Extension 26: The History of Milk of Magnesia

1. Prerequisites

The key ideas required to understand this section are:

Topic	Book page
Antacids	450
British Pharmacopeia	450

2. Background

Charles Henry Phillips (1820 – 1882) was an English pharmacist who founded the Phillips Camphor and Wax Company in Stamford, Conneticut. Here, in 1873, he concocted and patented a suspension of magnesium hydroxide which was sold under the brand name *Phillips' Milk of Magnesia* for use as a medicine.

After Phillips' death in 1882, his sons continued to run the company. Milk of Magnesia is now marketed by Pharma Omega; its most familiar form is as a liquid suspensions. It is an antacid – it neutralizes stomach acid and so relieves indigestion and heartburn. It is also effective as a laxative. Milk of Magnesia has an osmotic effect on the body and it draws water into the affected area. For example, when taken as a laxative to treat constipation, Milk of Magnesia draws water into the intestines. The water entering the intestinal tract causes movement in the intestines and safely relieves constipation. In the early years, adverts for the product also claimed it had other remarkable properties, such as alleviating some of the effects of smoking and improving the complexion. The former claim is no longer made, although some claim it reduces the oiliness of the skin.

Phillips' Milk of Magnesia is subject to the regulation of the UK Medicines and Healthcare Products Regulatory Agency (MHRA). In 2012, its sale was suspended because of concerns about the level of sulfate in the mixture. The European Pharmacopoeia, which set legal and scientific standards for the quality control of medicines, stipulates the maximum limits of sulfate that can be present in magnesium hydroxide suspensions. Too high a level of sulfate in milk of magnesia could form an unwanted side-product, magnesium sulfate. This sulfate can cause complications in the central nervous system; confusion, depression, slower respiratory rates and decreased blood pressure, as examples. Subsequently, the sulfate content was lowered.

3. Preparation

Instructions for the preparation of "magnesium hydroxide mixture" are given in the British Pharmacopeia (BP). Essentially, the mixture is prepared by reacting magnesium sulfate with sodium hydroxide.

You can investigate this preparation by duplicating some of the steps in a teaching laboratory (**but do not attempt to ingest any of the products, remember the manufacture of medicines is carried out under very strict conditions and by highly trained chemists!**)

First, magnesium sulfate solution is prepared from the reaction of dilute sulfuric acid with magnesium oxide:

$$MgO(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) \tag{1}$$
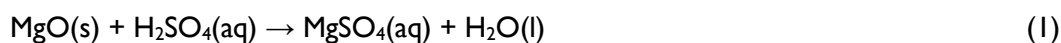
The resulting solution is heated until the water has evaporated, leaving $MgSO_4.7H_2O$.

A sample of the hydrated sulfate is then dissolved in water and sodium hydroxide solution added with stirring. A white precipitate of magnesium hydroxide precipitates out in suspension.which will settle out over time.

$$MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq) \tag{2}$$

Magnesium hydroxide forms a gelatinous precipitate that is very hard to filter off – it would clog most filters using standard laboratory apparatus, so more sophisticated techniques are employed such as drum filtration.

4. Assay

The product can be assayed for content and its purity determined. Magnesium hydroxide has very little solubility in water, so it cannot be titrated directly; a technique known as back titration is used.

The hydroxide will, however, react with dilute sulfuric acid in a neutralisation reaction:

$$Mg(OH)_2(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + 2H_2O(l) \tag{3}$$
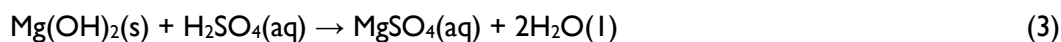
A known mass of a sample of the hydroxide is dissolved in a known amount of sulfuric acid, which is in excess, and the amount of unreacted acid that remains after the reaction is titrated against sodium hydroxide solution using methyl orange as indicator. In this way, the amount of acid that reacted with magnesium hydroxide can be calculated, leading to determination of the purity of the sample.

Revision questions

1. Can you write an ionic equation for equation (2)?

2. There were concerns about the level of sulphate in Milk of Magnesia.

    (i) Where do you think this originated from?
    (ii) How would you attempt to minimise this during the preparation of magnesium hydrozide?
    (iii) How could you confirm (qualitatively) that you had reduced the sulphate concentration?

3. The product from the preparation of magnesium hydroxide (0.20 g) is dissolved in 20 cm³ of deionized water and 20 cm³ of sulfuric acid of concentration 1.0 mol dm⁻³.

The resulting mixture is titrated with sodium hydroxide solution (1.0 mol dm⁻³), using methyl orange as indicator. It was found that 34 cm³ of sodium hydroxide were added until the endpoint, when the indicator turned yellow. What was the percentage purity of the sample?

Answers

1. $Mg^{2+}(aq) + 2OH^-(aq) \rightarrow Mg(OH)_2(s)$
(sodium and sulfate ions are spectator ions)
2.
    (i) During the preparation, aqueous sulfate is formed. The precipitate will be "wet" with this solution.

(ii) The precipitate should be washed well, several times, with deionised water after the initial filtration.

(iii) Test the filtrate from successive washings for the presence of sulfate – see p. 90. As the precipitate gets less dense, the sulfate ion concentration is reduced.

3. You need the equation for the neutralisation:

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$
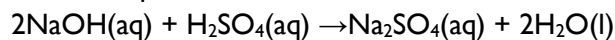
Amount in mol of sodium hydroxide used = volume × concentration of sodium hydroxide

= 34/1000 × 1.0 = 0.034 mol

This corresponds(from the equation) to 0.034/2 mol = 0.017 mol sulfuric acid.

Amount of acid originally added in mol = volume x concentration = 20/1000 × 1.0 = 0.020 mol

So: 0.020 – 0.017 mol of acid were used up reacting with magnesium hydroxide which is 0.003 mol.

This also corresponds to a complete reaction with 0.003 mol magnesium hydroxide (equation 3)

Mass of $Mg(OH)_2$ is 0.003 × 58 = 0.17 g

So, percentage of $Mg(OH)_2$ in original sample = $\frac{\text{mass of Mg(OH)}_2}{\text{mass of sample}} \times 100 = \frac{0.17}{0.20} = 85\%$

(Note that these are results you might expect in your lab. In an industrial assay, you would expect to get a purity of > 95%.)