An electrolyte is a substance which dissociates free ions when dissolved (aq) to produce an electrically conductive medium[i], such as acids and bases which can both be either strong or weak electrolytes. An acid is a compound which donates protons, opposed to a proton acceptor; a base. An acid produces hydrogen ions when dissolved in water, while a base produces hydroxide ions instead. For example the following neutralization reaction: *H*Cl (Acid) + Na*OH* (Base)→ H_2O + NaCl The Acid donates the H+ (hydrogen ion), and the base's OH- (hydroxide ion) accepts it to form a neutral solution. Both hydroxide and hydronium ions may be formed from water. A hydroxide ion is formed when a water molecule *loses* an H+ atom and becomes negatively charged (OH-), while a hydronium ion *gains* an H+ atom and becomes positively charged (H_3O^+).

The pH scale is a spectrum for expressing the hydrogen-ion concentration in moles per liter, in which all acids and bases fall upon; ranging from 0-14, the most acidic being 0 and the most basic being 14. The pH value itself is a negative logarithm of the hydrogen-ion concentration; calculated through the following formula: pH = - Log[H+]. Therefore, it measures the concentration of H+, the larger the exponent and the more basic, and the smaller the exponent and the more acidic. The pOH scale is very similar (pOH = -Log[OH-]), however it measures the concentration of OH- ions instead, so logically, these two scales are complementary, and together always add up to 14. Neutral solutions lie right in the middle of the pH scale, with a [H+] concentration of approximately 1.0×10^{-7} mol/L, while acidic solutions have a pH, or an [H+] concentration greater than 1.0×10^{-7} mol/L. Basic, or Alkaline solutions have a greater [OH-] concentration, so therefore the [H+] concentration is less than 1.0×10^{-7} mol/L.

Svente August Arrhenius was an individual who was born in Sweden in 1859 and received the Nobel Prize in 1903 for chemistry. Circa that period, Arrhenius was focused on *Investigating the Galvanic Conductivity of Electrolytes*, his thesis in 1884. From his results he concluded that electrolytes when dissolved in water, become to varying degrees dissociated into positively and negatively charged ions; the degree of this depending on the concentration of the solution (molarity), being more developed the greater the dilution.[ii] Basically, he discovered that in an aqueous solution, acids produce H+ ions, combining with a corresponding negative ion, and bases produce OH- ions. The work done by Arrhenius has come to help define many important terms and set the foundation for further exploration.

pH is extremely relevant and important in the context of the human body. During exercise for example, hemoglobin in the blood provides for oxygen, which is used in the breakdown of chemical energy in glucose into mechanical energy. Through this process, H+ molecules and CO_2 molecules are produced, increasing the H+ concentration therefore lowering the pH of the bloodstream. if the blood pH gets too low from the exemplary pH of 7.4, a condition known as acidosis results, and if the pH continues to drop below 6.8, death may occur.[iii] However, homeostasis is also self-maintained to some extent; if there is an excess of [H+] in the bloodstream, an excess of these ions will enter the cell, in an attempt to form equilibrium. Intracellular pH is relevant here; in all cells there is a substance, which has its own pH. These cells will help absorb the excess H+ concentration, in turn lowering their pH. As levels of CO_2 and H+ exceed the capacity of hemoglobin, according to Le Chậtelier's principle, equilibrium shifts which results in an overall shift in pH level, and acidosis. Hyperventilation and increased breathing help by adding O_2 and removing CO_2 which in turn helps lower the blood's pH, this process is referred to as respiratory acidosis. Homeostasis is further aided by organs which serve as buffers, or buffer substances themselves.

A buffer is a substance containing a weak acid and its conjugate base or a weak base and its conjugate acid.[iv] Therefore, buffers are able to maintain a stable pH with additions of small amounts of acids or bases.[v] Some organs within the human body nearly serve as buffers, such as the kidney; for example it absorbs H+ atoms and helps maintain a constant pH: this process is known as metabolic acidosis, and the failure to do so often results in acidosis. Natural buffers include sodium bicarbonate ($NaHCO_3$), which is abundant in the ocean and in many other ecosystems like in the Adirondacks; it helps balance and reduce the change in acidity due to acid rain.[vi] In most natural waters, buffers for a pH of 0-6 include CO_2 while in pH of 6-10, HCO_3- is far more predominant; and furthermore, in pH of 10-14 CO_3^{2-} is evident.[vii] This brings us to stomach pH.

During digestion, food entering the stomach catalyzes the excretion of hydrochloric acid, and the pH of the contents drop from 4 to 2; this means a hundred fold increase in the [H+] concentration, in order to break down the contents.[viii] The high acidity in the stomach environment has a very corrosive effect, which digests the food, however does not take as much a toll on the stomach. The lining of the stomach is resistant to high acid concentration, and in addition secrete mucous which provides more protection. This process thoroughly breaks down food which is in the stomach, however cells in the esophagus and small intestine are sensitive to the acidity, and this gives rise to stomach ulcers and acid reflux disease. This excess in [H+] production often results in pepsin disease, with the common term heartburn resulting. What antacid tablets do is neutralize this [H+] concentration by adding a weak base; of which the most common are hydroxides, bicarbonates and carbonates.[ix] The following reaction represents what would happen if calcium carbonate in an antacid tablet was added to an acidic [HCl] solution in the stomach:

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Hence, the neutralization of the stomach acid and the rise in pH; the products include carbon dioxide gas and water.

If I came across an unknown substance in the chemistry lab, which was revealed to be an acid, and I was asked to find the strength of the acid I would go about doing a number of things. I would test the acid with litmus paper and compare the color change to the pH reading. If this method for some reason failed, I would go about it another way. First, I would measure the amount of liquid solution there was, and then I would try to neutralize the solution. I would refer to the following formula to help calculate the strength of the acid: M × mL (of base) = M × mL (of acid). By observing the amount of base added before the solution was neutralized, as well as the molarity of the base I was using, a value may be calculated for what the mL acid × the molarity of that acid. I would divide this value by the mL of acid to find the molarity. Next, once the molarity was found I would either take the [H+] concentration (if it were a strong acid) or I would multiply that value by the Ka and then find the square root (if it were a weak acid). After finding the [H+] concentration I would either take the negative log if necessary, to find the pH of the substance, or I would judge the strength of the acid by its [H+] concentration.

For the Experiment portion, my group and I used a cabbage juice indicator solution to test the pH of many household substances. The purpose of this experiment was to test the pH of many household substances, and some non-household substances. To begin the experiment we added approximately 150 mL of tap water to a blender, with approximately six or seven cabbage leaves. We blended the components together, and extracted the substance it produced. We then strained out the excess cabbage leaf shreds and chunks, leaving only the liquid extract (pigment) as well as the water. This pigment would provide us with a dependable indicator as to the pH of the substances we would add. We poured between 15-20 mL's of this indicator solution into eight test tubes, observing that the initial color of this pigment was a deep purple. This color represented the color of an approximate neutral solution. Though the tap water was not distilled and may not have had an exact pH of 7, it was an approximation for a neutral value. We then commenced to fill each of these test tubes with household substances, hoping that the indicator solution would change color in order to reveal the approximate pH of the substances. The following table reveals our results from the changes in color of the substances.

Household substance	Initial Color of indicator	Final Color of indicator (color change if any)	Estimated pH based on color change
1) Lemon Juice	Deep purple	Changed to light opaque pink	Acidic 1-3
2)Ammonia	Deep purple	Changed to dark green with purple on bottom	Basic 13-14
3)Tonic Water	Deep purple	Changed to lighter see-through pink	Acidic 3-5
4)Hydrogen peroxide	Deep purple	Remains purple however lighter, begins to foam	More neutral 6-8
5) Liquid Detergent (water based)	Deep purple	Remains same color, tiny bit lighter	More neutral 7-8
6) Vinegar	Deep purple	Changes lighter and acquires slight pink tint	Acidic 4-6
7) Sodium Bicarbonate dissolved in water	Deep purple	Changes to true blue	Basic 8-10
8) Super Arm + Hammer washing Soda	Deep purple	Changes to dark green	Basic 10-13

We also tested the color change of two control solutions; to see what the base of comparison would be. The table below shows the results of this test:

Substance	Initial Color of indicator	Final Color of indicator (color change if any)
Sulfuric Acid (strong acid)	Deep purple	Red
Sodium Hydroxide (strong base)	Deep purple	Green- eventually turns to yellow

The results are rather self explanatory, producing a spectrum of pH values from household substances, ranging from a dark green which was very basic, to a more blue color which was a weaker base, to a neutral color of purple. On the acidic side of the spectrum, the colors ranged from red, which signified a strong acid, to a lighter pink which was a weak acid, to a transformation of pink into the neutral purple. We concluded that substances found in a household provide for a wide array of different pH values, from extremely acidic to extremely basic.

The pictures to follow show our experiment. The first picture shows all of the test tubes after the different substances had been added, and the color change has happened. The second picture shows the indicator solution mixed with the near neutral water. The third shows the piece of cabbage after we took some of the leaves.

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[ii] Svente Arrhenius, Biography. © 2005 The Nobel Foundation. Last modified April 14, 2005 <http://nobelprize.org/chemistry/laureates/1903/arrhenius-bio.html>

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[v] see source 1

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