The analysis of Milk of Magnesia by acid–base titration

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Abstract The physical and chemical properties of Milk of Magnesia provide examples of chemical concepts and opportunities for investigative practical work for secondary chemistry students of all ages. Basing the lesson on the contents of that familiar blue bottle, which is probably lurking somewhere in a cupboard in most households, also showcases the role of chemistry in the formulation and analysis of everyday pharmaceutical products. At key stage 3 (age 14), the formation of an insoluble suspension of magnesium hydroxide from the mixing of two clear solutions of soluble salts provides a very visual representation of a chemical reaction and the concept of chemical change. At GCSE level (age 15–16), direct titration of a sample of this product with dilute hydrochloric acid results in titre values consistent with the amount of magnesium hydroxide stated on the label and produces a most visual endpoint as the cloudy mixture simultaneously clarifies and changes colour. Using a back titration provides opportunities for calculations at A-level (age 16–19) and extends students' understanding of quantitative analytical methods.

It is always pleasing when an everyday household product can take centre stage during a chemistry lesson. Better still when the substance of interest has a strong connection to the local area or region. The analysis of Milk of Magnesia, originally developed by the Irish chemist Sir James Murray (Box 1), provides an excellent example of an everyday neutralisation reaction, and allows students the opportunity to consider the contents of a pharmaceutical product. Both direct and back titrations can be carried out to determine the exact amount of the active substance, magnesium hydroxide, present in the formulation. Getting close to the value stated on the bottle can enhance students' confidence and interest in chemistry and provide a context for a discussion of errors and product sampling.

Solubility and chemical change at key stage 3

The concepts 'solubility' and 'chemical change' feature strongly in the junior chemistry curriculum and can be addressed in this simple demonstration. The addition of a small amount of magnesium hydroxide to a beaker of water and stirring should demonstrate its low solubility in water, in comparison with other ionic substances such as sodium chloride with which students will be familiar. The cloudy suspension already resembles Milk of Magnesia and the class could discuss everyday examples of suspensions and solutions such as milk and dilute cordial. A centrifuge could be used to separate the insoluble solid to the bottom of the tube and leave the clear

Box 1 Milk of Magnesia and Sir James Murray

Born in 1788 in County Derry/Londonderry, Sir James trained as a doctor and, after qualification, he began work as an apothecary and physician at Belfast Dispensary and Fever Hospital (now the site of the Belfast City Hospital).

His career flourished under the patronage of the Marquis of Donegall, who owned Belfast Castle. During this time, he experimented with electrical apparatus and, in about 1809, he developed and marketed Murray's Fluid Magnesia. It was sold as a palatable laxative and a remedy for acidities, indigestion, heartburn, and gout. His most famous discovery has reportedly been put to other more unorthodox uses including mouth ulcer treatment, skin toning face masks and even helping a young actor to whiten his hair for a play.

Based on Garvin and O'Rawe (1993)

liquid above it. The low solubility of magnesium hydroxide can be further demonstrated by adding a small beaker containing approximately 100 cm^3 of 0.1 mol dm^{-3} sodium hydroxide to a larger beaker containing approximately 200 cm^3 of 0.1 mol dm^{-3} magnesium sulfate solution to form a cloudy white precipitate of magnesium hydroxide. This reaction (Figure 1) provides an example of a chemical change as it produces new substances, and a simple diagram can be used to represent the reaction and support the idea of the individual ions combining to form the new compounds magnesium hydroxide and sodium sulfate.

$$MgSO_4 + 2NaOH \rightarrow Mg(OH)_2 + Na_2SO_4 \qquad (1)$$

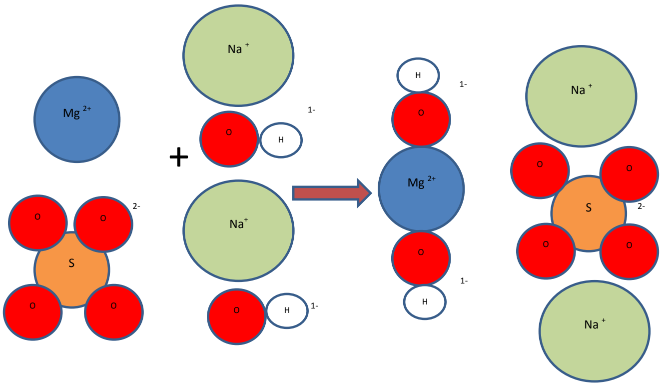

Figure 1 Magnesium sulfate and sodium hydroxide react to form magnesium hydroxide and sodium sulfate

Neutralisation

A sample of Milk of Magnesia can be used as an effective demonstration of neutralisation for the topic of acids and bases at key stage 3 (age 14). When dilute hydrochloric acid is slowly added to a mixture of Milk of Magnesia and water, students can see that the suspension becomes less dense as the magnesium hydroxide reacts with the HCl to form a salt that dissolves, until finally a clear solution is obtained. A sample of 5 cm^3 of Milk of Magnesia mixed in a beaker with approximately 20 cm^3 of water (to make it easier to see) will require about 14 cm^3 of 1.0 mol dm^{-3} HCl for complete reaction (Figure 2).

$$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O \qquad (2)$$

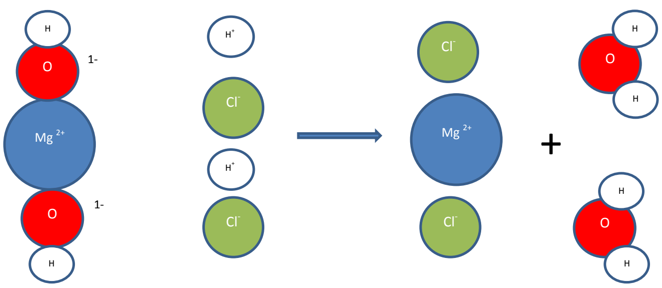

Figure 2 Magnesium hydroxide and hydrochloric acid react to form magnesium chloride and water

With reference to the word equation, students could be asked to decide which of the compounds that appear in the word equation are present in the beaker at each of the following stages:

- after a small portion of acid has been added but not enough to remove all the cloudiness;
- after just enouch acid has been added to remove all of the cloudiness;
- after extra (the word 'excess' can be introduced later) acid has been added beyond the point where all the cloudiness has been removed.

This should support students' understanding of the idea that neutralisation is only complete when the amount of acid added is equivalent to the amount of magnesium hydroxide present, and thus lay the foundations for carrying out titrations later in their studies.

The concept of neutralisation can be made more visual by adding a few drops of universal indicator solution to the suspension of Milk of Magnesia and repeating the gradual addition of 1.0 mol dm^{-3} HCl. The change in colour from blue through green to red 'indicates' when all the magnesium hydroxide has reacted and the solution has become neutral. Addition of more acid makes the solution acidic. Litmus can also be used or, keeping with the theme of everyday science, red cabbage indicator. (This indicator can be produced by heating chopped red cabbage leaves in water for five minutes, filtering and using the solution for testing.)

By firstly demonstrating this neutralisation without an indicator, students are required to focus more closely on the turbidity of the mixture, and this highlights the importance of close observation in chemistry. It may also help to avoid the possible misconception that neutralisation itself must involve an indicator.

Key stage 4

At key stage 4 (age 14–16) a sample of Milk of Magnesia may be directly titrated with 1.0 mol dm^{-3} hydrochloric acid using methyl orange as an indicator (Box 2). This activity could be used as part of a 'value for money' investigation, with comparison to other antacid products or, as in this case, to compare the mass of magnesium hydroxide present to the amount stated on the package.

Students should enjoy this aesthetically pleasing titration (Figure 3). During the early stages, the mixture takes on a yellow appearance not dissimilar to very thin scrambled egg. As the addition of acid is continued, the mixture becomes much thinner but is still clearly a suspension. Just before the endpoint, the mixture is quite transparent before turning to an orange solution at the equivalence point. Titre values are usually within 0.1 cm^3 of each other but there may

Box 2 Direct titration method

Apparatus

Burette, conical flask, pipette (5 cm^3), wash bottle, beaker (250 cm^3), small funnel

Reagents

Milk of Magnesia, 1.0 mol dm^{-3} HCl, methyl orange indicator solution

Health and safety

1.0 mol dm^{-3} hydrochloric acid is LOW HAZARD (CLEAPPS). Safety glasses and protective gloves should be used. Ensure that eyewash facilities are available.

Methyl orange is LOW RISK (CLEAPPS).

Method

1. Using a 5 cm^3 pipette, transfer 5 cm^3 of Milk of Magnesia into a conical flask. Using a wash bottle, wash any of the suspension adhering to the inside of the pipette into the flask. Add distilled water so that the total volume of water added is approximately 25 cm^3.
2. Add two drops of methyl orange indicator solution and titrate the initially yellow mixture with 1.0 mol dm^{-3} HCl. During the course of the titration, wash down the sides of the conical flask with distilled water.
3. As the titration proceeds, the mixture should become more transparent. The endpoint occurs when the colour of the solution changes to orange.
4. Repeat the procedure and calculate the mean titre value.

Results and calculation

Mean titre = 14.10 cm^3

Number of moles of 1.0 mol dm^{-3} HCl required
$= 14.10/1000 \times 1.0 = 1.41 \times 10^{-2}$

$Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

Number of moles of $Mg(OH)_2$ present = 0.705×10^{-2}

Mass of $Mg(OH)_2$ present $= 0.705 \times 10^{-2} \times 58.33$ (RFM)
$= 0.411\ g = 411\ mg$

The value stated on the bottle is 415 mg of $Mg(OH)_2$ per 5 cm^3 of sample.

Figure 3 The sample solution (from left to right) at the beginning, the middle and just at the endpoint of the titration

be quite a range of values depending on how effectively the students have washed the product from the pipette.

AS and A2 level

While Milk of Magnesia can be directly titrated with accuracy, an alternative method of analysis is via a back titration (Box 3). The CCEA (2016: 42) specification for A2 level requires students to '*use a back titration to determine the percentage of an active ingredient in an indigestion remedy*' within the unit 'Chemistry in Medicine'. The specification also includes back titrations for determining the purity of Group 2 oxides and carbonates. Although back titrations are not included in other specifications, the experiment includes many teaching points relevant to the A-level curriculum.

Teaching points

Carrying out both the direct and the back titration within the same practical session provides students with the opportunity to experience the colour change of methyl orange when titrating an alkaline solution with an acid (yellow to orange) and when titrating an acid

Box 3 Back titration method

Apparatus

Burette, 250 cm^3 volumetric flask, conical flasks, pipette (5 cm^3), pipette (25 cm^3), wash bottle, beaker (250 cm^3), small funnel

Reagents

Milk of Magnesia, 1.0 mol dm^{-3} HCl, 0.1 mol dm^{-3} NaOH, methyl orange indicator solution

Health and safety

1.0 mol dm^{-3} hydrochloric acid is LOW HAZARD (CLEAPPS). Safety glasses and protective gloves should be used. Ensure that eyewash facilities are available.

Methyl orange is LOW RISK (CLEAPPS).

0.1 mol dm^{-3} NaOH is LOW HAZARD (CLEAPPS).

Method

Sample analysis

1 Using a pipette (5 cm^3), transfer 5 cm^3 of Milk of Magnesia into a 250 cm^3 volumetric flask.

2 Using a wash bottle, wash any of the suspension adhering to the inside of the pipette into the volumetric flask and continue to do so until the effluent appears clear and free of suspension.

3 Using a pipette (25 cm^3), add 25 cm^3 of 1.0 mol dm^{-3} HCl to the volumetric flask. The suspension will react and dissolve with the acid to produce a clear solution. Make to volume with distilled water and invert several times to ensure thorough mixing.

4 Using a pipette (25 cm^3), transfer 25 cm^3 of the solution into each of two conical flasks.

5 Add two drops of methyl orange indicator solution to each conical flask.

6 Titrate one of the flasks at a time with 0.1 mol dm^{-3} NaOH solution, using the second flask as a colour reference, until the solution first becomes yellow.

7 Repeat the titration to obtain at least three titre values and calculate the mean titre value.

Blank analysis

1 Using a pipette (25 cm^3), add 25 cm^3 of 1.0 mol dm^{-3} HCl to a volumetric flask. Make to volume with distilled water and invert several times to ensure thorough mixing.

2 Using a pipette (25 cm^3), transfer 25 cm^3 of this solution into each of two conical flasks.

3 Add two drops of methyl orange indicator solution to each conical flask.

4 Titrate one of the flasks at a time with 0.1 mol dm^{-3} NaOH solution, using the second flask as a colour reference, until the solution first becomes yellow.

5 Repeat the titration to obtain at least three titre values and calculate the mean titre value.

Results and calculation

Blank analysis

Mean titre = 24.5 cm^3

Number of moles of 0.1 mol dm^{-3} NaOH required

$= 24.5/1000 \times 0.1 = 24.5 \times 10^{-4}$

Therefore 24.5×10^{-4} moles of HCl are present in the conical flask

Therefore 24.5×10^{-3} moles of HCl are present in the blank sample

Sample analysis

Mean titre = 9.75 cm^3

Number of moles of 0.1 mol dm^{-3} NaOH required

$= 9.75/1000 \times 0.1 = 9.75 \times 10^{-4}$

Therefore 9.75×10^{-4} moles of HCl are present in the conical flask

Therefore 9.75×10^{-3} moles of HCl are present in the sample of Milk of Magnesia

Therefore the number of moles of HCl neutralised by reaction with $Mg(OH)_2$

$= 24.5 \times 10^{-3} - 9.75 \times 10^{-3} = 14.75 \times 10^{-3}$

Number of moles of $Mg(OH)_2$ present in sample

$= 7.375 \times 10^{-3}$

Mass of $Mg(OH)_2$ present in sample (5 cm^3)

$= 7.375 \times 10^{-3} \times 58.33 = 0.430$ g (430 mg)

The amount stated on the label = 415 mg

% error = $(430 - 415)/415 \times 100 = 3.6\%$

The error in the result will arise from the combined error in the analysis of the blank and the analysis of the sample. The labelled value itself will also be prone to error and the actual value may vary from batch to batch.

solution with an alkali (red to yellow), and to see how the colour relates to the pH of the solution. The calculation based on the data from the back titration provides good practice in calculating the number of moles of each species and converting moles into grams and milligrams.

The direct titration of Milk of Magnesia provides an example of Le Chatelier's Principle in action. Magnesium hydroxide is sufficiently soluble to produce an alkaline solution in water, though the bulk of the compound is suspended (undissolved) in the conical flask (see equation 3). As the added acid neutralises the hydroxide ions present in the solution, more solid magnesium hydroxide dissolves to replace them. These hydroxide ions are in turn neutralised by the addition of more hydrogen ions. This continues until all of the magnesium hydroxide is neutralised.

$$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2OH^-(aq)$$
$$\downarrow 2HCl(aq)$$
$$MgCl_2(aq) + 2H_2O(l) \qquad (3)$$

The equilibrium between solid and dissolved magnesium hydroxide is brought to the right-hand side by the reaction of hydroxide ions with acid.

The use of Milk of Magnesia to relieve indigestion and constipation provides opportunities to discuss the human digestive system, particularly the presence of hydrochloric acid in the stomach and the role of osmosis in the large intestine. Milk of Magnesia is an example of a type of 'hyperosmotic laxative' and works by drawing water from nearby tissue by osmosis into the large intestine. This softens and moistens the stool and therefore helps increase bowel activity. You should expect to have a bowel movement within six hours of taking Milk of Magnesia.

This study of a household pharmaceutical product not only showcases the chemistry behind its medicinal properties and method of analysis, but also allows for a consideration of how its formulation must ensure that it looks, smells and sounds as attractive as possible. The class discussion could consider why peppermint oil is added, why the product is called 'Milk of Magnesia' and why it is packaged in an opaque rather than a transparent bottle. As well as extending students' knowledge and understanding of chemistry, this activity may help develop their appreciation of the role of marketing in the formulation and sale of pharmaceutical products.

Acknowledgements

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References

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