

## THE THEORY OF pH

The pH scale was suggested by Sorensen almost a century ago. We can think of it as a sort of shorthand for specifying the concentration of protons ( $H^+$  ions) in (usually aqueous) solution. Sorensen defined  $pH = -\log[H^+]$ . (Remember that [ ] means concentration or molarity.) In water, pH values from 0 (very acidic) to 14 (very basic) are possible. Thus pH 7, being right in the middle of the range, is neutral, that is, neither acidic nor basic. There are two tricky things about this scale: it's nonlinear and it seems "to go the wrong way" (high  $H^+$  corresponds to low pH). Both of these odd features stem from the fact that pH involves logarithms. First of all, a negative logarithm indicates an inversion or reciprocal of a number; that is why the scale seems "to go the wrong way." Also because of logarithms, the numbers seem much smaller:  $\log(10^7) = 7$ ,  $\log(10^{-2}) = -2$ ,  $\log(4.87 \times 10^3) = 3.68$ . So a change the pH from 7 to 5, increases the proton concentration by 100 times!

Why does pH warrant its own rather weird scale? Mostly its due to the central role that water plays in chemistry, biochemistry, and biology. In as much as water is the solvent in which most chemical reactions occur, it's very important to consider ionization of water. In addition to unionized  $H_2O$ ,  $OH^-$  and  $H^+$  can be present. The concentration of  $OH^-$  ions can be expressed by pOH which is defined as  $pOH = -\log [OH^-]$ . There is a simple relationship between these ions:  $pH + pOH = 14$ . The pH specifies the ionic state of the most important solvent. Another reason for the widespread use of the pH scale is the ease of constructing electrochemical devices ("pH meters") to measure it.

The pH is a scale of numbers that simplifies manipulation, calculation, and discussions involving the wide range of hydrogen ions in a solution. Determination of pH is c a determination of emf (voltage) difference in a pH cell. pH is dependent upon temperature; a temperature compensator varies the instrument definition of a pH unit from 54.20 mV to 66.10 mV at 60°C.

A pH meter consists of an electrode to sense pH, a reference electrode to complete the circuit while not affecting the pH measurement, an electronics module to link the electrode to the readout, and a display unit which could be a meter or a computer screen. Some pH meters also have a third electrode to sense the temperature in order to compensate for thermal effects. Often the electrodes are packaged into a "combination electrode" which is more convenient to use. The most frequent cause of pH meter malfunction is poorly-maintained electrodes.

In modern chemistry the "p" ("negative log of") notation has been extended to other species. Thus a chemist may use  $pS^-$ ,  $pNO_2^-$ ,  $pNa^+$ , and even  $p(\text{urea})$ . So-called ion-selective electrodes have been developed to measure all of those and many more.



## pH TITRATION OF OXALIC ACID

### Introduction:

The end-point of an acid-base titration is the point at which the acid and base are stoichiometrically equivalent. This point may be found by adding an indicator which changes color at the approximate pH of the equivalence point. A more elegant method uses a pH meter to monitor pH during the titration. A graph of pH versus titrant will exhibit an inflection point where the reactants are equivalent. In this experiment, a solution of oxalic acid will be titrated with NaOH to the end-point and beyond. A graph of pH versus volume of NaOH will be plotted and the concentration of the NaOH solution will be calculated.

### Purpose:

The purpose of this experiment is to titrate an NaOH solution to find its concentration.

### Equipment / Materials:

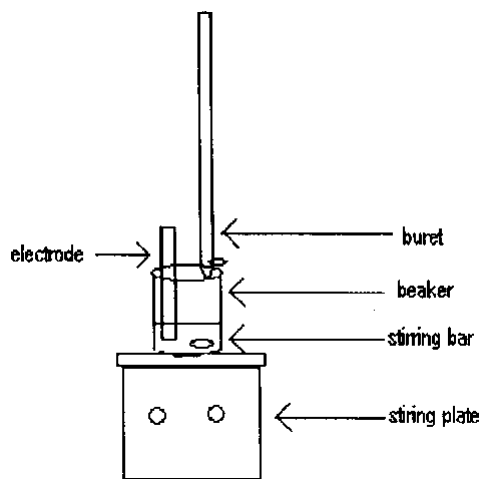
pH meter	oxalic acid
2-250 mL beakers	analytic balance
buret and clamp	scoopula
wash bottle	magnetic stirrer and stir bar
Kimwipes	dropper
Unknown NaOH	

### Safety:

- Always wear goggles in the lab.
- Oxalic acid is highly poisonous.

Procedure:

1. Plug in the pH meter and allow it to warm up for about 10 minutes.
2. The temperature knob should be set between 20 and 25°C.
3. Remove the cap from the electrode, and rinse the electrode with DI water. Blot the end with a Kimwipe.
4. Place the electrode in the pH 7 buffer, turn the knob to pH, and adjust the pH to 7.00 with the standardization knob.
5. Place the instrument on standby. Rinse and blot the electrode.
6. Place the electrode in pH 4.00 buffer, turn the knob to pH, and adjust the pH to 4.00 with the slope knob.
7. Place the instrument on standby. Rinse and blot the electrode.
8. Rinse the buret with - 5 mL of Unknown NaOH. Fill the buret past the 0.00 mL line and drain to the 0.00 nil, line. If the 0.00 mL line is passed, use a dropper to get the meniscus to the 0.00 ml, line.



11. Take an initial reading of the pH and the volume of NaOH to two decimal places.
12. Begin adding the NaOH slowly, stopping at frequent intervals to record the volume and the pH. Collect data at intervals of 0.2 pH units or 1 mL. Be particularly alert after pH 5.0 as the pH rises quite fast at this point.
13. Continue the titration until 50 ml, of the base has been added or until the pH reaches 12, whichever comes first.



Calculations:

Plot a titration curve of pH (y-axis) versus mL of NaOH (x-axis). Indicate the end-point and from it determine the concentration of the NaOH solution.

Questions:

1. Why must the pH meter be standardized?
2. Looking at the titration curve, is oxalic acid a monoprotic or polyprotic acid? Explain.
3. Explain why a polyprotic acid, when titrated, may only show one inflection point.

*Science in Motion*

*Materials List*

**Lab: pH Titration of Oxalic Acid**

Number of Lab Groups Prepared: \_\_\_\_\_

Equipment per lab group	Delivered	Returned
pH meter		
2-250 mL beakers		
Buret and clam		
Wash bottle		
Analytic balance		
Scoopula		
Magnetic stirrer and stir bar		
Consumables	Delivered	
Kimwipes		
Unknown NaOH		
Oxalic acid		
Dropper		