Science Education Collection

Using a pH Meter

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Overview

Source: Laboratory of Dr. Zhongqi He - United States Department of Agriculture

Acids and bases are substances capable of donating protons (H^+) and hydroxide ions (OH^-), respectively. They are two extremes that describe chemicals. Mixing acids and bases can cancel out or neutralize their extreme effects. A substance that is neither acidic nor basic is neutral. The values of proton concentration ($[H^+]$) for most solutions are inconveniently small and difficult to compare so that a more practical quantity, pH, has been introduced. pH was originally defined as the decimal logarithm of the reciprocal of the molar concentration of protons $(\text{pH}=-\log[\text{H}^+]=\log(^1/_{\text{H}^+})$, but was updated to the decimal logarithm of the reciprocal of the hydrogen ion activity $(\text{pH}=-\log a_{\text{H}^+}=(^1/_{a_{H^+}}))$. The former definition is now occasionally expressed as p[H]. The difference between p[H] and pH is quite small. It has been stated that pH = p[H] + 0.04. It is common practice to use the term 'pH' for both types of measurements.

The pH scale typically ranges from 0 to 14. For a 1 M solution of a strong acid, pH=0 and for a 1 M solution of a strong base, pH=14. Thus, measured pH values will lie mostly in the range 0 to 14, though values outside that range are entirely possible. Pure water is neutral with pH=7. A pH less than 7 is acidic, and a pH greater than 7 is basic. As the pH scale is logarithmic, pH is a dimensionless quantity. Each whole pH value below 7 is 10x more acidic than the next integer. For example, a pH of 4 is 10x more acidic than a pH of 5 and 100x (10 x 10) more acidic than a pH of 6. The same holds true for pH values above 7, each of which is 10x more basic (or alkaline) than the next lower whole value. For example, a pH of 10 is 10x more basic than a pH of 9.

Principles

The pH of a solution may be accurately and easily determined by electrochemical measurements with a device known as a pH meter with a pH (proton)-sensitive electrode (usually glass) and a reference electrode (usually silver chloride or calomel). Ideally, the electrode potential, *E*, for the proton can be written as

$$E = E_0 + 2.303\ (^{RT}/_F)\log a_{\text{H}^+} = E_0 - 2.303\ (^{RT}/_F)\ \text{pH}$$

where *E* is a measured potential, E_0 is the standard electrode potential at a_{H^+}= 1 mol/L, *R* is the gas constant, *T* is the temperature in kelvin, *F* is the Faraday constant.

The pH electrode uses a specially formulated, pH-sensitive glass in contact with the solution, which develops the potential (*E*) proportional to the pH of the solution. The reference electrode is designed to maintain a constant potential at any given temperature, and serves to complete the pH measuring circuit within the solution. It provides a known reference potential for the pH electrode. The difference in the potentials of the pH and reference electrodes provides a millivolt (mV) signal proportional to pH. In practice, a combined glass electrode has a built-in reference electrode. It is calibrated against buffer solutions of known hydrogen ion activity. Most pH sensors are designed to produce a 0-mV signal at 7.0 pH, with a (theoretically ideal) slope, or sensitivity, of -59.16 mV / pH at 25 °C. Two or more buffer solutions are used in order to accommodate the fact that the "slope" may differ slightly from ideal. Commercial standard buffer solutions usually come with information on the pH value at 25 °C and a correction factor to be applied for other temperatures.

Procedure

1. pH Calibration

1. Turn the meter's power on by pressing the "power" button.
2. Attach the automatic temperature compensation (ATC) probe if it is available and/or is not with the electrode.
3. Check that the measurement mode is pH. If not, press the "MODE" button until "pH" mode appears on the LCD display.
4. Consult the quick reference guide at the bottom of the meter or nearby for help if needed.
5. Always use fresh, unused, unexpired pH buffers for calibration. Buffers should be at the same temperature as the testing solutions.
6. Rinse the pH electrode with distilled water and then with the buffer being used for calibration (*i.e.*, pH 7.00).
7. Dip the pH electrode into a neutral pH buffer (*i.e.*, pH 7.00). Stir the buffer with a magnetic bar (at a moderate rate for ~30 s) for best results.
8. Press the "CAL/MEAS" (calibration [or Standardization]/easurement) button to select the 'calibration (standardization)' function. Set the buffer pH value on the meter display to 7.00.
9. When the "reading" is stable, press the "ENTER" button to accept. The primary reading will flash briefly before the secondary display begins scrolling through the remaining available buffers.
10. Rinse the pH electrode with distilled water and then with the buffer to be used for next calibration (*i.e.*, pH 4.01).
11. Dip the pH electrode into the next buffer of pH 4.01. The meter display should be locked on the buffer value.
12. When the "reading" is stable, press "ENTER" to accept. The primary reading will flash briefly, then display the percent efficiency (slope) before the secondary display begins scrolling through the remaining the available buffers.
13. Repeat steps 2.2-2.7 to calibrate the pH 10.01 buffer.

14. The meter will automatically return to measurement mode upon the successful completion of the 3-point calibration.
Notes: (1) The standard buffers of pH 4.01 and 10.01 may be replaced with other appropriate buffers per the testing samples' pH range. (2) For a single (neutral)-, or 2-point calibration, press the "CAL/MEAS" button to return to the measurements after completion of the calibration. (3) Calibration with more than 3 points may be used for more precise measurements. (4) It is recommended to perform the calibration at the beginning of each day. For very precise work the pH meter should be calibrated before each measurement. (5) Manually adjust the pH values of the buffers if the temperature differs from the standard room temperature and no ATC probe is attached.

2. pH Measurements

1. Confirm that the meter is on the pH measurement mode.
2. Thoroughly rinse the pH electrode between measurements with distilled water to prevent carryover contamination of the tested solutions. Gently blot the electrode on a laboratory cleaning tissue to remove the excess rinse water. Do not rub the bulb since this can cause a static charge buildup. Alternatively, rinse the electrode with the testing solution.
3. Dip the pH electrode into a testing solution or suspension. Stir the solution with a magnetic bar (~30 s) with the same stirring rate as for calibration for best results.
4. The pH is completed when the pH reading is stable.
5. If needed and available, press the "HOLD" button to freeze the measured reading. Press again to resume live reading.
6. Record the pH value (and temperature if needed) by writing down or pressing the "MEMORY" button (if applicable) to store the value into memory.
7. Repeat steps 3.2-3.6 for multiple measurements.
8. Thoroughly rinse and store the electrode in storage solution once all measurements are completed.
Notes: (1) The pH probe response time in each buffer should be no longer than 60 s, but may be longer for some testing solutions/ slurries. (2) The electrode probe should be cleaned using pH-electrode cleaning solution once a month, or whenever it is dirty. A 0.1 M HCl solution can be used for general cleaning. Diluted liquid detergent and household laundry bleach may be used for cleaning grease and bacterial contaminations. However, to avoid unexpected problems, the best practice is to always refer to the electrode manufacturer recommendations. (3) The pH electrode bulb should be moist at all times. Keep it in the electrode storage solution that comes with the electrode. Use pH 4 buffer solution if no storage solution is available. Use pH 7 buffer solution for a short time if neither are available.

Results

Figure 1 shows the pH of agricultural soils impacted by cropping management and groundwater irrigation. These soil samples were collected from 5 potato fields under different cropping rotation practices with or without groundwater irrigation. The lowest pH is observed in Field 4 soils in both rainfed and groundwater irrigated series. Groundwater irrigation consistently increased soil pH in all 5 fields. The pH information is essential for recommendation of liming the potato fields appropriately to promote optimal growth.

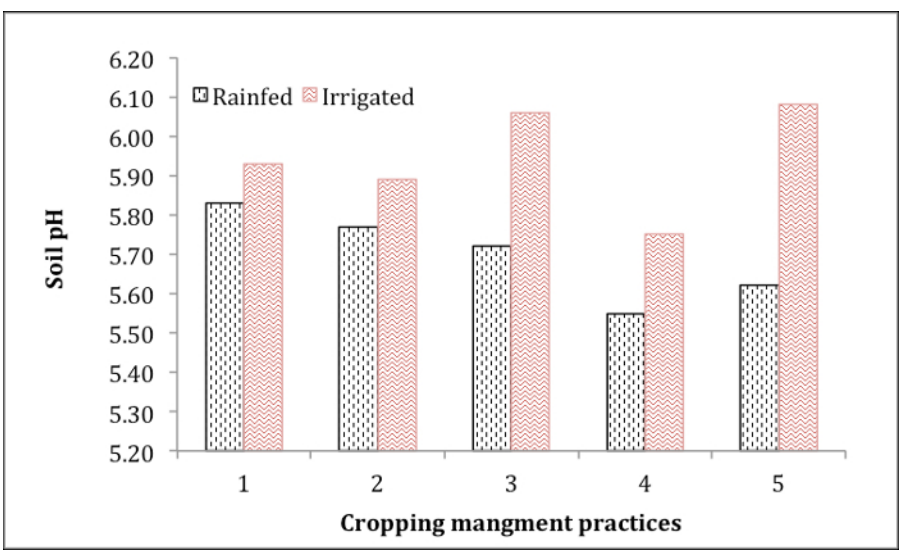

Figure 1. Soil pH of potato fields under different cropping management practices with or without underwater irrigation.

Applications and Summary

pH is one of the most commonly measured chemical parameters of aqueous solutions. It is a critical parameter in water and wastewater treatment for municipal and industrial applications, chemical production, agriculture research, and production. It is also critical in environmental monitoring, chemical and life sciences research, biochemical and pharmaceutical research, electronics production, and many more applications. **Figure 2** lists pH values of some common substances.

Pure water is neutral, with a pH of 7.00. When chemicals are mixed with water, the mixture can become either acidic or basic. Vinegar and lemon juice are acidic substances, while laundry detergents and ammonia are basic. Chemicals that are very basic or very acidic are considered "reactive." These chemicals can cause severe burns. Automobile battery acid is an acidic chemical that is reactive. Automobile batteries contain a stronger form of one the acids found in acid rain. Household drain cleaners often contain lye, a very alkaline chemical that is also reactive.

In living systems, the pH of different cellular compartments, body fluids, and organs is usually tightly regulated in a process called acid-base homeostasis. The pH of blood is usually slightly basic with a value of pH 7.365. This value is often referred to as physiological pH in biology and medicine. Plaque can create a local acidic environment that can result in tooth decay by demineralization. Enzymes and other proteins have an optimum pH range and can become inactivated or denatured outside this range.

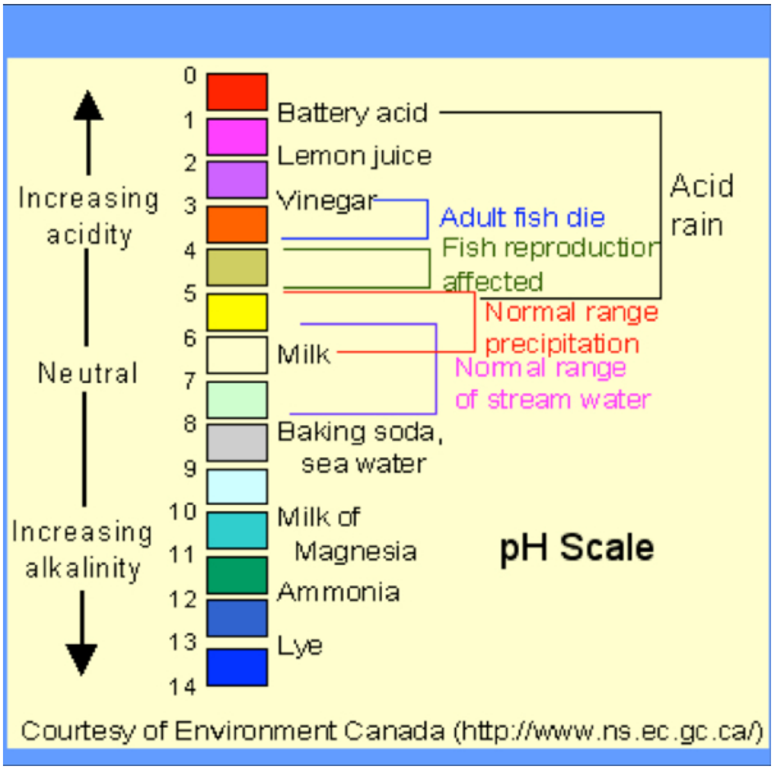

Figure 2. The pH scale and the pH values of some common items.