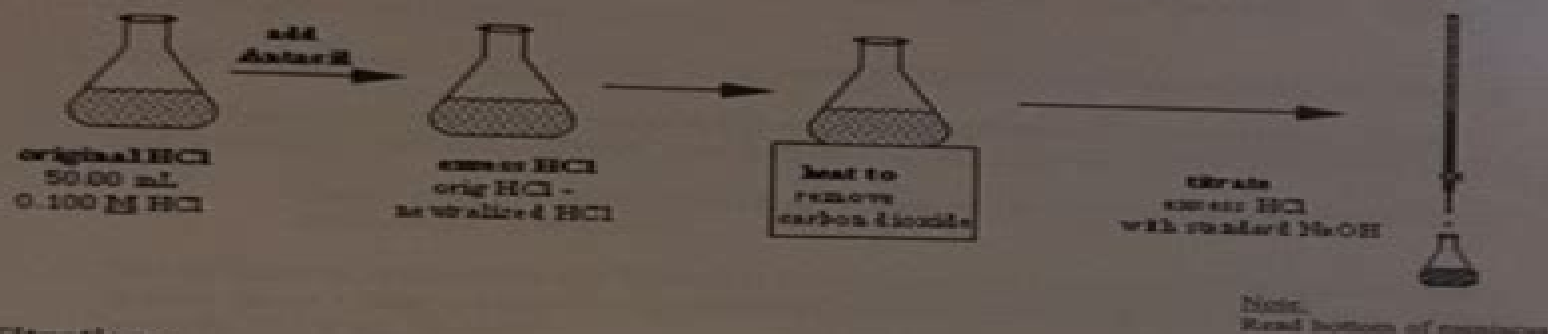


In this experiment, a portion of an antacid tablet will be dissolved in 50.00 mL of standard 0.100 M HCl solution simulating stomach acid. The antacid will neutralize some of the known amount of acid present in the sample of acid. The sample will then be heated gently, and the millimoles of acid that was *not* neutralized will be determined by titrating the sample against a standard NaOH solution. Finally, the per gram neutralizing ability of each antacid can be calculated using the original mass of antacid added to the simulated stomach acid.

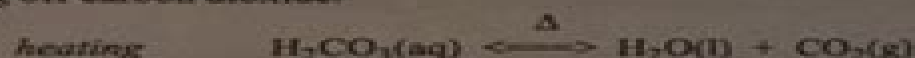


The Titration:

We will only work with the carbonate and bicarbonate antacids in this lab session. This means that a $\text{HCO}_3^-/\text{CO}_3^{2-}$ buffer can form when the antacid is added to the hydrochloric acid.



If this buffer is present during titration of excess acid, the endpoint will not be as sharp as desired. To circumvent this problem, you will heat the sample gently before titration. This will remove H_2CO_3 by boiling off carbon dioxide.



The removal of carbon dioxide as it bubbles from the solution will force the equilibrium to the right according to Le Chatelier's Principle until the buffer equilibrium is finally destroyed.

When the sample is subsequently titrated with standard NaOH, we get a classic acid-base reaction. Please remember that the HCl we are titrating is the excess HCl that was *not* neutralized by the antacid. This means that the more effective the antacid, the less HCl will be left to titrate and the smaller the titration volume will be.



Since there are a number of ions which can hydrolyze water present in your sample, at the endpoint of your titration, you will not have a neutral solution. Consequently, instead of phenolphthalein, we have selected an indicator which works better under these (acidic) conditions. The indicator you will use is methyl red, which is red in acid and yellow in base.

Effective Commercial Antacids Experiment

**Frederick A. Bettelheim, Joseph Marvin
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