The Reaction of Magnesium Metal with Water: A Kinetic study

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Introduction

The alkali and alkaline earth metals are the elements with the most pronounced metallic character. Like the other main group elements, this increases as one goes down the periodic table. All react vigorously (in some cases extremely vigorously) with acids. Most of them also react rapidly with water (cesium explosively). A number of reports exist in the literature showing the reactions of alkaline or alkaline earth metals with water.[1-3]

There are, however, two exceptions to the above generalisation. Beryllium does not react with water at all (even at red heat), and magnesium reacts with cold water only very slowly.[4] Such behavior of these two metals differs markedly both from that of their neighbors to the left (lithium and sodium), and from the one below (calcium)[3]. The reason for their low reactivity is the existing oxide protective covering in the case of Be, or in the case of Mg, the low solubility of the hydroxide.

A few years ago we reported the results of a preliminary study of the reaction of magnesium with both water and a solution of sodium chloride[5] in order to confirm that this latter explanation was correct. The measurement of the rates of reaction of magnesium in water and in potassium chloride solution can be easily followed by measuring the rate of evolution of hydrogen gas using simple equipment that should be available in most school laboratories, and we believe it would form a useful class demonstration. The details are given below.

Chemicals and equipment

Magnesium turnings, distilled water, dilute sulfuric acid (about 0.5 mol/L), a saturated solution of potassium chloride, a trough filled with distilled water, 2 test tubes, 2 one-hole corks, 2 doubly bent tubes, 2 graduated cylinders, 3 stands, 4 clamps with appropriate screw-holders.

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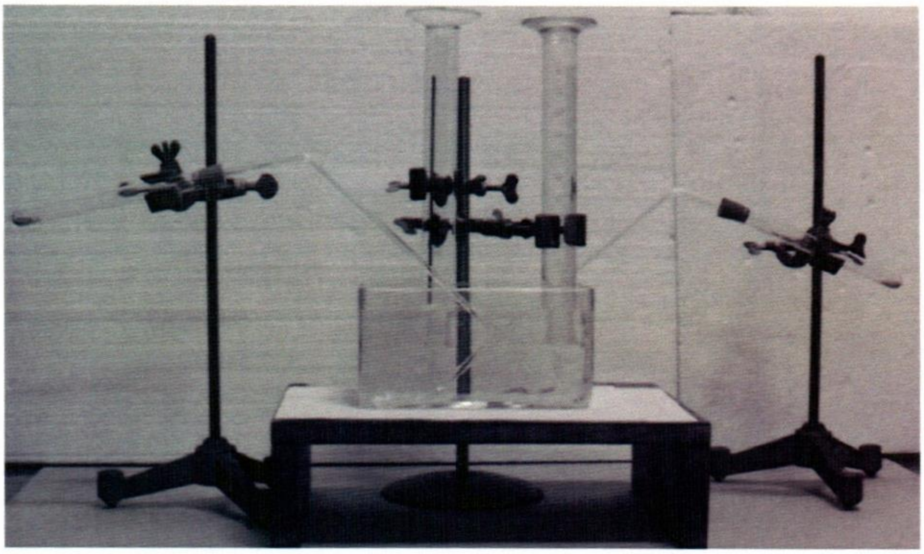

Fig. 1. A suggested experimental setup for the two experiments.

Method

The experiment is set up as shown above in Figure 1. To one of the test tubes is added 15 mL of distilled water and to the other 15 mL of saturated potassium chloride solution. To each is added exactly 0.5 g of dry magnesium turnings (previously cleaned by washing first with dilute sulfuric acid, then with water). The two systems are sealed and the hydrogen generated in each is collected in inverted measuring cylinders as shown above. The reaction is slow but steady and can be monitored over a period of about 2 weeks. The results of a typical run are presented in Figures. 2 and 3. Inspection of these shows that hydrogen generation is about 2 orders of magnitude faster in the case of KCl solution.

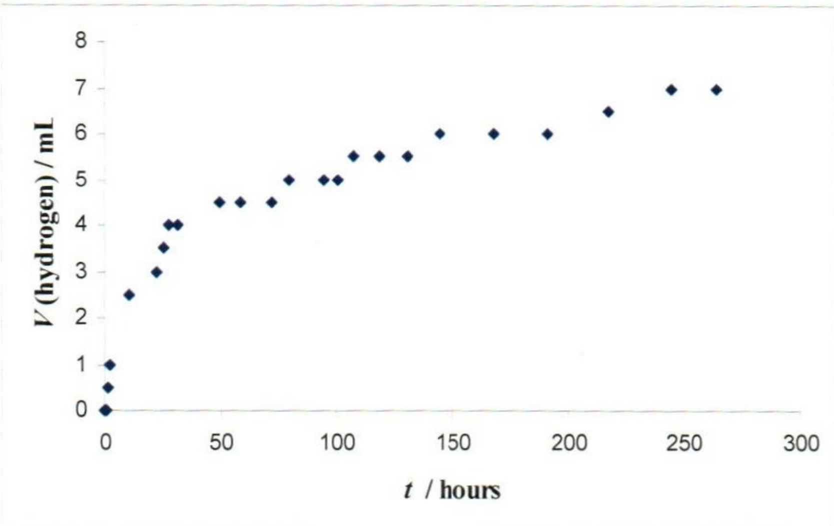

Fig. 2. Hydrogen generated from magnesium and distilled water

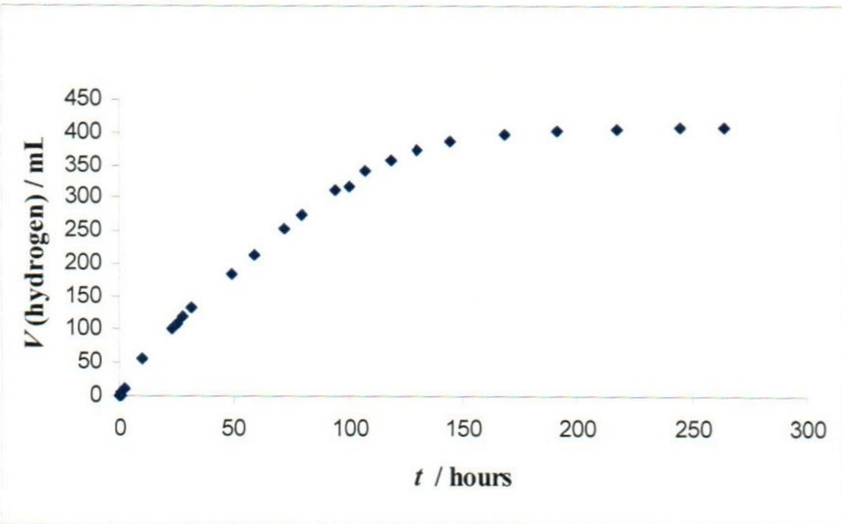

Fig. **3.** Hydrogen generated from magnesium and saturated KCl(aq)

Discussion

Comparison of the results illustrated in Figures 2 and 3 shows that the rate of evolution of hydrogen is much greater in the KCl solution. It is known that the rate of electrochemical corrosion is much higher in the presence of electrolytes, and aqueous KCl is an electrolyte. Perhaps the rate is higher for the same reasons. One logical explanation is that the product of our reaction is $Mg(OH)_2$ which is insoluble in water and this precipitates and adheres to the surface of the magnesium, slowing the reaction in water by limiting access. The salt effect[6] should result in the magnesium hydroxide being more soluble in the KCl solution, and this would leave the magnesium more open to attack. Alternatively it may be that in the strong KCl solution some of magnesium may form a species such as Mg(OH)Cl, and that this is more soluble than $Mg(OH)_2$.

Both of these explanations (and others) are feasible. We hope to study hydrogen generation in the presence of a number of different salts, including $MgCl_2$(aq) in order to investigate further.

References

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