AIM - To compare the solubility of some common salts in water?

Objective

The goal of this project is to measure the solubilities of some common chemicals: table salt (NaCl), Epsom salts (MgSO₄), and sugar (sucrose, C₁₂H₂₂O₁₁).

Introduction

A good part of the substances we deal with in daily life, such as milk, gasoline, shampoo, wood, steel and air are mixtures. When the mixture is homogenous, that is to say, when its components are intermingled evenly, it is called a solution. There are various types of solutions, and these can be categorized by state (gas, liquid, or solid). The chart below gives some examples of solutions in different states. Many essential chemical reactions and natural processes occur in liquid solutions, particularly those containing water (aqueous solutions) because so many things dissolve in water. In fact, water is sometimes referred to as the universal solvent. The electrical charges in water molecules help dissolve different kinds of substances. Solutions form when the force of attraction between solute and solvent is greater than the force of attraction between the particles in the solute. Two examples of such important processes are the uptake of nutrients by plants, and the chemical weathering of minerals. Chemical weathering begins to take place when carbon dioxide in the air dissolves in rainwater. A solution called carbonic acid is formed. The process is then completed as the acidic water seeps into rocks and dissolves underground limestone deposits. Sometimes, the dissolving of soluble minerals in rocks can even lead to the formation of caves.

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<td>Sea water, sugar solution</td>
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If one takes a moment to consider aqueous solutions, one quickly observes that they exhibit many interesting properties. For example, the tap water in your kitchen sink does not freeze at exactly 0°C. This is because tap water is not pure water; it contains dissolved solutes. Some tap water, commonly known as hard water, contains mineral solutes such as calcium carbonate, magnesium sulfate, calcium chloride, and iron sulfate. Another interesting solution property is...
exhibited with salt and ice. Have you ever had the chore of throwing salt on an icy sidewalk? When the ice begins to melt, the salt dissolves in the water and forms salt water. What happens to the freezing point of water when salt is added to it? Even some organisms have evolved to survive freezing water temperatures with natural "antifreeze." Certain artic fish have blood containing a high concentration of a specific protein. This protein behaves like a solute in a solution and lowers the freezing point of the blood. Going to the other end of the spectrum, one can also observe that the boiling point of a solution is affected by the addition of a solute. Do eggs cook faster or slower when salt is added to the pot of water? These two properties, namely freezing-point depression and boiling-point elevation, are called **colligative** properties (properties that depend on the number of molecules, but not on their chemical nature). Exploring these properties and others of aqueous solutions are just some of the many ways that you could expand the scope of this project.

Finally, if you enjoy learning about solutions or other areas of chemistry, consider a career in the physical sciences. One example is working as an **analytical** chemist. Such chemists analyze the chemical composition of substances. They conduct many experiments to identify special characteristics of substances for a wide variety of reasons. Perhaps they are charged with testing municipal drinking water for its purity, or perhaps they must test a forensic sample for evidence in a trial. Whatever the reason, it is challenging work that requires precision and creative thought.

In this project you will measure the aqueous solubility of some common household chemicals: table salt (NaCl), Epsom salts (MgSO₄), and sugar (sucrose, C₁₂H₂₂O₁₁). How much of each chemical can dissolve in a given volume of water?

**Terms, Concepts and Questions to Start Background Research**

To do this project, you should do research that enables you to understand the following terms and concepts:

- Solution
- Solute
- Solvent
- Soluble vs. insoluble
- Chemical structure of water
- Polar molecule
- Force of attraction
- Concentration of a solution
- Dilute vs. concentrated
- Sodium chloride (NaCl)
- Magnesium sulfate (MgSO₄)
- Sucrose (C₁₂H₂₂O₁₁)

**Questions**

- What is a saturated solution?
What is the difference between a saturated solution and an unsaturated solution?

Bibliography

- The experimental procedure is based on:
  - CBSE Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

- Distilled water
- Metric liquid measuring cup (or graduated cylinder)
- Three clean glass jars or beakers
- Non-iodized table salt (NaCl)
- Epsom salts (MgSO₄)
- Sugar (sucrose, C₁₂H₂₂O₁₁)
- Disposable plastic spoons
- Thermometer
- Three shallow plates or saucers
- Oven
- Electronic kitchen balance (accurate to 0.1 g)
Do your background research so that you are familiar with the terms, concepts, and questions, above.

**Determining Solubility: Method 1**

1. Measure 100 mL of distilled water and pour into a clean, empty beaker or jar.
2. Use the kitchen balance to weigh out the suggested amount (see below) of the solute to be tested.
   a. 50 g Non-iodized table salt (NaCl)
   b. 50 g Epsom salts (MgSO₄)
   c. 250 g Sugar (sucrose, C₁₂H₂₂O₁₁)
3. Add a small amount of the solute to the water and stir with a clean disposable spoon until dissolved.
4. Repeat this process, always adding a small amount until the solution will no longer dissolve.
5. Weigh the amount of solute remaining to determine how much was added to the solution. Save your saturated solutions for the second method.

**Determining Solubility: Method 2**

1. Label the underside of each saucer with tape, one for each solution.
2. Weigh the empty saucer and record the weight.
3. Pour in 10–15 mL of the appropriate saturated solution (corresponding to the label on the saucer).
4. Weigh the saucer + solution and record the weight.
5. Repeat steps 2–4 for each of the three solutions.
6. Put the saucers in a warm place (e.g., an oven on low heat) and allow the water to evaporate.
7. Re-weigh the saucers + dry crystals.
   a. Tip: make sure all the water has evaporated by weighing each saucer several times, with an interval back in the oven in between, to make sure the weight is no longer changing.

**Analyzing Your Results**

1. To make sure that your results are reproducible, you should repeat your solubility experiment at least three separate times for each chemical.
2. For each solubility determined by Method 1, you will have the original volume of water, the total mass of the solute, and the remaining mass of the solute. You can calculate how much of the solute was dissolved.
3. For each solubility determination by Method 2, you will have the mass of the dry solid after evaporation, and the mass of the original solution. You can calculate the mass of the water that evaporated.
4. Calculate the average solubility, in grams of solute per 100 mL of water, as determined by each method.
5. More advanced students should also calculate the standard deviation of the solubility, as determined by each method.
6. Compare the results of the two methods.
7. Compare your results to published solubilities for the three chemicals.

Variations

- Let’s say that instead of starting with pure water, you tried to dissolve Epsom salts (MgSO$_4$) in a saturated solution of NaCl. Do you think this would work? How much MgSO$_4$ would you expect to dissolve? Would it be more, less or the same amount as in an equal volume of distilled water? Design an experiment to find out.
- You could also try the experiment above with the other five pair-wise combinations of the three chemicals.
- Another variation you could try is an experiment on the how fast solutes dissolve. What can you do to increase the rate at which a solute dissolves in a solvent? How much more quickly does the solute dissolve, compared to when the solute is simply added to the solvent?
AIM-Science Projects {Chemistry}~Charles's Law: Volume vs. Temperature of a Gas at Constant Pressure

Objective

The goal of this project is to measure the relationship between the volume of a gas and its temperature, when the pressure of the gas is held constant.

Introduction

This is a modern version of a classic experiment by Jacques Charles (who was also interested in flying balloons). Charles studied the volume of a sample of air—sealed in a glass tube with a U-shaped curve—as he systematically changed the temperature by immersing the tube in a water bath. The air was trapped by a column of mercury, added to the open end of the tube. By changing the amount of mercury in the tube, Charles could maintain a constant pressure on the trapped air as the temperature was changed. Charles's apparatus was an example of a manometer, a device used to measure pressure.

You can learn more about how manometers work, and even run a simulated Charles's Law experiment by visiting the Chemistry Applet website (see Bibliography). This would be excellent preparation for doing the experiment on your own, so we highly recommend it. Note: to calculate the volume of gas in the applets, you will need to know that the inside diameter of the applet's manometer tube is 4.286 cm.

You can repeat Charles's experiments for yourself with an inexpensive, modern apparatus based on a disposable plastic syringe and a water bath. (Mercury is a dangerous neurotoxin, so we'll avoid working with it.)

Terms, Concepts and Questions to Start Background Research

To do this project, you should do research that enables you to understand the following terms and concepts:

- pressure,
- volume,
- Charles's Law,
- ideal gas,
- atmospheric pressure,
- manometer,
kelvins,
absolute zero.

Questions

- What assumption is made about the pressure of the gas in this experiment?
- What is the relationship between the degrees Celsius and kelvins?

Bibliography

- Here are the Chemistry Applet webpages mentioned in the Introduction. These will really help your understanding of Charles's Law if you take the time to do the virtual experiments! To calculate the volume of gas in the applets, you will need to know that **the inside diameter of the applet's manometer tube is 4.286 cm.**
- The Sizes.com website has an exhaustive index of units of measure, including both degrees Celsius and kelvins:
  - CBSE Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

- 35 ml syringe (available, for example, from Science First ("Gas Law Demonstrator Kit", Science First, Buffalo, NY. http://www.sciencefirst.com/vw_prdct_mdl.asp?prdct_mdl_cd=30170),
- a homemade clamp to hold syringe underwater, which can be made with:
  - two sturdy chopsticks (or two sturdy wood dowels) longer than the diameter of your cooking pot,
  - two rubber bands,
  - a weight (e.g., a can of soup);
- small piece of wire,
- thermometer (calibrated in °C, range at least 0–100°C),
- water,
- ice,
Experimental Procedure

Experimental Apparatus

1. Before starting the experiment, do your background research so that you are knowledgeable about the terms, concepts and questions, above.
2. With the plunger removed from the syringe, seal the tip of the syringe with a tight-fitting cap. If a suitable cap is not available, you can try epoxy or silicone sealant. Allow the epoxy or silicone the recommended curing time before proceeding with the experiment. (Note: if you seal the tip with the plunger in place, you will probably not be able to remove the plunger unless you destroy the seal. Why?)
3. When your sealed syringe is ready for use, insert the plunger to the 20 ml mark of the syringe along with a thin wire as shown in the diagram above. The wire will allow air to escape from beneath the plunger, equalizing the pressure in the syringe with the atmosphere. Use the lower ring of the plunger as your indicator.
4. Hold the plunger in place and carefully withdraw the wire.
5. Make sure that the plunger can move freely in the syringe, and that the tip of the syringe is well-sealed. Give the plunger a small downward push, and verify that it springs back. If it does not, you may need to lubricate the side of the plunger with a small amount of silicone lubricant or you may not have sealed the tip of your syringe properly.
6. When you are satisfied with the results of the previous step, record the initial volume of air in the syringe and the ambient temperature.
7. You will be immersing the syringe into a water bath, and observing the changes in volume of the gas as you change the temperature of
the water. Since the air in the syringe will make it buoyant, you need a way to hold the syringe under the water. If you have a ringstand and clamp, you're all set. Otherwise, you can put together a homemade clamp with materials you'll probably have around the house. Here's how:

a. Wrap a rubber band around the top of the syringe tube, just below the finger flanges.

b. Insert the chopsticks (as noted in Materials & Equipment, wood dowels can be substituted for chopsticks) through loops of this rubber band, one on either side of the syringe. Slide the syringe so that it is about 7–8 cm (3 in) from the ends of the chopsticks.

c. Wrap the second rubber band around the short ends of the chopsticks. This will make a "V" shape, with the syringe held tightly down near the point.

d. This second rubber band can also be used to hold the thermometer upright in the water. Keep the bulb immersed in the water, but not touching the side or bottom of the pot.

e. Place this assembly on the top of your cooking pot, so that the chopsticks are supported by the rim of the pot and the syringe sticks down into the pot.

f. To hold the syringe in place when the pot is filled with water, place your weight (e.g., a can of soup) on top of the wide end of the "V" made by the chopsticks.

g. Make any necessary adjustments to make the syringe and thermometer stable, and make sure that you can read the scale on the syringe.

Making the Measurements and Presenting Your Results

1. Remove the syringe and thermometer assembly from the pot and set them aside.

2. Place the pot on the stove, but don't turn on the burner yet. Fill the pot with ice cubes and enough water to immerse the syringe to somewhere between the 30 and 35 ml marks.

3. Replace the syringe and thermometer assembly, and weight it down securely.

4. Allow several minutes temperature in the water bath to stabilize and for the temperature of the air in the syringe to equilibrate with the water bath. Gentle stirring may help, but be careful not to break the thermometer or knock your weight off your clamp.

5. Record the temperature of the water bath and the volume of the air in the syringe. You may want to tap the plunger lightly to make sure it is free to move. (If necessary, carefully (and briefly) lift the syringe out of the water to read the volume. You may want to have an adult help you with this part.)

6. Turn the burner on (no higher than medium heat) to gradually heat the water. At regular intervals (e.g., every 10°C), turn the heat off and allow the temperature to stabilize. Again, record the temperature of the water bath and the volume of air in the syringe.
7. Repeat the previous step up to 80 or 90°C. The pot will be quite full, so it is best to avoid boiling the water.

8. As with any experiment, it is a good idea to repeat your measurements to be sure that your results are consistent. We suggest at least three separate trials. (Note that the temperatures used do not need to be exactly the same from trial to trial!)

9. Make a graph of gas volume vs. temperature for all of your data points. It’s a good idea to use a different symbol for each of your trials (if something was wrong with one particular trial, it may help you understand what went wrong).

Questions

1. What is the relationship between volume and temperature in your data set?
2. Can you extrapolate from your data to find the temperature that corresponds to a gas volume of zero? How confident are you with this result, and why?
3. Would your data look different if you used kelvins for the temperature axis instead of degrees Celsius?
4. Was the assumption of constant pressure valid?
5. What are the possible sources of error in your experiment?
AIM-Science Projects

{Chemistry}~Measuring the Amount of Acid in Vinegar by Titration with an Indicator Solution

Objective

The goal of this project is to determine the amount of acid in different types of vinegar using titration with a colored pH indicator to determine the endpoint.

Introduction

Vinegar is a solution made from the fermentation of ethanol (CH₃CH₂OH), which in turn was previously fermented from sugar. The fermentation of ethanol results in the production of acetic acid (CH₃COOH). There are many different types of vinegar, each starting from a different original sugar source (e.g., rice, wine, malt, etc.). The amount of acetic acid in vinegar can vary, typically between 4 to 6% for table vinegar, but up to three times higher (18%) for pickling vinegar (Wikipedia contributors, 2007).

In this project, you will determine the amount of acid in different vinegars using titration, a common technique in chemistry. Titration is a way to measure the unknown amount of a chemical in a solution (the titrant) by adding a measured amount of a chemical with a known concentration (the titrating solution). The titrating solution reacts with the titrant, and the endpoint of the reaction is monitored in some way. The concentration of the titrant can now be calculated from the amount of titrating solution added, and the ratio of the two chemicals in the chemical equation for the reaction. Let's go through the process with a specific example: the titration of acetic acid. But before we go over titration, here is a quick review of the chemistry of acids and bases.

It all has to do with hydrogen ions (abbreviated with the chemical symbol H⁺). In water (H₂O), a small number of the molecules dissociate (split up). Some of the water molecules lose a hydrogen and become hydroxyl ions (OH⁻). The "lost" hydrogen ions join up with water molecules to form hydronium ions (H₃O⁺). By convention (and for simplicity in writing chemical equations), hydronium ions are referred to as hydrogen ions H⁺. In pure water, there are an equal number of hydrogen ions and hydroxyl ions. The solution is neither acidic or basic.

An acid, like acetic acid, is a substance that donates hydrogen ions. When acetic acid is dissolved in water, the balance between hydrogen ions and hydroxyl ions is shifted. Now there are more hydrogen ions than hydroxyl ions in the solution. This kind of solution is acidic.
A base is a substance that accepts hydrogen ions. When a base is dissolved in water, the balance between hydrogen ions and hydroxyl ions shifts the opposite way. Because the base "soaks up" hydrogen ions, the result is a solution with more hydroxyl ions than hydrogen ions. This kind of solution is alkaline.

To measure the acidity of a vinegar solution, you can add enough hydroxyl ions to balance out the added hydrogen ions from the acid. The hydroxyl ions will react with the hydrogen ions to produce water. In order for a titration to work, you need three things:

1. a titration solution (contains hydroxyl ions with a precisely known concentration),
2. a method for delivering a precisely measured volume of the titrating solution, and
3. a means of indicating when the endpoint has been reached.

For the titrating solution, you'll use a dilute solution of sodium hydroxide (NaOH). Sodium hydroxide is a strong base, which means that it dissociates almost completely in water. So for every NaOH molecule that you add to the solution, you can expect to produce a hydroxyl ion.

To dispense an accurately measured volume of the titrating solution, you will use a buret. A buret is a long tube with a valve at the bottom and graduated markings on the outside to measure the volume contained in the buret. The buret is mounted on a ring stand, directly above the titrant solution (as shown below).

The illustration shows a buret (filled with titration solution) mounted on a ring stand above a beaker (containing the titrant solution) (G. Carboni, 2004).

Solutions in the buret tend to creep up the sides of the glass at the surface of the liquid. This is due to the surface tension of water. The
surface of the liquid thus forms a curve, called a meniscus. To measure the volume of the liquid in the buret, always read from the bottom of the meniscus. In the illustration below, I'd say that the fluid level is 14.58 mL.

Always read the fluid level in the buret from the bottom of the meniscus (G. Carboni, 2004).

In this experiment, you will use an indicator solution called phenolphthalein. (I love to say that word: fee-nol-fthay-leen!) Phenolphthalein is colorless when the solution is acidic or neutral. When the solution becomes slightly basic, phenolphthalein turns pinkish, and then light purple as the solution becomes more basic. So when your vinegar solution starts to turn pink, you know that the titration is complete.

Which type of vinegar do you think will have the most acetic acid? Find out for yourself with this project.

Terms, Concepts and Questions to Start Background Research

To do this project, you should do research that enables you to understand the following terms and concepts:

- Vinegar
- Acids and bases
  - Acetic acid (CH₃COOH)
  - Sodium hydroxide (NaOH)
- pH scale (see the Science Buddies Chemistry Resource, Acids, Bases, and the pH Scale)
- Indicator solutions (e.g., phenolphthalein)
- Stoichiometry
- Titration
- Meniscus
- Buret
- Titrant
Questions

- What value of pH is neutral?
- What range of pH values is acidic?
- What range of pH values is basic?
- How does a pH indicator work?
- At what pH does phenolphthalein change from colorless to pinkish?

Bibliography

- Here is a good introduction to acid-base chemistry:
  o This project is based on the vinegar titration experiment about 4/5 of the way down the page:
- These webpages have a quick review of exponents and logarithms:
- For more information about the pH scale, try these references:
- For more information on titration techniques see Chemistry Lab Techniques.
- CBSE Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

- Vinegar, at least three different types
Tip: it will be easier to see the indicator change color with lighter-colored vinegars.

- Distilled water
- Small funnel (do not use for food after using it for chemistry)
- The following items can be ordered from Science Kit & Boreal Laboratories:
  - Chemical safety goggles
  - Lab apron
  - Rubber (latex) gloves
  - 0.5% Phenolphthalein solution in alcohol (pH indicator solution)
  - 0.1 M sodium hydroxide solution
    - Sodium hydroxide is caustic, which means it will cause a chemical burn on bare skin.
    - You will have to order this chemical through your school.
    - This solution is fairly dilute and relatively safe to use with proper chemical safety precautions (chemical safety goggles, lab coat or apron, and rubber gloves).
    - Tip: since each molecule of sodium hydroxide (NaOH) can produce one hydroxide ion (OH\(^-\)), 0.1 N sodium hydroxide is the same as 0.1 M.
  - 125 mL Erlenmeyer flask
  - 25 or 50 mL buret
  - 10 mL graduated cylinder
  - Ring stand
  - Buret clamp

Experimental Procedure

Note: this project requires the use of a sodium hydroxide solution, which is caustic. You will have to order this chemical through your school. Proper safety precautions should be used when working with this solution, including:

- Chemical safety goggles
- Lab coat/apron
- Gloves

Performing the Titration

1. Do your background research so that you are knowledgeable about the terms, concepts, and questions, above.
2. Since you will be working with dilute sodium hydroxide, you should take proper safety precautions:
   a. Wear safety goggles for chemistry, an apron (or lab coat), and a pair of rubber gloves.
   b. If you spill sodium hydroxide on your skin, wash it off quickly with lots of running water.
3. Pour 1.5 ml of vinegar in an Erlenmeyer flask.
   a. These flasks are designed so that you can swirl the solution inside without spilling it.
b. Tip: you can also use a regular beaker and a stirring rod to keep the solution mixed as you titrate.

4. Dilute the vinegar with about 50 ml of distilled water.

5. Add 3 drops of 0.5% phenolphthalein solution.
   a. Phenolphthalein solution is colorless at acidic pH, and turns light purple at about pH 8.3.
   b. The vinegar solution is acidic, so it should remain colorless.

6. Use the buret clamp to attach the buret to the ring stand. The opening at the bottom of the buret should be just above the height of the Erlenmeyer flask you use for the vinegar/water/phenolphthalein solution.

7. Use a funnel to fill the buret with a 0.1 M solution of sodium hydroxide.

8. Note the starting level of the sodium hydroxide solution in the buret. Remember to read from the bottom of the meniscus. In the illustration below, I'd say that the fluid level is 14.58 mL.

   ![Image of buret](image)

   Always read the fluid level in the buret from the bottom of the meniscus. (G. Carboni, 2004)

9. Put the vinegar solution to be titrated under the buret. The illustration below shows an example of the experimental setup at the beginning of the titration.
At the start of the vinegar titration, the phenolphthalein is colorless. (G. Carboni, 2004)

10. Slowly drip the solution of sodium hydroxide into the vinegar solution. Swirl the flask gently to mix the solution, while keeping the opening underneath the buret. (Alternative is to use a beaker and stirring rod—but be careful not to hit the buret with the stirring rod.)

11. At some point you will see a pink color in the vinegar solution when the sodium hydroxide is added, but the color will quickly disappear as the solution is mixed. When this happens, slow the buret to drop-by-drop addition.

12. When the vinegar solution turns pink and remains that color even with mixing, the titration is complete. Close the tap (or pinch valve) of the buret. The illustration below shows how the solution color changes at the endpoint of the titration.
The endpoint of the titration is reached when the phenolphthalein in solution turns pinkish. (G. Carboni, 2004)

13. Note the remaining level of the sodium hydroxide solution in the buret. Remember to read from the bottom of the meniscus.
14. Subtract the initial level from the remaining level to figure out how much titrating solution you have used.
15. For each vinegar that you test, repeat the titration at least three times. If you are careful with all of your volume measurements, the results of your three repeated trials should agree within 0.1 mL.

Analyzing Your Results

Here's how to figure out how much acetic acid was in each sample.

1. Determine the number of moles of sodium hydroxide used to titrate the vinegar.
   a. Multiply the volume of added sodium hydroxide (in liters) by the concentration (in moles/liter).
   b. For example, if you added 12.5 mL of sodium hydroxide, the number of moles would be 0.0125 L × 0.1 moles/L = 0.00125 moles.

2. Determine the concentration of acetic acid in the vinegar.
   a. The number of moles of sodium hydroxide equals the number of moles of acetic acid in your vinegar sample ($M_s$).
   b. The sample volume was 1.5 mL ($V_s$).
   c. You can use a proportion to determine the number of moles of acetic acid ($M_x$) in a standard volume ($V_x = 1$ L) of vinegar: $M_s/V_s = M_x/V_x$.
   d. Continuing with the previous example, the number of moles of acetic acid would be 0.00125. Dividing by 0.0015 L gives 0.833 moles of acetic acid per liter, or a concentration of 0.833 M.
   e. You can also calculate the concentration in terms of grams of acetic acid per liter. To do this, multiply the molar concentration of acetic acid by the molecular mass of acetic acid, which is 60. In the case of our example, the concentration would be 0.833 × 60 = 50 g/L, or 5.0%.

3. Repeat the calculations for each type of vinegar you tested. Which vinegar had the highest concentration of acetic acid?

Variations

- Measure the acidity of solutions such as beer or wine.
- Measure the acidity of other beverages, such as: different fruit juices, soda, sport drinks, coffee, or teas.
- Measure the acidity of fermenting apple cider over time.
- Measure the acidity of uncorked wine over time (check once a day over the course of a week).
AIM-Science Projects

{Chemistry}~Measuring Surface Tension of Water with a Penny

Objective

The goal of this project is to investigate how added salt and added detergent affect the surface tension of water.

Introduction

Water molecules—good old H₂O—are made of one oxygen and two hydrogen atoms. The single oxygen and two hydrogen atoms are held together because they share electrons—this is called a covalent bond. The hydrogen atoms don't line up on opposite sides of the oxygen atom, as you might think. Instead they are at an angle of about 105° (if they were on opposite sides of the oxygen atom the angle would, of course, be 180°).

The oxygen atom tends to hold on to the shared electrons from the hydrogen atoms more tightly, so each end of the water molecule ends up with a partial charge. The oxygen portion of the molecule has a partial negative charge, and the hydrogen ends of the molecule have a partial positive charge. Another way of talking about the partial charges is to say that water molecules are polarized. Like a magnet, with a north and south pole, a water molecule has electrical poles. The oxygen atom is the negative pole, and each hydrogen atom is a positive pole.

These partial charges cause water molecules to interact with one another. Because opposite charges attract, water molecules tend to ‘stick’ to one another. The partial positive charges of the hydrogen atoms tend to align themselves with the partial negative charge of the oxygen atoms of neighboring water molecules. You can see models of this alignment in several of the references in the Bibliography section (Wiseth, date unknown; Hipschman, 1995a; Kimball, 2006). This tendency of water molecules to stick together due to the partial positive and negative charges is called hydrogen bonding.

Hydrogen bonding between water molecules leads to many interesting consequences at the visible, macroscopic level. For example: the boiling point of water, its surface tension, and its ability to dissolve salts are all related to hydrogen bonding.

The boiling point of water, 100°C, is unusually high for a molecule with such a low molecular weight. The boiling point is so high due to hydrogen bonding. On average, each water molecule interacts with about four others (each hydrogen atom interacts with the oxygen atom of separate water molecules, and each oxygen atom interacts with the
hydrogen atoms of two more water molecules). In water vapor, the molecules are too far apart for hydrogen bonding to occur, so boiling water means breaking up all of the hydrogen bonds in liquid water. Breaking those bonds takes energy, thus the high boiling point for water.

Hydrogen bonds also give liquid water a high surface tension. The water molecules on the surface have partners for hydrogen bonding only within the liquid; above the water surface there are no more molecules available for hydrogen bonding. This means that molecules at the surface experience a net force pulling them inward. If you fill a glass right up to the rim and then carefully add a few more drops of water, you can see that the glass can be overfilled without spilling. The surface tension of the water holds on to the 'extra' water as if there were a skin on the surface of the water.

Water is an excellent solvent for charged (polar) molecules like table salt, NaCl. In water, salt dissociates into positively charged sodium (Na\(^+\)) and negatively charged chloride (Cl\(^-\)) ions. The partial positive charge of the hydrogen ends of the water molecules surround the negatively charged chloride ions, and the partial negative charge of the oxygen ends of the water molecules surround the positively charged sodium ions. What effect will dissolved salt ions have on hydrogen bonds between water molecules?

Water behaves very differently when mixed with uncharged (nonpolar) molecules. An example of a nonpolar molecule is cooking oil. You may have heard the saying "oil and water don't mix," and this is why. Oil molecules are uncharged. Water molecules, as you have learned, are partially charged. The uncharged oil molecules disrupt the hydrogen bonding between water molecules. So when you try to mix oil and water, the oil ends up forming droplets within the water. The nonpolar oil molecules stick together and the polar water molecules stick together. Eventually, you get two layers, with the less dense oil floating on top of the denser water.

Nonpolar substances are sometimes called 'hydrophobic' (meaning 'water fearing'), and polar molecules are sometimes called 'hydrophilic' (meaning 'water loving') because of the two different interactions illustrated by salt and cooking oil.

Liquid detergents have dual properties. One end of the molecule is oily, and the other end is charged. In water, the oily ends of detergent molecules stick together, with the charged ends sticking out, into the water. Detergents can form small blobs in water (called micelles) and can also disperse, like oils, into a layer on the surface of the water (for illustrations, see Hipschman, 1995b). How do you think added detergent will affect the surface tension of water?

One way to find out is to count how many drops of water you can 'pile up' on top of a single penny. The Experimental Procedure section shows you how to do this with plain water, salt water, and water with detergent.
Terms, Concepts and Questions to Start Background Research

To do this project, you should do research that enables you to understand the following terms and concepts:

- surface tension,
- chemical structure of water,
- covalent bond,
- hydrogen bonds,
- polar solvent,
- non-polar solvent,
- hydrophobic,
- hydrophilic.

Questions

- What happens to salt when it is dissolved in water?
- How do you think adding salt to the water will affect the hydrogen bonds between water molecules?
  - What effect will this have on surface tension?
  - Do you think salt water will have more or less surface tension than plain tap water?
- What happens to detergent when it is dissolved in water?
- How will added detergent affect hydrogen bonds between water molecules?
  - What effect will this have on surface tension?
  - Do you think water with added detergent will have more or less surface tension than plain tap water?

Bibliography

- This animation shows how hydrogen bonding occurs between water molecules (requires Shockwave plug-in):
- These webpages about bubbles from San Francisco's Exploratorium explain how hydrogen bonds make water a "sticky" fluid, and how soap disrupts this "stickiness."
- Here is another brief description of hydrogen bonding in water:
• CBSE Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

• water,
• plastic transfer pipettes (or eyedropper),
  o one online source of transfer pipettes in small quantities is RachelsSupply.com, where you can get 10 fine-tipped pipettes (part #1N01) for $1.50 + shipping.
• salt,
• dishwashing detergent,
• clean glass jars (or beakers),
• measuring spoons.

Experimental Procedure

1. Holding the transfer pipette close to the surface of the penny, carefully pipet water droplets onto the penny, one at a time, counting each drop. Tips:
   a. The droplets should pool up on the penny, creating a big droplet of water.
   b. To make sure your count is accurate, hold the pipette far enough above the penny so that the drop has to fall a short distance before fusing with the droplet on the penny.
2. Stop pipetting when the droplet on the penny breaks up and overflows. The count for each trial is the number of drops that the penny could hold (in other words, count all of the drops except the one that caused the penny to overflow).
3. Repeat the measurement ten times for each solution that you test.
4. Test the following solutions:
   a. added salt: dissolve 1 teaspoon (6 grams) in 100 mL of water,
   b. added detergent: put 1 drop of liquid dishwashing detergent in 1 liter of water; do not shake–cap the container and gently tip it back and forth to mix.

Variations

• Try a series of increasing concentrations of salt (maximum solubility at room temperature is about 36 g salt/100 mL water). The best way to do this is by making a concentrated solution, and then making serial dilutions to make less-concentrated solutions. Does surface tension continue to change as more salt is added? Students who
have studied high school chemistry should compare molar ratios of NaCl and H₂O.

- Do you think that changing the temperature of the water would affect surface tension? How? Design an experiment to find out. Measuring surface tension on a penny is probably not the best design for this variation, because the temperature would not be well controlled. The volume of water is quite small, so the temperature could easily change. However, if you controlled the temperature of the water and the penny, you could probably get this to work. Another idea would be to find a different way to measure surface tension, using a larger volume of water.

- Does the surface of the penny matter? What happens if you coat the penny with a thin film of cooking oil? Wet a paper towel slightly with cooking oil. Wipe off the excess oil, then use the paper towel to wipe a thin film of oil on a penny. How many drops of water will the penny hold compared to a "normal" penny. Can you think of other surface treatments you could try? Could you make penny-sized disks of other materials to test? How important is the raised edge of the penny for holding the water? Does the 'heads' side hold more or less water than the 'tails' side?
Objective

The goal of this project is to determine which added material will make ice melt fastest.

Introduction

To make ice cream with an old-fashioned hand-crank machine, you need ice and rock salt to make the cream mixture cold enough to freeze. If you live in a cold climate, you've seen the trucks that salt and sand the streets after a snowfall to prevent ice from building up on the roads. In both of these instances, salt is acting to lower the freezing point of water.

For the ice cream maker, because the rock salt lowers the freezing point of the ice, the temperature of the ice/rock salt mixture can go below the normal freezing point of water. This makes it possible to freeze the ice cream mixture in the inner container of the ice cream machine. For the salt spread on streets in wintertime, the lowered freezing point means that snow and ice can melt even when the weather is below the normal freezing point of water. Both the ice cream maker and road salt are examples of freezing point depression.

Salt water is an example of a chemical solution. In a solution, there is a solvent (the water in this example), and a solute (the salt in this example). A molecule of the solute will dissolve (go into solution) when the force of attraction between solute molecule and the solvent molecules is greater than the force of attraction between the molecules of the solute. Water (H\(_2\)O) is a good solvent because it is partially polarized. The hydrogen ends of the water molecule have a partial positive charge, and the oxygen end of the molecule has a partial negative charge. This is because the oxygen atom holds on more tightly to the electrons it shares with the hydrogen atoms. The partial charges make it possible for water molecules to arrange themselves around charged atoms (ions) in solution, like the sodium (Na\(^+\)) and chloride (Cl\(^-\)) ions that dissociate when table salt dissolves in water.

Other substances that dissolve in water also lower the freezing point of the solution. The amount by which the freezing point is lowered depends only on the number of molecules dissolved, not on their chemical nature. This is an example of a **colligative property**. In this project, you'll investigate different substances to see how they affect the rate at which ice cubes melt. You'll test substances that dissolve in water (i.e., soluble substances), like salt and sugar, as well as substances that don't dissolve in water (i.e., insoluble substances), like
sand and pepper. Which substances will speed up the melting of the ice?

Terms, Concepts and Questions to Start Background Research

To do this project, you should do research that enables you to understand the following terms and concepts:

- Solution
- Solute
- Solvent
- Colligative properties
- Freezing point depression
- Phases of matter
  - Solid
  - Liquid
  - Gas
  - Plasma
- Phase transitions
  - Melting
  - Freezing
  - Evaporation
  - Condensation
  - Sublimation

Questions

- Which of the suggested test substances are soluble in water?
- Which of the suggested test substances are insoluble in water?

Bibliography

- For more information on colligative properties, see:

- For information on Avogadro's number and molecular weight, see:

- To try a simulated experiment on freezing point depression or boiling point elevation, see (Flash animation, requires browser plug-
in):

• Cbse Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

• Ice cubes
• Identical plates or saucers
• Timer
• Electronic kitchen balance (accurate to 0.1 g)
• Measuring cup
• Suggested materials to test for ice-melting ability
  o Table salt
  o Sugar
  o Sand
  o Pepper

Experimental Procedure

1. Do your background research so that you are knowledgeable about the terms, concepts, and questions, above.
2. You'll need a clean plate and several ice cubes for each of the substances to be tested.
3. Note the starting time, then carefully sprinkle one teaspoon of the substance to be tested over the ice cube.
4. After a fixed amount of time (say, 10 minutes), pour off the melted water into a measuring cup, and use the balance to measure the mass. Subtract the mass of the empty cup, and you'll have the mass of the melted water. Wait the same amount of time for each test.
5. Measure the remaining mass of the ice cube.
6. Repeat three times for each substance to be tested.
7. Use the same procedure to measure the melting rate for ice cubes with nothing added.
8. For each test, calculate the percentage of the ice cube that melted:

   \[
   \text{[mass of melt water]/[initial mass of ice cube]} \times 100
   \]

9. For each test, calculate the percentage of the ice cube remaining:

   \[
   \text{[remaining mass of ice cube]/[initial mass of ice cube]} \times 100
   \]
10. For each substance you tested, calculate the average amount of melted water produced (as a percentage of initial mass), and the average remaining ice cube mass (as a percentage of initial mass).

11. Did any substances speed up melting of the ice (compared to melting rate of plain ice cubes with nothing added)?

Variations

- Does the melting rate depend on the amount of solute added? Design an experiment to find out.
- Try your experiment in the refrigerator to simulate colder weather. Alternatively, if the outside temperature is wintry, take your experiment outdoors! Be sure to monitor the temperature regularly throughout your experiment.
- Do you think salt would melt ice in your freezer? Why or why not? Try it and find out.
AIM-Science Projects {Chemistry}~Make Your Own pH Paper

Objective

The goal of this project is to make your own pH indicator paper, and use it to measure the acidity and alkalinity of various solutions from around your house.

Introduction

In this project you'll learn how to make your own pH paper that you can use to find out if a solution is acidic or basic (alkaline). What does it mean for a solution to be acidic or alkaline?

It all has to do with hydrogen ions (abbreviated with the chemical symbol \(H^+\)). In water (\(H_2O\)), a small number of the molecules dissociate (split up). Some of the water molecules lose a hydrogen and become hydroxyl ions (\(OH^-\)). The "lost" hydrogen ions join up with water molecules to form hydronium ions (\(H_3O^+\)). For simplicity, hydronium ions are referred to as hydrogen ions \(H^+\). In pure water, there are an equal number of hydrogen ions and hydroxyl ions. The solution is neither acidic or basic.

An acid is a substance that donates hydrogen ions. Because of this, when an acid is dissolved in water, the balance between hydrogen ions and hydroxyl ions is shifted. Now there are more hydrogen ions than hydroxyl ions in the solution. This kind of solution is acidic.

A base is a substance that accepts hydrogen ions. When a base is dissolved in water, the balance between hydrogen ions and hydroxyl ions shifts the opposite way. Because the base "soaks up" hydrogen ions, the result is a solution with more hydroxyl ions than hydrogen ions. This kind of solution is alkaline.

Acidity and alkalinity are measured with a logarithmic scale called pH. Here's why: A strongly acidic solution can have one hundred million million (100,000,000,000,000) times more hydrogen ions than a strongly basic solution! The flip side, of course, is that a strongly basic solution can have 100,000,000,000,000 times more hydroxide ions than a strongly acidic solution. Moreover, the hydrogen ion and hydroxide ion concentrations in everyday solutions can vary over that entire range. In order to deal with these large numbers more easily, scientists use a logarithmic scale, the pH scale. Each one-unit change in the pH scale corresponds to a ten-fold change in hydrogen ion concentration. The pH scale ranges from 0 to 14. It's a lot easier to use a logarithmic scale instead of always having to write down all those zeros! By the way, notice how one hundred million million is a one with fourteen zeros after it? It's not coincidence, it's logarithms!
To be more precise, pH is the negative logarithm of the hydrogen ion concentration:

\[ \text{pH} = \log \frac{1}{[H^+]} = -\log [H^+] \]

The square brackets around the \( H^+ \) automatically mean "concentration" to a chemist. What the equation means is just what we said before: for each 1-unit change in pH, the hydrogen ion concentration changes ten-fold.

Pure water has a neutral pH of 7. pH values lower than 7 are acidic, and pH values higher than 7 are alkaline (basic). The table below has examples of substances with different pH values (Decelles, 2002; Environment Canada, 2002; EPA, date unknown).

<table>
<thead>
<tr>
<th>pH Value</th>
<th>( H^+ ) Concentration Relative to Pure Water</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>10 000 000</td>
<td>battery acid</td>
</tr>
<tr>
<td>1</td>
<td>1 000 000</td>
<td>sulfuric acid</td>
</tr>
<tr>
<td>2</td>
<td>100 000</td>
<td>lemon juice, vinegar</td>
</tr>
<tr>
<td>3</td>
<td>10 000</td>
<td>orange juice, soda</td>
</tr>
<tr>
<td>4</td>
<td>1 000</td>
<td>tomato juice, acid rain</td>
</tr>
<tr>
<td>5</td>
<td>100</td>
<td>black coffee, bananas</td>
</tr>
<tr>
<td>6</td>
<td>10</td>
<td>urine, milk</td>
</tr>
<tr>
<td>7</td>
<td>1</td>
<td>pure water</td>
</tr>
<tr>
<td>8</td>
<td>0.1</td>
<td>sea water, eggs</td>
</tr>
<tr>
<td>9</td>
<td>0.01</td>
<td>baking soda</td>
</tr>
<tr>
<td>10</td>
<td>0.001</td>
<td>Great Salt Lake, milk of magnesia</td>
</tr>
<tr>
<td>11</td>
<td>0.000 1</td>
<td>ammonia solution</td>
</tr>
<tr>
<td>12</td>
<td>0.000 01</td>
<td>soapy water</td>
</tr>
<tr>
<td>13</td>
<td>0.000 001</td>
<td>bleach, oven cleaner</td>
</tr>
<tr>
<td>14</td>
<td>0.000 000 1</td>
<td>liquid drain cleaner</td>
</tr>
</tbody>
</table>

In this project you will make your own pH paper from a colored indicator that you will extract from red cabbage by cooking it in water. Once you have the indicator solution, you can soak some coffee filter paper in it, then allow the paper to dry. When the paper is dry, you can cut it into strips, and you'll have pH paper that will change color. It will turn greenish when exposed to bases, and reddish when exposed to acids. How green or how red? That's your job! Use different solutions that you have around the house to find out how the color change corresponds to changes in pH.

Terms, Concepts and Questions to Start Background Research

To do this project, you should do research that enables you to understand the following terms and concepts:

- Acids
• Bases
• Logarithms
• pH
• pH indicators

Questions

• What value of pH is neutral?
• What range of pH values is acidic?
• What range of pH values is basic?
• What color is red cabbage pH paper when dipped in acidic solutions?
• What color is red cabbage pH paper when dipped in basic solutions?

Bibliography

• Here is two good websites about acids and bases, including information about indicators:

• This webpage has instructions for making several different colored pH indicators, including beet juice, phenolphthalein (from laxative tablets), red cabbage, and turmeric: Beckham, R., date unknown. "pH Indicators and Tests for Acids and Bases," Learn NC, School of Education, University of North Carolina, Chapel Hill [accessed July 12, 2007] http://www.learnnc.org/lessons/RichardBeckham5232002901.


• These webpages have a quick review of exponents and logarithms:

• For more information about the pH scale, try these references:
Materials and Equipment

To do this experiment you will need the following materials and equipment:

- Red cabbage leaves
- 1-quart cooking pot
- Water
- 1-quart bowl
- Strainer
- White coffee filters (cone-shaped ones are good)
  - Alternatively, you can use filter paper or chromatography paper.
- Acidic and basic solutions to test, for example:
  - Lemon juice, vinegar
  - Orange juice, soda
  - Tomato juice, acid rain
  - Black coffee, bananas
  - Milk, saliva
  - Pure water
  - Sea water, eggs
  - Baking soda solution
  - Milk of magnesia
  - Ammonia solution
  - Soapy water

Experimental Procedure

<table>
<thead>
<tr>
<th>Safety Notes:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Adult supervision required.</td>
</tr>
<tr>
<td>Do not mix strong acids and bases.</td>
</tr>
<tr>
<td>Use appropriate caution when testing the pH of household cleaning solutions (like ammonia). Avoid skin contact, and follow all precautions on the product label.</td>
</tr>
</tbody>
</table>

1. Do your background research so that you are knowledgeable about the terms, concepts, and questions, above.
2. Prepare a red cabbage indicator solution (the "Experiments with Acids and Bases" webpage (Carboni, 2004) has great pictures illustrating all of the steps)
   a. Slice a head of cabbage at approximately 3 cm (1 in) intervals, or peel the leaves from the head and tear them into pieces.
   b. Place the leaves in the cooking pot and cover with water.
c. Cook on medium heat for half an hour (low boil is good).

d. Allow the cooked cabbage to cool, then pour off the liquid into a bowl. You can pour through a strainer to catch the cabbage pieces, or hold them back with a large, flat ladle with holes—see the photographs on the "Experiments with Acids and Bases" webpage (Carboni, 2004).

e. The solution is a deep blue, but will change color when the pH changes. (You can experiment with using the liquid as a pH indicator.)

3. Here's how to make pH paper using the red cabbage solution and coffee filters:
   a. Soak the white coffee filters in the red cabbage solution for about 30 minutes.
   b. Drain the excess solution from the filters, and set them out in a single layer on some paper towels to dry overnight. To speed up the drying process, you can put them on a cookie sheet and put them in your oven at low temperature (150–200°F).
   c. When the coffee filters are dry, cut them into 3 cm × 8 cm (about 1 in × 3 in) strips.
   d. The strips are now ready to test the pH of various solutions. They start out blue, but will turn green in basic solutions and red in acidic solutions.

4. Use the strips to test the acidity/alkalinity of various solutions around your house. For example:
   a. Lemon juice, vinegar
   b. Orange juice, soda
   c. Tomato juice, acid rain
   d. Black coffee, bananas
   e. Milk, saliva
   f. Pure water
   g. Sea water, eggs
   h. Baking soda solution
   i. Milk of magnesia
   j. Ammonia solution
   k. Soapy water
   l. Note: if you test the pH of saliva, do not put the pH paper in your mouth! Instead, spit some saliva into a clean container and dip the paper into the saliva.

5. After testing, put the pH strips in order of increasing pH of the solution tested.
   a. You can use the table in the Introduction as a guide.
   b. The Variations section has some additional suggestions for independent confirmations of the pH readings.

6. Do you see a gradual change in color as the pH of the tested solutions varies? Can you match specific colors to certain pH levels? Over what range of pH does the color continue to change? How accurately do you think you can determine the pH of a solution with your test papers? Within 1, 2, or 3 pH units?

Variations
• Compare the performance of your homemade pH paper with commercial pH paper (can be found in a well-stocked tropical fish store). Or, buy an inexpensive pH meter and use it to calibrate your homemade pH paper. Use the table in the Introduction to make a series of different solutions, form low to high pH. Measure the pH of each solution with the pH meter (rinse off the tip between solutions), and write down the results. Now check each solution with your pH paper. Can you see color differences that correspond to the measured changes in pH? Over what pH range do you see color changes? How large does the shift in pH need to be in order to see a change in color?

• Try making pH indicator solutions (and/or indicator papers) from other natural dyes: for example beet juice, phenolphthalein, or turmeric powder (Beckham, date unknown; Krampf, 2006). Test your household solutions with each of the indicators. Does the additional information from multiple indicators give you a better measure of the pH of your solutions?

• Does the pH of your saliva change after eating various types of food? If so, how much time does it take to return to normal? Design an experiment to find out. Again, do not put the pH paper in your mouth. Instead, spit some saliva into a clean container and dip the pH paper into the saliva. Also, don't try changing the pH of your saliva with anything non-edible!

• What is the pH of rainwater in your area? Can you measure it with your pH paper or pH indicator solutions?
AIM- Science Projects {Chemistry}~What's the Point of Boiling?

Objective

The goal of this project is to separate pure water from fruit juice using a simple stovetop distillation apparatus.

Introduction

This project uses the technique of distillation. Distillation is when you boil a liquid, and then capture the vapor that escapes from the liquid and cool it. The cooled vapor condenses back into liquid. The condensed liquid is called the distillate. Do you think this process changes the liquid?

What if the liquid you boil has substances dissolved in it? For example, what if you started with a solution of sugar water? If you boiled the sugar water, you know from experience that there would be steam rising up from the pot on the stove. If you condensed that steam back into liquid, do you think the condensed liquid (the distillate) would contain sugar or not?

In this project, you will learn how to build a simple stove top distillation apparatus with stuff that you probably have in your kitchen right now. All you need is a deep pot with a sloping lid, a coffee cup, a bowl, some ice, and a stove. Of course, you'll also need a liquid to distill. Colored fruit juice will work fine, or you could make a solution of sugar water. Add food coloring to it if you like. The Experimental Procedure section, below, shows you how to put it all together to find out what happens.

Terms, Concepts and Questions to Start Background Research

To do this project, you should do research that enables you to understand the following terms and concepts:

- Boiling point
- Phases of matter:
  - Solid
  - Liquid
  - Vapor
- Condensation
- Solvent
- Solute
- Distillate

Questions
What happens to solute molecules when the solvent evaporates or boils?

How will the distillate compare to the original juice for:
  - color?
  - taste?
  - pH?

Bibliography

- Cbse Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

- Stove
- Deep cooking pot with sloped lid
- Ceramic coffee cup
- Ceramic bowl
- Ice
- Hot mitts
- Colored fruit juice (e.g., orange juice, grape juice, cranberry juice, etc.)

Experimental Procedure

1. Do your background research so that you are familiar with the terms, concepts, and questions, above. For more information on distillation methods.
2. The line drawing below is an illustration of the stove-top distillation apparatus used in this experiment.
3. Here are the steps for using the distillation apparatus.
   a. Pour the colored fruit juice into the bottom of the pot. Save at least 200 ml of the original juice for comparison to the distillate.
   b. Place the ceramic coffee cup, open side up, in the center of the deep pot. (That's correct, right in the juice!)
   c. Place a bowl on top of the coffee cup. (The bowl will catch the condensed liquid that drips down from the lid.)
   d. Put the cover on the pot, upside down.
   e. Put ice in the cover of the pot.
   f. Turn on the burner to medium heat. You want the juice to boil moderately (not a rolling boil).
   g. Allow the pot to boil for 10 minutes or so (enough time to collect a sufficient amount of distillate for testing).
   h. When done, turn off the burner. Allow the pot to cool for a few minutes.
   i. Put on hot mitts and carefully remove the cover from the pot.
   j. Still wearing hot mitts, lift the bowl off of the coffee cup and set it down on a heat-resistant surface.
   k. Remove the coffee cup.
   l. After it cools, pour the remaining juice from the pot into a clear container.

4. How do the original juice, the remaining juice from the pot, and the distillate compare in terms of color?

5. Ordinarily in a chemistry experiment, you would not taste any of the solutions. In this case, since you are using clean kitchen utensils, and edible fruit juice, a taste test is OK. Let the liquids cool to room temperature before tasting them! How do the three different liquids compare for taste?
   a. Which liquid is sweetest?
   b. Which is least sweet?
   c. You should be able to explain why.
Objective
The goal of this project is to measure the effect of reactant particle size on the rate of a chemical reaction.

Introduction
You may have seen a television commercial for Alka-Seltzer tablets, or heard one of their advertising slogans: "Plop, plop, fizz, fizz, oh what a relief it is!®" When you drop the tablets in water, they make a lot of bubbles, like an extra-fizzy soda. And like a soda, the bubbles are carbon dioxide gas (CO$_2$). However, with Alka-Seltzer®, the CO$_2$ is produced by a chemical reaction that occurs when the tablets dissolve in water.

The main ingredients of Alka-Seltzer tablets are aspirin, citric acid, and sodium bicarbonate (NaHCO$_3$). When sodium bicarbonate dissolves in water, it dissociates (splits apart) into sodium (Na$^+$) and bicarbonate (HCO$_3^-$) ions. The bicarbonate reacts with hydrogen ions (H$^+$) from the citric acid to form carbon dioxide and water. The reaction is described by the following chemical equation:

$$3\text{HCO}_3^- + 3\text{H}^+ \rightarrow 3\text{H}_2\text{O} + 3\text{CO}_2$$

So how does particle size come into this? In order for the reaction shown above to take place, the ingredients in the tablet first have to dissolve. The tablet has a large surface area, so this step should be pretty fast, right? What effect do you think particle size will have on the speed of the bicarbonate reaction? You can find out for yourself by plopping prepared Alka-Seltzer® tablets (whole tablets, halved tablets, quartered tablets, and powdered tablets) into water at the same temperature, and timing how long it takes for the chemical reaction to go to completion.

Terms, Concepts and Questions to Start Background Research
To do this project, you should do research that enables you to understand the following terms and concepts:

- Molecules
- Temperature
- Reactants
- Products
- Reaction rate

Questions
Do you think changing the particle size will have a measurable effect on the chemical reaction rate?
Will smaller particles speed up or slow down the reaction?

Bibliography

- Cbse Blog: http://cbse-sample-papers.blogspot.com

Materials and Equipment

To do this experiment you will need the following materials and equipment:

- At least 12 Alka-Seltzer® tablets (if you plan to do additional variations to the project, you'll want to get a larger box)
- Sheet of blank paper
- Hammer
- Piece of scrap wood
- Thermometer (good range would be -10°C to 110°C
  - E.g. catalog # WW6332000 from Science Kit & Boreal Lab or catalog #15V1460 from Wards Natural Science)
  - Standard kitchen candy thermometer will also work fine
- Clear 12 ounce (355 mL) drinking glass (or larger)
  - Note: Use Pyrex glass when working with water heated on the stove or in the microwave)
- Measuring cup
- Masking tape
- Something to stir with (a teaspoon or a chopstick, for example)
- Tap water
- Stop watch (you can also use a clock or watch with a second hand)
- A helper
- Lab notebook
- Pencil

Experimental Procedure

1. Do your background research and make sure that you are familiar with the terms, concepts, and questions, above.
2. In this experiment, you will be measuring the time it takes for one Alka-Seltzer® tablet to react completely in water. You will investigate how the reaction time changes as you vary the particle size of the reactants.
3. You'll use the same glass for repeated trials, so it is convenient to mark the desired water level.
a. Use the measuring cup to add 8 ounces (236 mL) of water to
the glass. (If you’re using metric volume units, rounding up to
250 mL is fine.)
b. Use a piece of masking tape on the outside of the glass to
mark the water level. Place the tape with its top edge even
with the water level in the glass.
c. Now you can use the masking tape to fill the glass to the right
level for each trial.

4. For observing the reaction, you will use the same volume of water at
the same starting temperature. The only variable that you should
change is the particle size of the tablets. You will use four different
particle sizes for the Alka-Seltzer® tablets:
a. A whole tablet
b. A tablet broken in half
c. A tablet broken in quarters
d. A tablet ground into powder. To do this, fold a single tablet to
be ground inside a clean piece of paper. Place the folded
paper on a piece of scrap wood, and use the hammer to firmly
pound the tablet about ten times. Stop immediately if the paper
shows signs of tearing: you don’t want to lose any of the
powder.

5. Here is how to measure the reaction time:
a. Fill the glass with water to the level of the masking tape.
b. Measure the temperature of the water, and record it in your lab
notebook. Each trial should be carried out at the same
temperature, so adjust the water temperature (by adding warm
or cold water) as necessary.
c. Remove the thermometer. (It’s not a good idea to use the
thermometer as a stirring rod. It might break.)
d. Have your helper get ready with the stop watch, while you get
ready with an Alka-Seltzer®. Have your helper count one—
two—three. On three, the helper starts the stop watch and you
drop the tablet (or tablet pieces) into the water.
e. You’ll immediately see bubbles of CO₂ streaming out from the
tablet.
f. Stir the water gently and steadily. Use the same stirring
method and speed for all of your experimental trials. The tablet
will gradually disintegrate. Watch for all of the solid white
material from the tablet to disappear.
g. When the solid material has completely disappeared and the
bubbles have stopped forming, say "Stop!" to have your helper
stop the stopwatch.
h. Record the reaction time in your lab notebook.
i. Tip: be careful when opening the packets and handling the
Alka-Seltzer® tablets. The tablets are thin and brittle, so they
break easily. You need to have four whole tablets for this
experiment.

6. For each of the four particle sizes, you should repeat the experiment
three times, for a total of 12 trials. You can organize your data in a
table like the one below.
<table>
<thead>
<tr>
<th>Particle Size</th>
<th>Temperature (°C)</th>
<th>Reaction Time (s)</th>
<th>Average Reaction Time (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Trial #1</td>
<td>Trial #2</td>
</tr>
<tr>
<td>Whole Tablet</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tablet Broken in Half</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tablet Broken in Quarters</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Powdered Tablet</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

7. Calculate the average reaction time for each of the four particle sizes.
8. Make a bar graph showing the average reaction time, in seconds, (y-axis) vs. particle size (x-axis).
9. How does reaction time change with particle size?
Objective
The objective of this project is to use paper chromatography to analyze ink components in permanent black markers.

Introduction
Matter makes up everything in the universe. Our body, the stars, computers, and coffee mugs are all made of matter. There are three different types of matter: solid, liquid, and gas. A solid is something that is normally hard (your bones, the floor under your feet, etc.), but it can also be powdery, like sugar or flour. Solids are substances that are rigid and have definite shapes. Liquids flow and assume the shape of their container; they are also difficult to compress (a powder can take the same shape as its container, but it is a collection of solids that are very small). Examples of liquids are milk, orange juice, water, and vegetable oil. Gases are around you all the time, but you may not be able to see them. The air we breathe is made up of a mixture of gases. The steam from boiling water is water's gaseous form. Gases can occupy all the parts of a container (they expand to fill their containers), and they are easily compressed.

Matter is often a mixture of different substances. A heterogeneous mixture is when the mixture is made up of parts that are dissimilar (sand is a heterogeneous mixture). Homogeneous mixtures (also called solutions) are uniform in structure (milk is a homogeneous mixture). A sugar cube floating in water is a heterogeneous mixture, whereas sugar dissolved in water is a homogeneous mixture. You will determine whether the ink contained in a marker is a heterogeneous or homogeneous mixture, or just one compound.

In a mixture, the substance dissolved in another substance is called the solute. The substance doing the dissolving is called the solvent. If you dissolve sugar in water, the sugar is the solute and the water is the solvent.

For this project, you will be making a small spot with an ink marker onto a strip of paper. The bottom of this strip will then be placed in a dish of water, and the water will soak up into the paper.

The water (solvent) is the mobile phase of the chromatography system, whereas the paper is the stationary phase. These two phases are the basic principles of chromatography. Chromatography works by something called capillary action. The attraction of the water to the paper (adhesion force) is larger than the attraction of the water to itself (cohesion force), hence the water moves up the paper. The ink will also be attracted to the paper, to itself, and to the water differently, and
thus a different component will move a different distance depending upon the strength of attraction to each of these objects. As an analogy, let's pretend you are at a family reunion. You enjoy giving people hugs and talking with your relatives, but your cousin does not. As you make your way to the door to leave, you give a hug to every one of your relatives, and your cousin just says "bye." So, your cousin will make it to the door more quickly than you will. You are more attracted to your relatives, just as some chemical samples may be more attracted to the paper than the solvent, and thus will not move up the solid phase as quickly. Your cousin is more attracted to the idea of leaving, which is like the solvent (the mobile phase).

Chromatography is used in many different industries and labs. The police and other investigators use chromatography to identify clues at a crime scene like blood, ink, or drugs. More accurate chromatography in combination with expensive equipment is used to make sure a food company's processes are working correctly and they are creating the right product. This type of chromatography works the same way as regular chromatography, but a scanner system in conjunction with a computer can be used to identify the different chemicals and their amounts. Chemists use chromatography in labs to track the progress of a reaction. By looking at the sample spots on the chromatography plate, they can easily find out when the products start to form and when the reactants have been used up (i.e., when the reaction is complete). Chemists and biologists also use chromatography to identify the compounds present in a sample, such as plants.

Terms, Concepts and Questions to Start Background Research

- adhesion, cohesion forces
- capillary action
- stationary phase, mobile phase
- hydrophilic, hydrophobic
- R_f value
- paper chromatography
- solvent
- solution

Questions

- Why do different compounds travel different distances on the piece of paper?
- How is an R_f value useful?
- What is chromatography used for?

Bibliography

- Basic chemistry concepts: [http://www.chem4kids.com](http://www.chem4kids.com)
Materials and Equipment

- water
- at least 15 identically sized strips of paper (5 for each pen)
  Note: chromatography paper or laboratory filter paper is preferable, but you can use a paper towel. The problem with paper towels is that they may be too absorptive and smear the ink. For more information on which papers work and which don't.
- ruler
- pencils
- at least three different types of black markers (including one permanent marker), or at least three different colors of marker (including one permanent marker)
- a wide-mouth jar for the solvent

Experimental Procedure

Note: To make sure you can compare your results, as many of your materials as possible should remain constant. This means that the temperature, type of water used, size of paper strips, where the ink is placed onto the paper etc. should remain the same throughout the experiment.

1. Cut paper strips about one by four inches in area (they must all be the same size).
2. Take one of the paper strips and use a ruler and pencil to draw a line across it horizontally two cm from the bottom. This is the origin line (see illustration, below).
3. Pour a small amount of water into your glass (there should be barely enough for the paper strip to hang inside of the jar and just touch the water).
4. Using one of the markers, place a small dot of ink onto the line (see illustration, above).
5. Use the pencil to label the strip, so that you know which marker it represents.
6. Tape the paper to a pencil and hang it into the jar of solvent so that the bottom edge is just barely touching (see illustration, below).

![Paper Strip in Jar](image)

7. Let the water rise up the strip until it is almost at the top.
8. Remove the strip from the jar and mark how far the solvent rose with a pencil.
9. Analyze the ink component(s):
   - Measure the distance the solvent and each ink component traveled from the starting position, then calculate the $R_f$ value for each component (some of the ink components might not have moved at all!).
10. Repeat this experiment for each brand or color of marker five times.

Questions

- Did the different inks separate differently? By looking at the $R_f$ values, can you tell if any of the ink components from the different markers are the same?
- If the ink components separated differently for each marker, why did this happen (think about the strength of attractions)?
AIM-DETERMINE THE MASS OF ALUM CRYSTALS

INTRODUCTION

Aluminium because of its low density, high tensile strength and resistance to corrosion is widely used for the manufacture of airplanes, automobiles, lawn furniture as well as for aluminium cans. Being a good conductor of electricity, it is also used for the transmission of electricity. Aluminium foil is used for wrapping cigarettes, confectionery items, etc. Aluminium is also used for making utensils. The recycling of aluminium cans and other aluminium products is a very positive contribution to saving our natural resources. Most of the recycled aluminium is melted and recast into other aluminium metal products or used in the production of various aluminium compounds, the most common of which are the alums. Alums are double sulphates having general formula $X_2SO_4\cdot M_2(SO_4)_3 \cdot 24H_2O$ where,

$X =$ monovalent cation such as Na$^+$, K$^+$, NH$_4^+$, etc.

$M =$ trivalent cation such as Al$^{+3}$, Cr$^{+3}$, Fe$^{+3}$, etc.

Some important alums and their names are given below:

Potash Alum: $K_2SO_4\cdot Al_2(SO_4)_3 \cdot 24H_2O$

Soda Alum $Na_2SO_4\cdot Al_2(SO_4)_3 \cdot 24H_2O$

Chrome Alum $K_2SO_4\cdot Cr_2(SO_4)_3 \cdot 24H_2O$

Ferric Alum $(NH_4)_2SO_4\cdot Fe_2(SO_4)_3 \cdot 24H_2O$

Alums are isomorphous crystalline solids, which are soluble in water. Potash alum is used in paper making, in fire extinguisher, in foodstuffs and in purification of water. Soda alum is used in baking powders and chrome alum is used in tanning leather and waterproofing fabrics. Ferric alum is used in antiseptics. The shape of a potash alum crystal is octahedral.
REQUIREMENTS

APPARATUS:
- Conical flasks
- Filter paper
- Piece of aluminium foil
- Burner
- Funnel

CHEMICALS:
- Potassium Hydroxide (KOH)
- Sulphuric Acid (H₂SO₄)
- Ethanol (C₂H₅OH)

THEORY

1. Aluminium metal is treated with hot aqueous KOH solution. Aluminium dissolves as potassium aluminate, KAl(OH)₄, salt.

\[2\text{Al(s)} + 2\text{KOH(aq)} + 6\text{H}_2\text{O(l)} \rightarrow 2\text{KAl(OH)}_4(aq) + 3\text{H}_2(g)\]

2. Potassium aluminate solution on treatment with dil. sulphuric acid first gives ppt. of Al(OH)₃, which dissolves on addition of small excess of H₂SO₄ and heating.

\[2\text{KAl(OH)}_4(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow 2\text{Al(OH)}_3(s) + \text{K}_2\text{SO}_4(aq) + 2\text{H}_2\text{O(l)}\]

\[2\text{Al(OH)}_3(s) + 3\text{H}_2\text{SO}_4(aq) \rightarrow \text{Al}_2(\text{SO}_4)_3(aq) + 6\text{H}_2\text{O(l)}\]

3. The resulting solution is concentrated to near saturation and cooled. On cooling crystals of potash alum crystallize out.

\[\text{K}_2\text{SO}_4(aq) + \text{Al}_2(\text{SO}_4)_3(aq) + 24\text{H}_2\text{O(l)} \rightarrow \text{K}_2\text{SO}_4.\text{Al}_2(\text{SO}_4)_3.24\text{H}_2\text{O(s)}\]

PROCEDURE

1. Prepare 50ml of 4M KOH solution.

2. Add small pieces of aluminium foil (about 1 gm) in the conical flask containing the KOH solution. Since during this step hydrogen gas is evolved, this step must be done in a well-ventilated area.
3. After all of aluminium has reacted, filter the solution to remove any insoluble impurities.

4. Allow the filtrate to cool. Now add slowly conc. H₂SO₄ until insoluble Al(OH)₃ just forms in the solution.

5. Gently heat the mixture until the Al(OH)₃ ppt. dissolves. Leave the solution overnight for the crystallization to continue.

6. Take out the crystals and wash them with 50/50 ethanol-water mixture.

7. Determine the mass of the alum crystals.

**OBSERVATIONS**

- Mass of aluminium metal = 1 gm
- Mass of potash alum = 13 gm
- Colour = White
- Shape of crystals = Octahedral

**PRECAUTIONS**

1. A few drops of conc. sulphuric acid should be added while preparing saturated solution of aluminium sulphate to prevent its hydrolysis.

2. Aluminium sulphate solution should be clear and not turbid.

3. Cool the conc. solution slowly to get large crystals. Rapid disturbance of solution may change the shape, size and quantity of the crystals.

4. Conc. solution should be cooled undisturbed. A slight disturbance of solution may change the shape, size.
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AIM-TO STUDY THE RATE OF EVAPORATION OF DIFFERENT LIQUIDS

INTRODUCTION

When a liquid is placed in an open vessel, it slowly escapes into gas phase, eventually leaving the vessel empty. This phenomenon is known as evaporation. Evaporation of liquids can be explained in terms of kinetic molecular model. Although there are strong intermolecular attractive forces which hold molecules of a liquid together, the molecules having sufficient kinetic energy can escape into gas phase if such molecules happen to come near the surface. In a sample of liquid all the molecules do not have same kinetic energy. There is a small fraction of molecules which have enough kinetic energy to overcome the attractive forces and escape into gas phase.

Evaporation causes cooling. This is due to the reason that the molecules, which undergo evaporation, are high-energy molecules; therefore the kinetic energy of molecules which are left behind is less. Since the remaining molecules have lower average kinetic energy therefore, temperature must be lower. If the temperature is kept constant the remaining liquid will have the same distribution of molecular kinetic energies and the high-energy molecule will keep on escaping from the liquid into the gas phase. If the liquid is taken in an open vessel, evaporation will continue until whole of the liquid evaporates.

REQUIREMENTS

APPARATUS:

- THREE PETRIDIISHES OF DIAMETER 10 CM WITH COVERS
- 10 ML PIPETTE
- STOP WATCH

CHEMICALS:

- ACETONE
- BENZENE
- CHLOROFORM
PROCEDURE

- Clean and dry the petridishes and mark them as A, B, C.
- Pipette out 10 ml of acetone to petridish A and cover it.
- Pipette out 10 ml of benzene in petridish B and cover it.
- Pipette out 10 ml of chloroform in petridish C and cover it.
- Uncover all the three petridishes simultaneously and start the stopwatch.
- Note the respective time when the liquids evaporate completely from each petridish.

OBSERVATIONS

<table>
<thead>
<tr>
<th>Petridish Mark</th>
<th>Liquid Taken</th>
<th>Time taken for complete evaporation</th>
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<tr>
<td>A</td>
<td>Acetone</td>
<td>53 min</td>
</tr>
<tr>
<td>B</td>
<td>Benzene</td>
<td>42 min</td>
</tr>
<tr>
<td>C</td>
<td>Chloroform</td>
<td>30 min</td>
</tr>
</tbody>
</table>

CONCLUSION

The rate of evaporation of the given three liquids is in the order:

Chloroform > Benzene > Acetone
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AIM
To analyse some fruits & vegetables juice for the contents present in them.

INTRODUCTION
Fruits and vegetable are always a part of balanced diet. That means fruits vegetables provide our body the essential nutrients, i.e. carbohydrates, proteins, vitamins and minerals. Again their presence in these is being indicated by some of our general observations, like freshly cut apples become reddish black after some time. Explanation for it is that iron present in apple gets oxidixed to iron oxide. So, we can conclude that fruits and vegetables contain complex organic compounds, for e.g., anthocin, chlorophyll, esters(flavouring compounds), carbohydrates, vitamins and can be tested in any fruits or vegetable by extracting out its juice and then subtracting it to various tests which are for detection of different classes of organic compounds. Detection of minerals in vegetables or fruits means detection of elements other than carbon, hydrogen and oxygen.

MATERIAL REQUIRED
- Test Tubes
- Burner
- Litmus paper
- Laboratory reagents
- Various fruits
- Vegetables juices

CHEMICAL REQUIREMENTS
- pH indicator
- Iodine solution
- Fehling solution A and Fehling solution B
- Ammonium chloride solution
- Ammonium hydroxide
- Ammonium oxalate
- Potassium sulphocynaide solution

PROCEDURE
The juices are made dilute by adding distilled water to it, in order to remove colour and to make it colourless so that colour change can be easily watched and noted down. Now test for food components are taken down with the solution.

**TEST, OBSERVATION & INFERENCE**

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<tr>
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<tr>
<td><strong>Test for acidity:</strong></td>
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<tr>
<td>Take 5ml of orange juice in a test tube and dip a pH paper in it. If pH is less than 7 the juice is acidic else the juice is basic.</td>
<td>The pH comes out to be 6.</td>
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<tr>
<td><strong>Test for Starch:</strong></td>
<td></td>
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<tr>
<td>Take 2 ml of juice in a test tube and add few drops of iodine solution. It turns blue black in colour than the starch is present.</td>
<td>Absence of blue black in colour.</td>
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<tr>
<td><strong>Test for Carbohydrates (FEHLING’S TEST):</strong></td>
<td></td>
<td></td>
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<tr>
<td>Take 2 ml of juice and 1 ml of fehling solution A &amp; B and boil it. Red precipitates indicates the presence of producing sugar like maltose, glucose, fructose &amp; Lactose.</td>
<td>No red coloured precipitates obtained.</td>
<td>Carbohydrates absent.</td>
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<tr>
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<td>Yellow precipitate is obtained.</td>
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</tr>
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**CONCLUSION**

From the table given behind it can be conducted that most of the fruits & vegetable contain carbohydrate & vegetable contain carbohydrate to a small extent. Proteins are present in small quantity. Therefore one must not only depend on fruits and vegetables for a balance diet.
AIM-ANALYSIS OF VEGETABLES AND FRUIT JUICES

AIM

To analyse some fruits & vegetables juice for the contents present in them.

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<tr>
<td><strong>Test for Starch:</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Take 2 ml of juice in a test tube and add few drops of iodine solution. It turns blue black in colour than the starch is present.</td>
<td>Absence of blue black in colour.</td>
<td>Orange juice is acidic.</td>
</tr>
<tr>
<td><strong>Test for Carbohydrates (FEHLING’S TEST):</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Take 2 ml of juice and 1 ml of fehling solution A &amp; B and boil it. Red precipitates indicates the presence of producing sugar like maltose, glucose, fructose &amp; Lactose.</td>
<td>No red coloured precipitates obtained.</td>
<td>Carbohydrates absent.</td>
</tr>
<tr>
<td><strong>Test for Iron:</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Test for Calcium:</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Take 2 ml of juice add Ammonium chloride and ammonium hydroxide solution. Filter the solution and to the filtrate add 2 ml of Ammonium Oxalate solution. White ppt or milkiness indicates the presence of calcium.</td>
<td>Yellow precipitate is obtained.</td>
<td>Calcium is present.</td>
</tr>
</tbody>
</table>

**CONCLUSION**

From the table given behind it can be conducted that most of the fruits & vegetable contain carbohydrate & vegetable contain carbohydrate to a small extent. Proteins are present in small quantity. Therefore one must not only depend on fruits and vegetables for a balance diet.