

pH Primer – Use of pH Meter and Indicators

This lab is just to start to get familiar with the pH meters and different indicators. We will be doing 3 different major parts.

1. pH of Salt Solutions
2. Determine the pH of a 1×10^{-10} M HCl solution.
3. Calculate the percent by volume of acetic acid in vinegar.

You will be using pH meters which are ~\$500. The probes alone are ~\$300 so we must be careful. I will talk about the use and care of the pH meters. Please keep them clean and do not abuse them.

pH of Salt Solutions

You will follow the procedure given out in class. Make sure you write a chemical equation that will explain the observed pH on the basis of the pH measurement. Also using your textbook select an indicator that would allow you approximate the pH of each solution.

Determine the pH of a 1×10^{-10} M HCl solution

Using volumetric flasks and pipets you and your group will come up with a way to make a 1×10^{-10} M HCl solution from 1 M HCl. Before proceeding you must get my approval of the procedure. Once you have made you 1×10^{-10} M HCl solution, determine the pH of the solution using the pH meter.

Calculate the percent by volume of acetic acid in vinegar

Using pH meter determine the pH of standard household vinegar. Knowing that the density of vinegar is 1.0492 g/mL at 20 °C and that the K_a of acetic acid is 1.78×10^{-5} , calculate the percent by volume of acetic acid in vinegar.

This lab should be recorded in your notebooks and the carbon copies turned in as your lab report. Make sure that you have all of your data, calculations, and written procedure in your lab notebook.

pH of Salt Solutions

Salts that can be considered to have been formed from the complete neutralization of strong acids with strong bases—such as NaCl (which can be considered to have been produced by the neutralization of HCl with NaOH)—ionize completely in solution. Such strong acid/strong base salts do not react with water molecules to any appreciable degree (do not undergo hydrolysis) when they are dissolved in water. Solutions of such salts are neutral and have pH = 7. Other examples of such salts are KBr (from HBr and KOH) and NaNO₃ (from HNO₃ and NaOH).

However, when salts formed by the neutralization of a weak acid or base are dissolved in water, these salts furnish ions that tend to react to some extent with water, producing molecules of the weak acid or base and releasing some H⁺ or OH⁻ ion to the solution. Solutions of such salts will *not* be neutral but, rather, will be acidic or basic themselves.

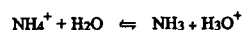
Consider the weak acid HA. If the sodium salt of this acid, Na⁺A⁻, is dissolved in water, the A⁻ ions released to the solution will react with water molecules to some extent



The solution of the salt will be *basic* because hydroxide ion has been released by the reaction of the A⁻ ions with water. For example, a solution of sodium acetate (a salt of the weak acid acetic acid) is basic because of reaction of the acetate ions with water molecules, releasing hydroxide ions:



Conversely, solutions of salts of weak bases (such as NH₄⁺Cl⁻, derived from the weak base NH₃) will be *acidic*, because of reaction of the ions of the salt with water. For example,



Many salts of transition metal ions are acidic. A solution of CuSO₄ or FeCl₃ will typically be of pH 5 or lower. The salts are completely ionized in solution. The acidity comes from the fact that the metal cation is hydrated. For example, Cu(H₂O)₆²⁺ better represents the state of a copper(II) ion in aqueous solution. The large positive charge on the metal cation attracts electrons from the O-H bonds in the water molecules, thereby weakening the bonds and producing some H⁺ ions in the solution:



C. pH of Salt Solutions

Note: If the number of pH meters in the laboratory is limited, your instructor may ask you to perform this qualitative portion of the experiment using *universal indicator*, rather than the pH meter. Universal indicator is a dye that exhibits different colors over the entire pH range. Universal indicator does not permit as sensitive a measurement of pH as would be possible with a pH meter or with the method using several indicators as discussed in Part

A of this experiment. For a quick “ballpark” estimate of pH, however, universal indicator can be very useful.

In this case, obtain 10 drops of each salt solution, and test for pH with one drop of universal indicator. Use the color chart provided with the indicator to estimate the pH of each salt solution. Write an equation accounting for the pH observed.

If sufficient pH meters are available in the lab, first calibrate the meter using a pH reference buffered solution, and then obtain 20–25 mL (in a small beaker) of one of the 0.1 M salt solutions listed in the table that follows. Rinse the pH electrode with distilled water and immerse the electrode in the salt solution.

Salt	Formula
sodium chloride	NaCl
potassium bromide	KBr
sodium nitrate	NaNO ₃
potassium sulfate	K ₂ SO ₄
sodium acetate	NaC ₂ H ₃ O ₂
ammonium chloride	NH ₄ Cl
potassium carbonate	K ₂ CO ₃
ammonium sulfate	(NH ₄) ₂ SO ₄
copper(II) sulfate	CuSO ₄
iron(III) chloride	FeCl ₃

Allow the electrode to stand in the solution for 2–3 minutes so that it may come to equilibrium with the solution; then record the pH of the salt solution.

Rinse the pH electrode with distilled water, and determine the pH of each of the other salt solutions listed in the table.

On the basis of the pH measured for each salt solution, write a chemical equation that will explain the observed pH.