way of removing hardness of water is distillation.

Hardness of water can pose problems not only while washing but also in the industry especially when using water to supply steam boilers. Evaporated water coming from rivers or lakes leaves precipitate which stick to hot sides of the boiler forming **fur**. A layer of fur has much lower conductivity than steel and therefore amount of fuel used is greatly increased. Fur can be also deposited on sided and heater of household kettles or water pipes. That is why methods of softening water (described above) should be applied to avoid deposition of fur.

In household the most popular way of removing fur is dissolving it with acid. Different types of acids are used for different purposes. Ethanoic acid and malic acid are used for removing fur from kettles while orthophosphoric (V) acid and hydrochloric acid are preferred in washing bathroom devices.

1.3 Solubility product³

Many salts which we refer to as insoluble dissolve but only to small extent. In a saturated solution, dissolved ions and undissolved salt are in equilibrium



the product of the molar concentrations of calcium ions and carbonate ions is called

the solubility product K_{sp} of calcium carbonate.

$$K_{sp} = C_{\text{Ca}^{2+}} \cdot C_{\text{CO}_3^{2-}}$$

The solubility product of a salt is the product of the molar concentrations of the ions in a saturated solution of the salt, raised to the appropriate powers.

The solubility can be expressed in grams of solute dissolved in dm^3 of solution or in the moles of solute in 1dm^3 of solution at a stated temperature.

Solubility product can be applied to:

- A. Qualitative analysis
- B. Precipitation titration

1.4 Electrolytes⁴

Electrolyte is a substance which is able to produce ions.

Ions can be produced in a process of:

- dissociation (decomposition into ions in a presence of polar solvent e.g. water, alcohol) e.g.
 $\text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^-$
- melting e.g.
 $\text{PbCl}_2 \rightleftharpoons \text{Pb}^{2+} + 2\text{Cl}^-$

Degree of dissociation is a ratio of molecules that underwent dissociation (n) to the total number of molecules. It is expressed in a formula:

$$\alpha = n / N \quad [\cdot 100\%]$$

Due to degree of dissociation three types of electrolytes can be distinguished:

- 1) Strong - $\alpha \approx 1$ e.g. HCl, HBr, H_2SO_4 , KOH, NaOH, salts
- 2) Medium - $0.1 < \alpha < 1$ e.g. HF, LiOH, HCOOH, H_3PO_4
- 3) weak - $\alpha < 0.1$ e.g. $\text{NH}_3 \cdot \text{H}_2\text{O}$, $\text{Mg}(\text{OH})_2$, HNO_2 , H_2S , CH_3COOH

³ Description based on: E.N. Ramsden, *A-level Chemistry*. Stanley Thornes (Publishers) Ltd., Cheltenham 1994

⁴ Description based on: E.N. Ramsden, *A-level Chemistry*. Stanley Thornes (Publishers) Ltd., Cheltenham 1994

1.5 Titration⁵

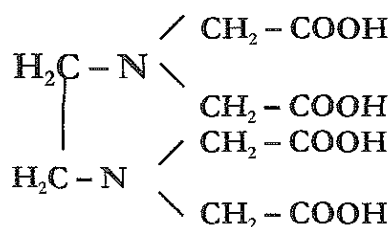
Volumetric analysis is a means of finding the concentration of a solution. The method is to add a solution of e.g. an acid to a solution of e.g. a base, in a measured way until there is enough of the acid to neutralise the base. This method is called titration. The concentration of one of the two solutions must be known, and the volume of both must be measured. Titre – the solution with known concentration is added drop by drop from a burette to the conical flask containing a solution with unknown concentration (the volume of a solution must be measured). The titre must be added until 'end point' (equivalence point) is reached. The change of a colour of an indicator shows when exactly the right volume of solution has been added to achieve equivalence point.

There are several types of titration:

- Acid – base titration – which bases on the neutralisation reaction
- Redox titration – which bases on the reactions of oxidation and reduction
- Precipitation titration – which bases on the reactions of formation of hardly soluble precipitates
- Complexometric titration – which bases on the formation of complex ions
- Conductometric titration – which involves flow of the current which measures conductivity of ions in the solution

Complexometric titration⁶

This type of titration bases on forming stable complexes between determined ion and a substance used as a titre. Metals are the substances that can be determined using complexometric titration as they can easily form complexes. As a titre - organic compounds which are ligands are used. The most commonly used organic compound is ethylenediaminetetraethanoic acid - EDTA (versenic acid). Its formula is:



⁵ Description based on: E.N. Ramsden, *A-level Chemistry*. Stanley Thornes (Publishers) Ltd., Cheltenham 1994

⁶ J. Piaseczki i J. Kotowska, *Jakościowa i ilościowa analiza chemiczna*. Wydawnictwo Brasika, Szczecin 1996

As versenic acid is hardly soluble in water more often, as a titre, disodium versenate is used (called complexon III).

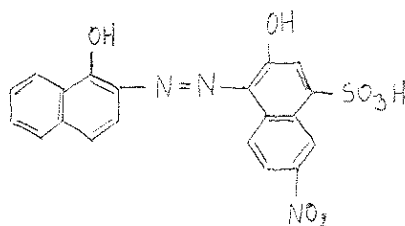
Solutions of versenic acid and disodium versenate are colourless that its why in complexometric titration proper indicators must be used.

Substances the most often used as indicators in complexometric titration are:

▫ Eriochrome Black T

An indicator used in titration of metal ions (e.g. magnesium or zinc). Depending on the pH eriochrome black can appear in three colours. In slightly alkaline solutions (pH 7 – 11) used when titrating metals it has blue colour. In buffered solutions (pH 9 – 10) blue form of eriochrome black forms violetish complexes with some of the metals (e.g. Mg, Zn). Since eriochrome black solutions are unstable the indicator is used in a form of mixture with sodium chloride (powder form) or with ethanol (liquid form).

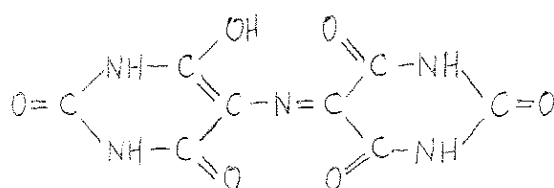
Formula:



▫ Murexide

An indicator used in titration with EDTA for determination of calcium ions in alkaline environment (pH 12 – 13) and nickel and cobalt ions in ammonia environment. Murexide is violet in pH 12 – 13 and its complex formed with calcium ions is pink.

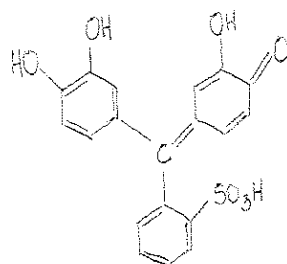
Formula:



▫ Pyrocatechol violet

An indicator used for titration of acidic solutions. It is usually used for titration of bismuth and thorium. In pH 1.5 – 6 solution of this indicator is yellowish and its complexes with metal ions are blue.

Formula:



IV. EXPERIMENTAL PART

1.1 Method used for complexometric determination of calcium ions

1.1.1. Equipment needed :

- Burette
- Pipette
- Conical flask
- Measuring flasks
- Mortar
- Scale

1.1.2. Reagents:

- 0.05M solution of disodium versenate
- 2M solution of NaOH
- Mixture of murexide and NaCl
- 1M solution of HCl
- 1M solution of HCOOH
- 1 M solution of CH₃COOH
- Calcium carbonate - CaCO₃ – crystals
- distilled water

/preparation of reagents is presented in the appendix, pages 25-26/

1.1.3. Procedure:

- I. Take 10cm^3 of solution containing calcium ions and pour it into a conical flask
- II. Add 5cm^3 of 2M solution of NaOH into the conical flask
- III. Add pinch of murexide mixture - solution changes colour into pink
- IV. Titrate using 0.05M solution of disodium versenate till pink solution changes colour into violet
- V. Titration repeat three or four times to have precise results
- VI. Amount of calcium in the solution must be calculated using the following formula:

$$x = \frac{V\text{cm}^3 \cdot ZM \cdot 40.08 \text{ g} \cdot c}{1000 \text{ cm}^3 \cdot 1M}$$

Where:

x - number of grams of calcium

V - amount of titre used

Z M - molar concentration of disodium versenate

c - commensurability

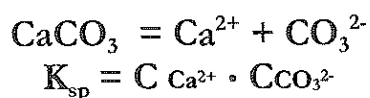
Formula is based on reaction where:

1000 cm^3 of 1M solution of disodium versenate \rightarrow (reacts with) 40.08g of calcium

$V\text{ cm}^3$ of Z M solution of disodium versenate \rightarrow (reacts with) x g of calcium

1.2 Checking solubility of calcium carbonate according to solubility product

Basing on the solubility product I checked how many moles of CaCO_3 can be soluble in 1000 cm^3 of water.



$$C_{\text{Ca}^{2+}} = x$$

$$C_{\text{CO}_3^{2-}} = x$$

$$C_{\text{CaCO}_3} = x$$

$$K_{\text{sp}} \text{ of } \text{CaCO}_3 = 4.8 \cdot 10^{-9}$$

K_{sp} – solubility product

C – molar concentration of ions

$$K_{\text{sp}} = x^2$$

$$x = \sqrt{4.8 \cdot 10^{-9}}$$

$$x = 6.9 \cdot 10^{-5}$$

$$x = 0.00007 \text{ mole dm}^{-3}$$

Now I could change moles into grams to find how many grams of CaCO_3 I should dissolve in 1 dm^3 of water.

$$M_{\text{CaCO}_3} = 100.09 \text{ g/mole}$$

$$1 \text{ mole of } \text{CaCO}_3 - 100.09 \text{ g}$$

$$0.00007 \text{ mole of } \text{CaCO}_3 - x \text{ g}$$

$$x = 0.007 \text{ g of } \text{CaCO}_3$$

I could also calculate how many grams of calcium should be in a solution containing 0.007 grams of CaCO_3 .

$$\text{In } 100.09 \text{ g of } \text{CaCO}_3 - 40.08 \text{ g of Ca}$$

$$\text{In } 0.0070 \pm 0.0001 \text{ g of } \text{CaCO}_3 - x$$

$$\underline{x = 0.0028 \text{ g of Ca}}$$

Having counted the amount of calcium carbonate that can be dissolved in 1000 cm^3 of water I could prepare solution.

I weighed $0.0070 \pm 0.0001 \text{ g}$ of calcium carbonate and put into measuring flask for 1000 cm^3 . I filled the flask to reach 1000 cm^3 .

Now using complexometric titration and procedure described before I could find out how many grams of calcium there are in my solution. The results are as follow:

Titration – 3 trials

	1 st trial	2 nd trial	3 rd trial
Volume of titre used [cm ³ ±0.1]	2.0 ± 0.1 cm ³	2.0 ± 0.1 cm ³	2.0 ± 0.1 cm ³

Using the formula presented in procedure I could count how many grams of calcium are present in the solution.

$$x = \frac{V \text{ cm}^3 \cdot ZM \cdot 40.08 \text{ g} \cdot c}{1000 \text{ cm}^3 \cdot 1M}$$

Where:

x - number of grams of calcium

V - amount of titre used

Z M - molar concentration of disodium versenate

C – commensurability

$$X = (2.0 \pm 0.1 \cdot 0.05 \cdot 40.08 \cdot 100 \cdot : 1000 = 0.401 \pm 0.020 \text{ g of Ca}$$

.....
Error analysis

$$\Delta X = 0.1/2.0 \cdot 0.401 = 0.020$$

Now I could compare the theoretical amount of Ca that should be in solution containing 0.007g of calcium carbonate with the amount of Ca that I determined using titration.

Theoretical number of Ca: 0.0028g

Determined number of Ca: 0.401 ± 0.020g

As both values greatly differ (by 0.3982 g) this suggests that the method which was used is too less sensitive and could not accurately pick out such small number of grams of Ca that was present in the solution.

1.3 Determination of the concentration of calcium carbonate which produces the most accurate results when being dissolved

Knowing that I cannot use the amount of calcium carbonate calculated using solubility product I had to follow the procedure given in the literature⁷. The procedure suggests:

- a. amount of CaCO_3 should equal 0.9000 to 1.0099 g for 1000cm^3 of solution
- b. weighed amount of calcium carbonate should be transferred into measuring flask
- c. *calcium carbonate should be dissolved in 25cm^3 of 1M HCl*
/preparation of 1M hydrochloric acid – see appendix, page 25/
- d. flask should be filled to reach 1000cm^3

I decided to take 5 different concentrations of calcium carbonate to check on which concentration the amount of calcium can be determined the most accurately. The concentrations are:

1. 0.08 g of calcium carbonate per 100cm^3 of solution
2. 0.09 g of calcium carbonate per 100cm^3 of solution
3. 0.10 g of calcium carbonate per 100cm^3 of solution
4. 0.11 g of calcium carbonate per 100cm^3 of solution
5. 0.12 g of calcium carbonate per 100cm^3 of solution

Basing on the procedure for complexometric titration I could determine amount of calcium in each solution.

The results of titration are:

Concentration of calcium carbonate in the solution [g / $100\text{cm}^3 \pm 0.0001$]	Volume of titre used [$\pm 0.1 \text{ cm}^3$]		
	1 st trial	2 nd trial	3 rd trial
0.0802	1.8	1.7	1.7
0.0907	1.9	1.9	1.9
0.1001	2.0	2.0	2.0
0.1100	2.4	2.4	2.4
0.1207	2.6	2.6	2.6

⁷ J. Piaseczki i J. Kotowska, *Jakościowa i ilościowa analiza chemiczna*. Wydawnictwo Brasika, Szczecin 1996

Using the formula I could calculate the exact amount of calcium in each solution. I also calculated the theoretical amount of calcium that should be present in each solution. The results are presented in the table.

Calculations

$$x = \frac{V \text{ cm}^3 \cdot ZM \cdot 40.08 \text{ g} \cdot c}{1000.0 \text{ cm}^3 \cdot 1M}$$

Where:

x - number of grams of calcium

V - amount of titre used

Z M - molar concentration of disodium versenate

C – commensurability

For 0.0802g of calcium carbonate

$$x = \frac{1.7 \pm 0.1 \text{ cm}^3 \cdot 0.05M \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1M} = 0.034 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/1.7 \cdot 0.034 = 0.002$$

Theoretical value

In 100.09 grams of calcium carbonate - 40.08 grams of calcium

In 0.0802g grams of calcium carbonate - x

$$x = 0.032 \text{ g of calcium}$$

For 0.0907g of calcium carbonate

$$x = \frac{1.9 \pm 0.1 \text{ cm}^3 \cdot 0.05M \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1M} = 0.038 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/1.9 \cdot 0.038 = 0.002$$

Theoretical value

In 100.09 grams of calcium carbonate - 40.08 grams of calcium

In 0.0907g grams of calcium carbonate - x

$$x = 0.036 \text{ g of calcium}$$

For 0.1001g of calcium carbonate

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

Theoretical value

In 100.09 grams of calcium carbonate - 40.08 grams of calcium

In 0.1001g grams of calcium carbonate - x

$$x = 0.040 \text{ g of calcium}$$

For 0.1100g of calcium carbonate

$$x = \frac{2.4 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.048 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.4 \cdot 0.048 = 0.002$$

Theoretical value

In 100.09 grams of calcium carbonate - 40.08 grams of calcium

In 0.1100g grams of calcium carbonate - x

$$x = 0.044 \text{ g of calcium}$$

For 0.1207g of calcium carbonate

$$x = \frac{2.6 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.052 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.6 \cdot 0.052 = 0.002$$

Theoretical value

In 100.09 grams of calcium carbonate - 40.08 grams of calcium

In 0.1207g grams of calcium carbonate - x

$$x = 0.048 \text{ g of calcium}$$

Results are presented in the table:

Concentration of calcium carbonate in the solution [g /100cm ³ ± 0.0001]	Amount of calcium in the solution [g ± 0.002]	Theoretical amount of calcium in the solution [g]
0.0802	0.034	0.032
0.0907	0.038	0.036
0.1001	0.040	0.040
0.1100	0.048	0.044
0.1207	0.052	0.048

First three results fit in the uncertainty range but the third one is the most precise. That is why I decided to use the concentration of calcium carbonate 0.10g/100cm³ for the further investigations. The last two results differ from the theoretical ones which suggest that calcium carbonate was not fully dissolved while applying 2.5 cm³ of 1 M solution of hydrochloric acid.

1.4 Investigation on the effect of strong electrolyte – hydrochloric acid (HCl) on the solubility of calcium carbonate

Next part of my investigation was to check how different amount of hydrochloric acid can influence on the solubility of calcium carbonate. I based on the concentration of calcium carbonate which turned to be the most accurate when finding the amount of calcium in the solution – 0.10 g/100cm³. The concentration of hydrochloric acid should be constant – 1 molar solution. /preparation of 1M hydrochloric acid – see appendix, page 25/

Quantities of hydrochloric acid I decided to take:

1. $0.5 \pm 0.1 \text{ cm}^3$ of 1M HCl per 100cm^3 of solution
2. $1.5 \pm 0.1 \text{ cm}^3$ of 1M HCl per 100cm^3 of solution
3. $2.5 \pm 0.1 \text{ cm}^3$ of 1M HCl per 100cm^3 of solution
4. $4.5 \pm 0.1 \text{ cm}^3$ of 1M HCl per 100cm^3 of solution
5. $6.5 \pm 0.1 \text{ cm}^3$ of 1M HCl per 100cm^3 of solution

Although I had calculated amount of calcium in the solution where 2.5 cm^3 of 1 M HCl was used in the previous part of my investigation, I decided to check it once again when doing next titrations.

The results of titration are following:

Amount of acid [$\text{cm}^3 \pm 0.1$]	Volume of titre used [$\pm 0.1 \text{ cm}^3$]		
	1 st trial	2 nd trial	3 rd trial
0.5	0.8	0.8	0.8
1.5	1.8	1.8	1.8
2.5	2.0	2.0	2.0
4.5	2.0	2.0	2.0
6.5	2.0	2.1	2.0

Having the results of titration I could calculate the amount of calcium present in the solution. From the previous part of my experiment I knew that the amount of calcium ions in the solution containing 0.10g of calcium carbonate should equal 0.040g. Using such data I could check on which concentration of hydrochloric acid whole amount of calcium carbonate was dissolved. The data were processed and shown in the table.

$$x = \frac{V\text{cm}^3 \cdot ZM \cdot 40.08 \text{ g} \cdot c}{1000 \text{ cm}^3 \cdot 1M}$$

Where:

x - number of grams of calcium

V - amount of titre used

Z M - molar concentration of disodium versenate

C - commensurability

For 0.5 cm³ of HCl

$$x = \frac{0.8 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.016 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/0.8 \cdot 0.016 = 0.002$$

For 1.5 cm³ of HCl

$$x = \frac{1.8 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.036 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/1.8 \cdot 0.036 = 0.002$$

For 2.5 cm³ of HCl

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 4.5 cm³ of HCl

$$x = \frac{2.0 \text{ cm}^3 \pm 0.1 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 6.5 cm³ of HCl

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

Amount of HCl in the solution [cm ³ ± 0.1]	Amount of calcium in the solution [g ± 0.002]	Theoretical amount of calcium in the solution [g]
0.5	0.016	0.040
1.5	0.036	
2.5	0.040	
4.5	0.040	
6.5	0.041	

The results show that 0.10 g of calcium carbonate can be dissolved completely in 2.5 or more cm³ of 1 M solution of hydrochloric acid. In the lowest concentrations of HCl per 100cm³ of water the salt is dissolved only to some extent.

1.5 Investigation on the effect of a medium electrolyte – methanoic acid (HCOOH) on the solubility of calcium carbonate

In this part of my investigation I checked how different amount of electrolyte of medium strength – methanoic acid can influence on the solubility of calcium carbonate. Similarly as in the previous part I used solution of 0.10g of calcium carbonate in 100cm³. The concentration of methanoic acid should be constant – 1 molar solution. /preparation of 1M methanoic acid – see appendix, page 26/

Quantities of methanoic acid I decided to take:

1. 2.5 ± 0.1 cm³ of 1M HCOOH per 100cm³ of solution
2. 4.5 ± 0.1 cm³ of 1M HCOOH per 100cm³ of solution
3. 6.5 ± 0.1 cm³ of 1M HCOOH per 100cm³ of solution
4. 8.0 ± 0.1 cm³ of 1M HCOOH per 100cm³ of solution
5. 10.0 ± 0.1 cm³ of 1M HCOOH per 100cm³ of solution

I decided for higher concentrations of the acid than in the case of hydrochloric acid as I noticed that it is hardly possible to dissolve 0.10g of calcium carbonate in 0.5 cm³ of methanoic acid.

The results of titration are following:

Amount of acid [cm ³ ± 0.1]	Volume of titre used [± 0.1 cm ³]		
	1 st trial	2 nd trial	3 rd trial
2.5	1.8	1.7	1.8
4.5	2.0	2.0	2.0
6.5	2.1	2.0	2.0
8.0	2.0	2.0	2.0
10.0	2.0	2.0	2.1

Then I calculated the amount of calcium present in the solution. From the previous part of my experiment I knew that the amount of calcium ions in the solution containing 0.10g of calcium carbonate should equal 0.041g. I calculated the amount of calcium ions in each solution. The results are presented in the table.

For 2.5 cm³ of HCOOH

$$x = \frac{1.8 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.036 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/1.8 \cdot 0.036 = 0.002$$

For 4.5 cm³ of HCOOH

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 6.5 cm³ of HCOOH

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 8.0 cm³ of HCOOH

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 10.0 cm³ of HCOOH

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

The results are presented in the table:

Amount of HCOOH in the solution [cm ³ ± 0.1]	Amount of calcium in the solution [g ± 0.002]	Theoretical amount of calcium in the solution [g]
2.5	0.036	0.040
4.5	0.040	
6.5	0.040	
8.0	0.040	
10.0	0.040	

I noticed that calcium carbonate was not fully dissolved in solution containing 2.5cm³ of 1 M solution of methanoic acid. While using higher amount of acid the salt was dissolved completely and the method of calcium determination for such amount was very precise.

1.6 Investigation on the effect of a weak electrolyte – ethanoic acid (CH₃COOH) on the solubility of calcium carbonate

In the last part of my experiment I checked the influence of a weak electrolyte on the solubility of calcium carbonate. As in both previous examples I used 0.10g of calcium carbonate in 100cm³ solution. I used 1 molar solution of ethanoic acid to dissolve the salt. /preparation of 1M ethanoic acid solution – look appendix, page 26/

Quantities of hydrochloric acid I decided to take:

1. 0.5 ± 0.1 cm³ of 1M CH₃COOH per 100cm³ of solution
2. 1.5 ± 0.1 cm³ of 1M CH₃COOH per 100cm³ of solution
3. 2.5 ± 0.1 cm³ of 1M CH₃COOH per 100cm³ of solution
4. 4.5 ± 0.1 cm³ of 1M CH₃COOH per 100cm³ of solution
5. 6.5 ± 0.1 cm³ of 1M CH₃COOH per 100cm³ of solution

Again I used the amount of the acid the same as in the case of hydrochloric acid as according to the literature calcium carbonate is easily dissolved in this acid.

The results of titration are following:

Amount of acid [cm ³ ± 0.1]	Volume of titre used [± 0.1 cm ³]		
	1 st trial	2 nd trial	3 rd trial
0.5	0.8	0.7	0.8
1.5	1.7	1.8	1.8
2.5	2.0	2.1	1.9
4.5	2.0	2.0	2.0
6.5	2.0	2.0	2.1

Similarly as in the previous parts I could now calculate the amount of calcium ions present in each solution. Again I could compare my results with the theoretical value of calcium ions that should be present in each solution. Then I could notice where calcium carbonate was dissolved completely. The results are as follow:

For 0.5 cm³ of CH₃COOH

$$x = \frac{0.8 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.016 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/0.8 \cdot 0.016 = 0.002$$

For 1.5 cm³ of CH₃COOH

$$x = \frac{1.8 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.036 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/1.8 \cdot 0.036 = 0.002$$

For 2.5 cm³ of CH₃COOH

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 4.5 cm³ of CH₃COOH

$$x = \frac{2.0 \text{ cm}^3 \pm 0.1 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

For 6.5 cm³ of CH₃COOH

$$x = \frac{2.0 \pm 0.1 \text{ cm}^3 \cdot 0.05 \text{ M} \cdot 40.08 \text{ g} \cdot 10}{1000.0 \text{ cm}^3 \cdot 1 \text{ M}} = 0.040 \pm 0.002 \text{ g}$$

Error analysis

$$\Delta x = 0.1/2.0 \cdot 0.040 = 0.002$$

The results are presented in the table below:

Amount of CH ₃ COOH in the solution [cm ³ ± 0.1]	Amount of calcium in the solution [g ± 0.002]	Theoretical amount of calcium In the solution [g]
0.5	0.016	0.040
1.5	0.036	
2.5	0.040	
4.5	0.040	
6.5	0.040	

According to my results, 0.10g of calcium carbonate is the most effectively dissolved when the amount of 1 molar solution of ethanoic acid is 2.5cm³ or more. When the amount of 1 molar solution of ethanoic acid used is smaller than 2.5cm³ then calcium carbonate is not dissolved completely.

2.0 ANALYSIS OF THE OBTAINED RESULTS, CONCLUSION AND EVALUATION

Checking solubility of calcium carbonate according to solubility product

In this part of my experiment I was trying to check the solubility of calcium carbonate due to the solubility product. According to this method 0.007 g of calcium carbonate should dissolve in 1000cm³ of water. I used complexometric titration to check if it really does dissolve. Unfortunately the method I used was not sensitive enough to detect such a small amount of ions so the result was very imprecise. That is why I couldn't use such an amount of calcium carbonate into the further investigation.

Determination of the concentration of calcium carbonate which produces the most accurate results while being dissolved

This part was to find the concentration of calcium carbonate on which the results of solubility are the most precise. According to my data two of proposed concentrations proved to be precise - 0.0907 g of CaCO₃/ 100cm³ and 0.1001 g of CaCO₃/ 100cm³. Though both results fixed within the range of an error the second one was more precise. That is why I decided for this concentration for my further investigation - on the effect of strong, medium and weak electrolyte on the solubility of calcium carbonate.

Investigation on the effect of strong, medium and weak electrolytes on the solubility of calcium carbonate

According to my results all of the electrolytes – strong (hydrochloric acid) medium (methanoic acid) and weak (ethanoic acid) could dissolve calcium carbonate. Taking into consideration ethanoic acid and hydrochloric acid I can say that they both are similarly effective in dissolving the salt. In all of the solutions the amount of calcium carbonate was 0.10g (± 0.01) and all the acids were 1molar solutions. I noticed that the whole amount of calcium carbonate was dissolved in 2.5 cm³ and higher volumes of 1 molar solution of hydrochloric and ethanoic acid but for methanoic acid such amount was too little. The smallest amount of 1 molar solution of methanoic acid that could dissolve 0.10 g of calcium carbonate was 4.5cm³. That would mean that salts of methanoic acid (calcium methanoates) are less soluble in water than calcium salts

of hydrochloric and ethanoic acid. That would mean that the most effective way of dissolving calcium carbonate would be using either a strong electrolyte – meaning hydrochloric acid, or a weak electrolyte – meaning ethanoic acid as the amount that should be used would be much smaller than in the case of methanoic acid.

My investigation was very time consuming and the main weakness in my opinion was the number of solutions I had to prepare. The rest of the experimental part was interesting as I was eager to find out the results.

The method I used i.e. complexometric titration was the good choice. With this method I could easily determine the amount of calcium ions in the solution and I could make a lot of repetitions as one trial of titration do not consume much time. The only problem was to determine the exact point of equivalence as the colour had to change from pink into the violet. The other problem I encountered while doing the experiment was preparation of solutions containing certain amount of calcium carbonate. I had to measure very precise amount of the salt which took some time and I couldn't lose any of it when putting into measuring flask. Firstly I tried to measure it on the filter paper but I noticed that some of the crystals stuck into the surface of the paper. That is why I decided to use transparent foil. I could rinse off all the crystals from the foil into the flask so none of them was lost.

My investigation proved that the choice of agent used for getting rid of fur from household devices is right. Producers usually use ethanoic acid for products for cleaning e.g. kettles as it is not very toxic and one can use only small amount of it to get rid of the fur.

APPENDIX

Preparation of reagents:

1. Murexide and NaCl mixture

Rub 0.1 g of murexide with 20 g of sodium chloride using a mortar. The indicator must be hold in bottle made of dark glass.

2. 2.0 molar NaOH

Solution was prepared from 6.0 molar solution of sodium hydroxide.

Using mixing rule amount of water and 6.0M solution of NaOH that should be mixed to obtain 2.0M solution was calculated.

$$\begin{array}{ccc} 6.0M & \rightarrow & 2cm^3 \\ \searrow & & \nearrow \\ & 2.0M & \\ \nearrow & & \searrow \\ 0M & \rightarrow & 4cm^3 \end{array}$$

I needed about $600cm^3$ of 2.0M solution that is why I had to mix $200 cm^3$ of 6.0M Na OH with $400cm^3$ of water

3. 1.0 molar solution HCl

Solution was prepared from 2.0 molar solution of hydrochloric acid.

Using mixing rule amount of water and 2.0M solution of HCl that should be mixed to obtain 1.0M solution was calculated.

$$\begin{array}{ccc} 2.0M & \rightarrow & 1cm^3 \\ \searrow & & \nearrow \\ & 1.0M & \\ \nearrow & & \searrow \\ 0M & \rightarrow & 1cm^3 \end{array}$$

I needed about $1000cm^3$ of 1.0M HCl that is why I had to mix $500 cm^3$ of 2.0M HCl with $500cm^3$ of water

4. 1.0 molar solution of HCOOH

Solution was prepared from 26.5 molar solution of HCOOH

Using mixing rule amount of water and 26.5 M solution of HCOOH that should be mixed to obtain 1.0M solution was calculated.

$$\begin{array}{ccc} 26.5\text{M} & \rightarrow & 1.0\text{cm}^3 \\ \searrow & & \nearrow \\ & 1.0\text{M} & \\ \nearrow & & \searrow \\ 0\text{M} & \rightarrow & 25.5\text{cm}^3 \end{array}$$

I needed about 250.0 cm³ of 1.0M HCl that is why I had to mix 9.5 cm³ of 26.5M HCl with 242.5 cm³ of water

5. 1.0 molar solution of CH₃COOH

Solution was prepared from 8.0 molar solution of CH₃COOH

Using mixing rule amount of water and 8.0M solution of CH₃COOH that should be mixed to obtain 1.0M solution was calculated.

$$\begin{array}{ccc} 8.0\text{M} & \rightarrow & 1.0\text{cm}^3 \\ \searrow & & \nearrow \\ & 1.0\text{M} & \\ \nearrow & & \searrow \\ 0\text{M} & \rightarrow & 7.0\text{cm}^3 \end{array}$$

I needed about 100.0 cm³ of 1.0M CH₃COOH that is why I had to mix 12.5 cm³ of 8.0M CH₃COOH with 87.5 cm³ of water

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