**Classic Acid-Base Titration**

**Modified for AP Environmental Science “Acid Rain” demo**

**Materials\***

Each lab station should be set up with the following materials-

50 mL beaker

250 mL beaker

250 mL Erlenmeyer flask

100 mL graduated cylinder

pH meter

PASCO pH PC-link (to connect meter to your computer)

1 50-mL Buret

Ring stand and buret clamp

Wash bottle

Chalk

Mortar and pestle

Plastic pipets

Straws

General classroom supplies-

Weighing papers

Potassium hydrogen phthalate (KHP) powder

Sodium Hydroxide (NaOH) solution

Scales

0.05 M sulfuric acid (“concentrated acid rain”)

Distilled H2O

\*At the end of the lab, please make sure all materials are organized as you found them and your lab station is cleaned up. Used plastic pipets and straws may be thrown away.

**Safety Information**

 The materials you will use in today’s lab are flammable and corrosive and can damage your clothes, skin, and eyes. Do not, for any reason, light a match in the lab room. You MUST wear goggles and gloves at all times in the lab. Hair must be pulled back and sleeves rolled up. After you have handled the chemicals we will use today, take special precaution not to touch your face area or clothes while you are still wearing your gloves. Even a trace amount of chemical could be corrosive to your skin or cause holes in your clothes. At the end of the lab, please dispose of gloves and clean your goggle glasses as part of the lab cleanup. Lab benches also need to be thoroughly wiped down with soap and water. Thanks!

**Part I-** Simulating Carbonic Acid Formation

*Background*

 Atmospheric carbon (carbon dioxide) has the ability to react with water and create a weak acid, carbonic acid. While this acid is not as severely effecting as other components of acid rain (for example, sulfuric acid), it is still considered a potential acid rain hazard. In the below demonstration, you will simulate carbonic acid formation.

*Directions*

1. Wash all beaker glassware at your station.
2. Send someone to obtain distilled H2O in your washbottle. From this bottle, add 100 mL of distilled water to the 250 mL beaker.
3. Connect the PASCO sensor to your computer and use the *DataStudio* program to obtain the initial pH of the water.
4. Leave the sensor in the beaker and begin to blow bubbles into the water. For the next 5 minutes, record the pH reading at 1-minute intervals (use a stopwatch online @ <http://www.online-stopwatch.com/>). You may switch blowers with new straws if you wish throughout the time period but try to maintain a constant stream of bubbles in the water.
5. Record your data in the table below.

|  |  |
| --- | --- |
| Time Since Start | Measured pH |
| 0 (initial pH) |  |
| 1 min |  |
| 2 min |  |
| 3 min |  |
| 4 min |  |
| 5 min (final pH) |  |

1. Write the chemical formula represented by this experiment.
2. Clean all used glassware.

**Part II-** Acid-Base Titration

*Background*

 Titrations are typically used in chemistry to determine the amount of acid or base in a sample solution. Usually, a known volume of acid would be “titrated” by slowly adding dropwise a standard solution of a strong base (the standard is one whose concentration is accurately known). The titrant, or the sample whose concentration is unknown, reacts with the standard via a neutralization reaction. In a neutralization reaction, an acid and a base recombine to form water and another neutral product. The exact volume of base needed to neutralize the acid is measured, and once this volume is reached, the reaction is at the equivalence point and can be considered neutral. Indicators are usually added to acid-base titrations to detect the equivalence point. The endpoint of the titration is the point at which the indicator changes color and signals that the equivalence point has indeed been reached. The pH at the equivalence point will be exactly 7, so an indicator that changes color at pH 7 will help us know when the two clear liquids have neutralized each other. In this reaction, we will use phenolphthalein, which changes to a pink color when the solution reaches a pH of 7.

 The progress of an acid-base titration (neutralization) can be followed by measuring the pH of the solution being analyzed as a function of the volume of the titrant added (this means the volume of the titrant is the *independent* variable and the pH is the *dependent* variable on your graph). A graphed plot of the resulting data is called a pH curve or titration curve. Titration curves can be referenced for a precise determination of the equivalence point without the use of an indicator.

*Directions*

1. On an analytical balance, accurately weigh 0.4 to 0.6 grams of KHP on previously tared weigh paper. Record the mass of the KHP in the data table section (attached).
2. Transfer the KHP into your Erlenmeyer flask.
3. Fill your graduated cylinder with 40-mL of distilled water. Use this water to rinse any solid that stuck to the flask into the bottom of the flask. Pour the remainder of the 40 mL into the Erlenmeyer so you have a total of 40 mL + KHP in your flask.
4. Check to make sure your buret titration colum is correctly set up.
5. Place your 40 mL beaker under the buret tip. Briefly open the buret stopcock to allow any air bubbles to escape from the tip. Close the stopcock. Remove your 40 mL beaker you were using to catch the excess NaOH.
6. Measure the precise volume of the solution in the buret and record this value in the data table section. Note- volumes are read from the top down in a buret. Always read from the bottom of the meniscus, including 2 decimal places (see diagram).



1. Position the buret over the Erlenmeyer flask so that the tip of the buret is within the flask but at least 2 cm above the liquid surface.
2. Add 3 drops of phenolphthalein solution the KHP solution in the flask. Measure the pH and record this value in provided table the data section (this is the initial pH)- see basic diagram below. Remove the pH sensor from the liquid.



1. Begin the titration by adding 1 mL of the NaOH solution the Erlenmeyer flask, then closing the buret stopcock and swirling the flask. Measure the pH and record. For the remainder of the titration, continuously swirl as you add the NaOH.
2. Repeat step 11 until 3 mL of NaOH has been added to the flask. Be sure to continuously swirl the flask.
3. Reduce the incremental volumes of NaOH solution to 0.1 mL until the pink color starts to persist. Take the pH after each addition. Reduce the rate of addition of NaOH solution to drop by drop until the pink color persists for 15 seconds. Take the pH again once the pink color forms (you do not need to measure it at the end of each drop). Remember to constantly swirl the flask. **Be very cautious at this point. If you over-titrate, you must start over.**
4. Measure the volume of NaOH remaining in the buret, estimating to the nearest 0.01mL. Record this value as the “final reading” in the data section.
5. Dispose of your titrated solution in the sink and clean your flask and other used glassware. Leftover NaOH may remain in the buret with the stopcock sealed.
6. Calculate the total volume of NaOH used by subtracting your final reading from your initial reading. This is the amount of NaOH needed to neutralize your acid.

**Part III-** Showing the effects of acid rain on human-made structures & Tennessee bedrock.

*Background*

 The approximate average rainfall pH in middle Tennessee is 4.8. This is deemed ‘acidic’ by acid deposition standards, which set the maximum pH to be considered “acid rain” at ≤5.0. While the pH of the acid you will use below is approx. 1-2, it will quickly simulate the long-term effects of a weaker acid on structures containing calcium carbonate (limestone, shells, chalks, etc). As you know, the predominant bedrock here in middle Tennessee is limestone and many local structures are built with this material.

*Directions*

1. Using a mortar and pestle, crush half of a piece of chalk into small pieces. The chalk should be crushed, not finely ground or powdered.
2. Place the chalk in a small beaker.
3. Using the graduated cylinder, measure and pour 25 mL of 0.05M sulfuric acid (“concentrated acid rain”) into the beaker.
4. Observe the reaction within the beaker for 5 minutes.
5. Record your observations below.
6. Pour contents of beaker into the central waste beaker. Clean your beaker by initially flushing with water and then washing using traditional cleaning technique.

**Data Recording Section**

**Mass of KHP \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**Start value of NaOH in buret (initial reading) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

**End value of NaOH in buret @ equivalence point (final reading) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_**

|  |  |
| --- | --- |
| mL NaOH added to KHP | pH |
| 0 (initial pH) |  |
| 1 |  |
| 2 |  |
| 3 |  |
| 3.1 |  |
| 3.2 |  |
| 3.3 |  |
| 3.4 |  |
| 3.5 |  |
| 3.6 |  |
| 3.7 |  |
| 3.8 |  |
| 3.9 |  |
| 4.0 |  |
|  |  |
|  |  |
|  |  |
|  |  |
|  |  |