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| ***� Why Acid Rain Makes Plants Go Brrr�*** ��� An Introduction |

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|  | **I. Introduction to Acid Rain.� Pure water has a pH of 7.0, which is neutral on the pH scale; however, natural, unpolluted rainwater actually has a pH of about 5.6, which is more acidic than pure water (Casiday).� The natural acidity of rainwater comes from the presence of three substances (CO2, NO, and SO2) found in the troposphere, the lowest portion of the atmosphere (Casiday). As is seen in the table below, carbon dioxide (CO2), formed in the decomposition of organic matter and through cellular respiration, is present in the greatest concentration and therefore contributes the most to the natural acidity of rainwater (Casiday).**       |  |  | | --- | --- | | **Gas** | **Concentration** | | **Carbon dioxide**  **CO2** | **355 ppm** | | **Nitric oxide**  **NO** | **0.01 ppm** | | **Sulfur dioxide**  **SO2** | **0-0.01 ppm** |   **Item I-I**    Carbon dioxide forms carbonic acid through a reaction with water.� Carbonic acid can then proceed to break down and release a hydrogen ion (H+), also called a proton, and a hydrogen carbonate ion (HCO3-).� The ability of H2CO3 to release a proton (H+) is what classifies it as an acid, thus lowering the pH of a solution (Casiday).� The reactions just described are shown again below.� (Also, for an introduction to the pH scale, proceed to the next section of the introduction.)   |  | | --- | |  |   **Item I-II**   |  | | --- | |  |   **Item I-III**  Nitric oxide (NO) also contributes to the natural acidity of rainwater and is formed during lightning storms by the reaction of nitrogen and oxygen.� Then, NO gains an oxygen atom from the air and becomes nitrogen dioxide (NO2) which in turn reacts with water to give nitric acid (HNO3) (Casiday).� This acid breaks down in water to yield hydrogen ions and nitrate ions (NO3-) in a reaction similar to the breakdown of carbonic acid shown above in Items I-II and I-III, also lowering the pH of the solution (Casiday).� The reactions just outlined are shown formulaically below.� Also, the very small amounts of sulfur dioxide (SO2) that contribute to the natural acidity of rainwater are formed chemically through volcanic activity and the decay of organic matter (Shaw).� The chemical reactions that relate to sulfur dioxide are shown later in items I-VIII, I-IX, and I-X.��     |  | | --- | |  |   **Item II-IV**   |  | | --- | |  |   **Item III-V**   |  | | --- | |  |   **Item I-VI**    Up to this point, our discussion has focused entirely on the causes of the natural acidity of rainwater.� Unfortunately, these are not the only origins of the acidity of acid deposition.� Acid deposition is a term used to describe any deposit that has a pH below 5.6 (Phamornsuwana).� This reference is used because precipitation in the pH range of 5.6-7.0 can attribute its acidity to the natural causes previously discussed.� However, when the pH of precipitation falls below the 5.6 reference mark, we must begin to attribute its acidity to unnatural, human sources.� We will first discuss the causes of unnatural acid deposition and focus on the effects later in the introduction.  Excessively acidic deposition is caused by the same chemical constituents examined above.� The difference, however, is that the quantities of sulfur dioxide and nitrogen oxides far exceed those typically encountered under natural circumstances.� More than 90% of the sulfur and 95% of the nitrogen emissions occurring in eastern North America are of human origin (Environment).�� The detrimental effects of the increased levels of these chemicals are caused not only through rainfall, but also by snow, fog, or dry particles of sulfate and nitrate salts (Wilson).� In some areas of the United States, the pH of rainwater can be as low as 3.0, almost 100 times more acidic than normal rainwater.� (((5.6-3.0)/3.0) x100=86.667)� Some typical concentrations of acid-producing chemicals, after the inclusion of human contributions, are shown below (Casiday).� Note the increases in value from Item I-I.�     |  |  | | --- | --- | | **Gas** | **Concentration** | | **Nitric oxide**  **NO** | **0.2 ppm** | | **Sulfur dioxide**  **SO2** | **0.1 - 2.0 ppm** |   **Item I-VII**  Roughly a quarter of the acidity of rain is accounted for by nitric acid (HNO3) (Casiday). In addition to the natural reactions that result in the small quantities of nitric acid in rainwater, combustion involved in transportation and energy production emits immense quantities of nitric oxide gas (Environment). Motor vehicles contribute about 45% of the nitrogen oxides emitted in the United States, approximately the same amount as from coal burning utilities (Carolina 1).� This gas then forms nitric acid via the equations labeled Item I-V and I-VI.� The other 75% of the acidity of rain is caused by the sulfuric acid (H2SO4) found in the affected rainwater.� Although sulfuric acid may be produced naturally in small quantities from organic decay and volcanic activity as was discussed above, the majority is produced by human industry, particularly the combustion of sulfur-containing fossil fuels by power plants and the base-metal smelting industry (Environment). When these fossil fuels are burned, the sulfur contained in them reacts with oxygen from the air to form sulfur dioxide (SO2) (Wilson). This compound then reacts with atmospheric oxygen and water to form sulfuric acid (H2SO4) (Wilson).� Combustion of fossil fuels accounts for approximately 80% of the total atmospheric SO2 in the United States (Casiday).   |  | | --- | |  |   **Item I-VIII**  Sulfuric acid breaks down readily in water, to give an H+ ion and an HSO4- ion.� The HSO4- ion may further breakdown to give H+ and SO42-.� Consequently, the presence of H2SO4 causes the concentration of H+ ions to increase dramatically, and so the pH of the rainwater drops relatively quickly when mixed with sulfuric acid.�   |  | | --- | |  |   **Item I-IX**   |  | | --- | |  |   **Item I-X**    ����������� Now that we have discussed both the natural and anthropogenic causes of acid deposition, we will take a brief look at its effects.� Since this is such an overwhelmingly large area, we will focus solely upon the known effects of acid deposition on plant life. Acid harms tree leaves, decreasing a tree�s ability to photosynthesize.� Furthermore, it damages the tree�s bark, leaving it vulnerable to insects and disease.� Acid also leaches nutrients, such as calcium and magnesium, from the soil, depriving already weakened trees of essential minerals. Finally, acid releases toxic aluminum ions from the soil, which can damage plant roots (Carolina 2).� Plant life, as does all life, thrives on the balance that nature has established through millions of years of evolution and natural selection.� Each individual and each population fulfills a unique purpose and occupies a unique niche in the biosphere.� Each individual and each population has established through trial and error a perfectly balanced, dynamic network of supply and demand.�� Acid deposition, however, poses an imminent danger to the balance of this network.� Natural acidity in precipitation is dealt with by the natural buffering capacity of soils and lakes.� When the pH of acid deposition oversteps its natural bounds and exceeds the level at which it can be neutralized, buffering alkaline chemicals will begin to be depleted (Environment).� They are used at a faster pace than that at which they can be restored.� When this occurs, the first step to destroying the balance of the biosphere has been taken.� The pH of soils cannot be neutralized quickly enough through natural processes and can drop well below safe levels.� This dangerous and detrimental effect of acid deposition will be the basis for our experiment.� We will examine one particular subsequent effect of a decreased soil pH caused by excessive acid deposition:� the breakdown of aluminosilicates to yield toxic levels of soluble aluminum.� This will be discussed in more detail later in the introduction.�  ����������� Below are the 1999 isopleth maps showing the deposition levels of SO4, NO3, and pH for the United States.� This data is very similar to that for the previous six years.� Therefore, it would be redundant to show all three maps for the available seven years.� (National)� These maps should show, more qualitatively than quantitatively, the distribution of these various compounds and their relative severity throughout the country.� It is apparent that those areas in the Eastern portion of the country suffer more from the baneful effects of acid deposition as a result of their increased levels of industry.� We must henceforth keep in mind that this experiment is most relevant in those areas of the country and similar areas around the world.�      **Item I-XI**  **Item I-XII**  **Item I-XIII**    **II. Introduction to the pH Scale.�** �pH is a unit of measure that indicates the level of alkalinity or acidity of a substance or solution.� The pH scale covers the range from zero to fourteen.� The term pH is derived from �p�, the mathematical symbol of the negative logarithm, and �H�, the chemical symbol for Hydrogen.� Its formal definition is the negative logarithm of the Hydrogen Ion activity (Omega).� Therefore, pH = -log(H+).� If a substance has a higher concentration of H+ ions relative to �OH ions, it is acidic and below 7 on the pH scale.� The converse is true as well.� The change in one unit pH represents a tenfold change in the concentration of protons (Omega).� An ion, whether positive or negative, is a charged particle.� Ions can be created when an atom or molecule gains or loses at least one electron.�� The presence of ions in solution allows electrical energy to be passed through the solution as a conductor (Omega). Different compounds form ions in solution in different amounts.� These amounts depend on the ability of the atoms to take or give up electrons. Compounds will dissociate (or ionize) in a given solution to form hydrogen (H+) or hydroxyl (OH-) ions in the solution (Omega).� Molecules that break down easily form strong acids or bases when in aqueous solutions.� Buffers are solutions that have stable pH values and a tendency to resist changes in that pH level (Omega).� In an aqueous solution, hydrogen ions normally combine with the water solvent to form the hydronium ion (H3O+).� pH measurements of these solutions are therefore measurements of the hydronium ion concentration (Omega). The terms "hydronium ion" and "hydrogen ion" are, therefore, able to be used interchangeably in reference to pH.�    **III. Introduction to Perlite.�** �Perlite, the growing medium for our experiment, is a mineral of volcanic origin and a chemically inert material that is extracted in New Zealand, the U.S., and other countries by scraping the earth's surface (Aquabotanic). The method through which perlite is processed includes a grinding and heat treatment, which results in very lightweight, white sterile fragments (Aquabotanic).� Perlite has a very low cation exchange capacity, low water-holding capacity and neutral pH (Aquabotanic).� Because perlite contains only minute amounts of plant nutrients, liquid feeding will be our ideal mode of fertilization (Aquabotanic). Perlite contains, on average, 47.5% oxygen, 33.8% silicon, 7.2% aluminum, 3.5% potassium, 3.4% sodium, 3.0% bound water, 0.6% iron and calcium, and 0.2% magnesium and trace elements (Perlite Institute).� Each of these elements is most often found mixed with glassy silicates, except of course silicon (Perlite Institute).� We chose perlite for this experiment because it has almost exactly the amount of aluminum determined to be the arbitrary average reference for soil, 7.1%.� It will act as a controlling factor in the experiment because it is inert and will allow each plant to be subjected to exactly the same nutrient schedule.� It will also eliminate any confounding variables inherent with the use of naturally occurring soil, such as unequal distribution of nutrients, aluminum, etc.� It also has low levels of magnesium and other elements that could also contribute to toxicity.� Any element below three percent will be deemed negligible in our experiment.� A picture of a perlite particle measuring roughly 4mm across is shown below:  **Item III-I**      **IV. Introduction to Aluminum Toxicity.�** Aluminum (Al) is the most abundant metallic element on the earth�s surface, comprising more than eight percent of the earth�s crust.� Only oxygen and silicon are more abundant (Spectrum).� The characteristic range of aluminum in soils is from roughly 1% to 30% (Lindsay).� Years ago, the arbitrary aluminum reference concentration for all soils was established at 7.1% (Lindsay).� Perlite conveniently contains 7.2%, almost exactly the same concentration as that specified by the arbitrary aluminum reference concentration.  ����������� Soil acidity is common to soils where precipitation is great enough to leach sufficient amounts of exchangeable bases from the soil�s top layer.� This leaching of exchangeable bases diminishes the acid neutralizing capacity of the soil and causes its pH to fall (Brady).� This lowered pH of the soil can have many effects on the behavior of the aluminum it contains.�  Aluminum bonds with oxygen to form many different compounds.� It is largely found bonded to oxygen ions in tetrahedral coordination in natural rocks.� As the rocks begin to deteriorate, the aluminum tends to demonstrate more octahedral bonding (Do).� The dissociation of aluminum can be facilitated by the input of acids from any of the anthropogenic or natural sources previously discussed (McBride).� The hydrated aluminum ion is classified as an acid itself because it contains protons (H+) that are removable from the water molecules that surround it.� However, these compounds are traditionally written, for convenience�s sake, without the formulas for water although it is present in the compounds.� The octahedrally bonded aluminum complex with six water groups (-OH2) occurs in solutions with pH 5.0 or below.� When the pH increases, the aluminum, hydrogen, oxygen complex loses (-OH2)�s by giving of (H+) ions and retaining��� �(-OH) ions.� Both the protons (H+) and the (-OH) groups are absorbed by the soil via cation exchange sites (Jackson).� The octahedrally bonded complex found at lower pH is held less strongly by electrostatic attraction and hydrogen bonding.� It is easily absorbed from the soil and the silicates it is bound up in because it is not subjected to the attraction to cation exchange sites in the soil that the compounds with more (-OH) groups must deal with.� The acid-based reactions that increase the solubility of aluminum as pH decreases are shown below (McBride):       |  | | --- | | **Acidic** | | Al(H2O)6 3+ + H2O �� Al(H2O)5 OH 2+ + H+ | | Al(H2O)5 OH2+ + H2O ��� Al(H2O)4 (OH)2 + + H+ | | Al(H2O)4 (OH)2+ + H2O �� Al(H2O)3 (OH)3 0� + H+ | | Al(H2O)3 (OH)30 + H2O �� Al(H2O)2 (OH)4- + H+ | | **Basic** |   **Item IV-I**    **Al(H2O)6 3+ is the primary soluble form of aluminum whose phytotoxic qualities we are going to examine in this research study.� From this point on, however, we will write it as Al3+.�**  **����������� The effects of aluminum on plants have been an important area of research for a long time in the agriculture industry.� Toxic levels of soluble aluminum (Al3+) have been shown to diminish both the height and weight of plants.� Al3+ is known to have a toxic effect on a plant�s roots, diminishing its metabolism by decreasing mineral nutrition and hydration.� The single most apparent symptom of aluminum toxicity is the inhibition of root growth, which has become an accepted measure of aluminum stress for plants (Herrera-Estrella).� This decrease in root elongation and proliferation eventually leads to decreased nutrient extraction and hydration.� There is no evidence of aluminum�s necessity in plants.� It has been discovered to be a factor in the production of a blue pigment in a plant called Hydrangea, but that is about it (Chenery).�**  **As well as diminished root proliferation, dwindling absorption of phosphorus is an important symptom of aluminum toxicity.� Aluminum bonds to and renders useless the phosphorus within the roots of plants, and thereby destroys the metabolic rate of plants (Naidoo).� It has been discovered that aluminum and phosphorus were mainly concentrated around the outer cells of the root caps (Naidoo).� There can even be seen a cluster formed by the aluminum and phosphorus at the cell surface when viewing under a microscope a plant that was exposed to conditions favorable for aluminum toxicity (Naidoo).� It was determined that 85% to 95% of the aluminum in the roots could be found near or on cell walls (Clarkson).� It has also been documented that the precipitate occurs not as a smooth film, but as scattered clumps along the cell surface (McCormick).� Needless to say, a large clump of aluminum and phosphorus sticking to the cell surface would, according to the fluid-mosaic model of cell membranes, certainly cause some extreme disturbances in a plant�s metabolic activity.�**  **����������� Yet another area of study involves the aluminum that penetrates within the cell and causes cell division and metabolic activity to stop immediately.� This reaction has to be a direct result of an interference with metabolic activity because simply decreasing the available phosphorus would not have such an immediate effect (Clarkson).� This might be caused by the aluminum bonding to the phosphates in compounds such as ATP.� The phosphates in nucleic acids and lipids can also be affected.� It can act as a competitive inhibitor and may cause the cell protoplasm to coagulate (Clarkson).�**  **����������� For the soluble aluminum (Al3+) to affect the plant, it must enter where lateral roots begin to grow.� As they push their way through the endodermis of the root, the aluminum can seep in the temporary hole.� The hole in the endodermal layer then provides access to the inner portion of the root.� Once within the plant, which has a neutral or basic composition, the aluminum can also reverse its acid-based reaction and convert with oxygen and hydrogen molecules back to its large, tetrahedrally bonded compound.� In the process, it releases protons, which will cause further damage to the plant by disrupting its natural pH balance (Do).�**  **We should also note that aluminum cannot easily be measured in soil because it is only the soluble aluminum that is available for plant uptake and therefore able to have detrimental effects on plant metabolism.� No correlation has been found between the levels of soluble aluminum and total aluminum in soil.� This is why a secondary, and rather roundabout, approach must be taken to measuring levels of aluminum toxicity.� The method we will be using involves testing plants� metabolism by recording their ability to react to rapid decreases in temperature.� A plant with a normal metabolic rate should be able to react to temperature drops with sufficient speed to survive.� However, a plant whose metabolism has been slowed via aluminum toxicity may be unable to close its guard cells in time to shield its stomata.� The stomata are pores in the plant that function as gaseous exchange sites through which CO2 enters and water vapor is lost (Raven).� If a plant cannot close its guard cells quickly enough to prevent the water within the plant from freezing, it will likely die as a result of a rapid decrease in temperature.� This variable (survival or demise of the plant) is one method by which we will attempt to determine if aluminum toxicity indeed endangers plants by slowing their metabolism.� The second method will be the measurement and analysis of the root length of the tested plants.�**      **V. Conclusions to the Introduction.� Now that we have gained all the necessary reference knowledge that will be necessary to recall throughout the study, we may proceed with confidence.� It has been well established that the problem of aluminum toxicity�s detrimental effects on plant metabolism are of significant importance.� To measure the effect of aluminum toxicity on the metabolism of our chosen test plant, a winter vegetable called the cherry belle radish, we will first chill both a control group watered with neutral water and a test group watered with simulated acid rain in a refrigerator for an hour and then freeze both groups for an hour.� The same procedure will then be used on a third group watered with simulated acid rain and treated with a buffer.� We will then conduct our study by measuring two dependent variables (root length and survival) when subjected to one independent variable (presence or absence of simulated acid rain).� We believe that the decreased metabolic rate of the group watered with simulated acid rain will render the plants unable to react quickly enough to the rapid changes in temperature by closing their guard cells to avoid freezing.� The most exciting thing to us about our project is the fact that, to our knowledge, it has never been done before.� This appears to be an entirely new experiment.� Nowhere in our research did we see any experimentation similar to our own being done.� To learn more about our study, please proceed to the abstract, hypothesis, prediction, and procedure.�** |  |
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