

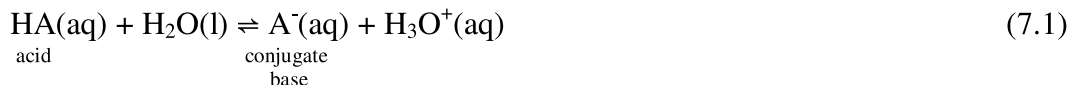
PRACTICAL NUMBER 7

ACID STRENGTH, ACID CONCENTRATION AND pH

INTRODUCTION

In this experiment you will determine the pH values of a number of different acids of various strengths over a range of concentrations. You will attempt to relate these measurements to the following approximate expressions relating pH to the concentration of the acid (its *analytical concentration*, C_a) and the pK_a value for the acid.

For a Brønsted-Lowry acid in aqueous solution we have:



For a strong acid the equilibrium lies well over to the right (virtually no intact HA molecules remain) and so the concentration of H_3O^+ is equal to the *analytical* concentration of the acid, the concentration which would be determined by titration, or in other words the concentration assuming no dissociation of the acid. This is represented by C_a . Hence:

$$\begin{aligned} [\text{H}_3\text{O}^+] &= C_a \\ \therefore \lg[\text{H}_3\text{O}^+] &= \lg C_a \\ \therefore -\lg[\text{H}_3\text{O}^+] &= -\lg C_a \\ \therefore \text{pH} &= -\lg C_a \end{aligned} \quad (7.2)$$

For a weak acid the equilibrium lies well to the left (virtually all HA molecules remain intact). We must apply the equilibrium expression:

$$\frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]} = K_a$$

But $[\text{HA}] \simeq C_a$ and $[\text{A}^-] \simeq [\text{H}_3\text{O}^+]$

$$\begin{aligned} \therefore \frac{[\text{H}_3\text{O}^+]^2}{C_a} &= K_a \\ \therefore [\text{H}_3\text{O}^+] &= K_a^{1/2} C_a^{1/2} \text{ and taking logs of both sides :} \\ \lg[\text{H}_3\text{O}^+] &= \lg(K_a^{1/2} C_a^{1/2}) = \lg K_a^{1/2} + \lg C_a^{1/2} = \frac{1}{2} \lg K_a + \frac{1}{2} \lg C_a \\ \therefore -\lg[\text{H}_3\text{O}^+] &= -\frac{1}{2} \lg K_a - \frac{1}{2} \lg C_a \\ \therefore \text{pH} &= \frac{1}{2} pK_a - \frac{1}{2} \lg C_a \end{aligned} \quad (7.3)$$

METHOD

You will prepare several series of dilutions from stock solutions of various 1.0 M acids and then measure the approximate pH of each of the solutions as follows:

- 1) You will need a source of distilled water of consistent pH. Check the pH of the water from your wash bottle. It should be around 5 or 6. (Why not 7?) Compare this with the pH of the water from the aspirator. If there is any significant difference consult your supervisor. Record the pH of the water used.

- 2) Take a rack of 6 test tubes and place sufficient 1.0 M HCl(aq) to make the measurement (see step 6 below) in the first.
- 3) Add a further 10 cm³ of 1.0 M HCl(aq) to a 100 cm³ measuring cylinder and then add distilled water to the 100 cm³ mark. Stopper the measuring cylinder and then shake the contents. Pour enough of the solution into the second test tube to make a second measurement.
- 4) Discard most of the rest of the solution from the measuring cylinder until just 10 cm³ of the solution remain and then top up the cylinder once again with distilled water to the 100 cm³ mark. Put some of this solution in the third tube for the third measurement.
- 5) Proceed in a similar manner until all the tubes contain hydrochloric acid solution of progressively decreasing concentration.
- 6) Record the concentration (C_a) of each solution in the results table.
- 7) Measure the pH values of the solutions with the indicator papers provided, or (if available) with a pH meter. If indicator papers are to be used, a white tile with dimples in it may be used. Place a piece of indicator paper a dimple and add few drops of solution to it. Be sure to use a clean teat pipette and ***start with the most dilute solution***. This is because the dilute solutions are the most easily affected by contamination. For the same reason you have to be particularly careful about cleanliness with these solutions.
- 8) Various types of pH paper are available. Each type covers a different range of pH and has its own colour chart. A given type cannot measure pH values outside its range. The pH value may be found by comparing the colour of the paper with the colour chart provided. If it is found that using, for example, pH paper covering a range of 1 to 4, a result at one end of the range is obtained (for example 4), then this value may not be accurate and should be checked using another pH paper covering the pH value better (eg 1 to 10 paper). In addition, if solutions of indicator are provided, it may be possible to test with these too.
- 9) Record your measured pH values in the results table.
- 10) Repeat the above determinations using 1.0 M solutions of ethanoic acid, ammonium chloride, and aluminium chloride (or magnesium chloride). Make a careful note on your paper if other solutions have been substituted.
- 11) Complete the table by calculating $-\lg C_a$ for each dilution.

RESULTS

pH of water used in the experiment: _____ = _____

	<i>Tube 1</i>	<i>Tube 2</i>	<i>Tube 3</i>	<i>Tube 4</i>	<i>Tube 5</i>	<i>Tube 6</i>
<i>Conc. of acid: $C_a/\text{mol dm}^{-3}$</i>						
<i>$-\text{Lg}C_a$</i>						
<i>pH of hydrochloric acid solution</i>						
<i>pH of ethanoic acid solution</i>						
<i>pH of ammonium chloride solution</i>						
<i>pH of aluminium (or magnesium) chloride solution</i>						

PROBLEMS

- 1) Plot graphs of pH (vertical axis) against $-\text{Lg}C_a$ (horizontal axis) for all the solutions between a single pair of axes (i.e. all on one graph). Use a different style (or colour) of point and line for each of the substances and include a key with your graph. Also label each of the lines with the name of the substance used. Note that the graphs should be straight lines with different slopes and intercepts, but that if the water itself is at all acidic or basic, then the pH values for the more dilute solutions will probably be in error.
- 2) Probably the pH of the water used (if it is distilled water) will show that it is slightly acidic. Why should we expect this?
- 3) Is it possible that the pH of any of your solutions could be higher than that of the water used in making them up? Explain.
- 4) Explain why there are differences in pH between different acids having the same concentration. Are your results broadly what you would expect to observe? Explain.
- 5) Could a solution of a weak acid have a lower pH than a solution of a strong acid? (Look at your results!) Explain.
- 6) Which of the acids is/are weak? Calculate the pK_a values of the *weak* acids from the fact that the intercept on the vertical axis is equal to $\frac{1}{2}\text{pK}_a$.
- 7) Explain, with an equation, a reaction that makes solutions of aluminium chloride (or magnesium chloride) acidic. (See the introduction for practicals 9 and 10 under *aquocations*.) On the basis of your answer suggest why it is impossible to obtain pure anhydrous aluminium (or magnesium) chloride by heating the hydrated salt to drive off the water?

- 8) Write equilibrium equations showing how and why solutions of ammonium chloride and ethanoic acid are acidic.
- 9) In question (6) you were asked to calculate the pK_a of the weak acids. Explain why the intercept on the pH axis gives the pK_a value for the acid. Why does this not apply to **strong** acids?
- 10) Explain the difference between the **strength** of an acid and the **concentration** of an acid.
- 11) Calculate the degree of dissociation (α) for 1 M and 10^{-5} M solutions of ethanoic acid. What do your answers illustrate?