# Pre-Lab Information

Purpose Explore the concept of pH and how to determine if a substance is acidic or basic using a laboratory procedure.

Time Approximately 60 minutes

Question How is pH used to determine if a solution is acidic or basic?

Hypothesis If a solution is an acid then it will have a pH below 7, and if the solution is a base if will have a pH above 7, since the pH value is determined by the concentration of H+ and OH- ions.

Variables Independent variable: amount of H+ and OH– ions in each solution

 Dependent variable: pH level of each substance

Summary In Part I of this experiment, students will use dried beans to model how ion concentrations relate to the pH scale. In Part II, students will make serial dilutions of laundry detergent and vinegar to determine the effect of these dilutions on the pH of the solution. In Part III, students will use a natural pH indicator to establish a reference scale for pH. Finally, in Part IV, students will test the pH of common household solutions using the natural pH indicator as well as commercial indicator strips.

# Safety

* Always wear a lab coat and safety goggles when performing an experiment.
* Behavior in the lab needs to be purposeful. Exercise caution when using the hot plate, as it can cause burns.
* Label your glassware and paper cups clearly and appropriately. Many chemicals appear to be the same.
* Check glassware, such as beakers and test tubes, for cracks and chips prior to use.
* Wear the right gear, such as chemically resistant gloves and oven mitts, when performing the experiment.
* Report all accidents—no matter how big or small—to your teacher.

# Lab Procedure

1. **Gather materials.**

|  |  |  |
| --- | --- | --- |
| * Dried black beans
* Dried white navy beans
* Funnel
* Distilled water
* Vinegar
* Ammonia
* Mass balance
* 13 commercial pH indicator strips with reference chart
 | * 20 g red cabbage
* Thermometer
* Hot plate
* Oven mitt or hot pad
* Glass stirring rod
* 1 filter paper
* 250 mL Erlenmeyer flask
* Eight 250 mL beakers
* Seven 100 mL beakers
 | * 10 mL graduated cylinder
* 13 small paper cups
* Bleach
* Glass cleaner (clear or colorless)
* Lemon juice
* Milk
* Soda pop (colorless)
* Baking soda solution
 |

**PART I: Demonstration of H+ and OH– Concentration with pH**

Dried black beans will represent H+ ions.

 Dried white navy beans will represent OH– ions.

1. **Create a model solution of pH = 7 with equal concentrations of H+ and OH**–**.**

*\*Note that the table is arranged with pH = 7 in the middle of the three rows.*

* 1. Count out 50 black beans and 50 white navy beans and add them to a 100 mL beaker. Cover the beaker, and carefully shake the beans to mix them.
	2. Draw a picture of what you see in the appropriate row (pH = 7) of Table A and record the approximate ratio of black beans (H+ ions) to white navy beans (OH– ions) by dividing the number of black beans by the number of white navy beans. This is the number of H+ ions per ion of OH– at pH 7.
	3. Consult Table B, which provides the relative concentrations of H+ and OH– ions at specific pHs. Compare the concentrations of H+ and OH– at pH 7. Divide the H+ concentration by the OH– concentration. Record this calculated ratio in Table A next to your approximated ratio. Compare your observed or approximated ratio with the calculated ratio. Are they the same or different? Why? Record your comparison in Table A.
1. **Create a model solution of pH = 6 with more H+ ions than OH– ions.**
	1. Empty the beaker from the previous step. Count out 100 black beans and add them to the beaker. Then add 1 white navy bean to the beaker.
	2. Draw a picture of what you see in Table A and use your picture to visually approximate the ratio of black beans (H+ ions) to white navy beans (OH– ions). Record this ratio in Table A.
	3. Consult Table B, and compare the concentration of each ion at pH 6. Divide the H+ concentration by the OH– concentration. Record this calculated ratio in Table A next to your approximated ratio. Compare your observed or approximated ratio with the calculated ratio. Are they the same or different? Why? Record your comparison in Table A.
2. **Create a model solution of pH = 8 with more OH– ions than H+ ions.**
	1. Empty the beaker from the previous step. Count out 100 white navy beans and add them to the beaker. Then add 1 black bean to the beaker.
	2. Draw a picture of what you see in Table A and use your picture to visually approximate the ratio of black beans (H+ ions) to white navy beans (OH– ions). Record this ratio in Table A.
	3. Consult Table B, and compare the concentration of each ion at pH 8. Divide the H+ concentration by the OH– concentration. Record this calculated ratio in Table A next to your approximated ratio. Compare your observed or approximated ratio with the calculated ratio. Are they the same or different? Why? Record your comparison in Table A.
3. **Calculate H+/OH– ratios for more extreme pH solutions.**
	1. Refer to Table B and answer the following questions. Record all answers in Table C before moving on to Part II of the lab.
	2. Find pH = 2 in Table B.
		* 1. Record the concentration of H+ and OH– ions in Table C. Divide the concentration of H+ by the concentration of OH–. How many times more H+ ions does a pH = 2 solution have than OH– ions?
			2. How many more times H+ ions are there in a pH = 2 solution than in a pH = 6 solution?
	3. Find pH = 11 in Table B.
		* 1. Record the concentration of H+ and OH– ions in Table C. Divide the concentration of H+ by the concentration of OH–. How many times fewer H+ ions does a pH = 11 solution have than OH– ions?
			2. How many times fewer H+ ions are there in a pH = 11 solution than in a pH = 6 solution?

**PART II: Determining the Effect of Dilution on pH**

1. **Establish a reference point for the pH scale.**
	1. Pour 100 mL of distilled water into a 250 mL beaker labeled “distilled water.”
	2. Predict the pH of the distilled water and record it in Table D. Note that the column for water is in the middle of the table as your reference point.
	3. Dip an indicator strip into the water, and record the resulting color in Table D.
	4. Compare the color to the indicator strip’s color key to determine the pH range (e.g., 5–6) and record.
2. **Make serial dilutions of a common acid.**
	1. Pour 100 mL of lemon juice into a 250 mL beaker labeled “A1” for Acid # 1.
	2. Using a 10 mL graduated cylinder, take 1 mL of the solution in beaker A1 and add it to 99 mL of distilled water in a 250 mL beaker labeled “A2” for Acid #2. Stir well with a clean glass stirring rod. Rinse the graduated cylinder well with water.
	3. Using the 10 mL graduated cylinder, take 1 mL of the solution in beaker A2 and add it to 99 mL of distilled water in a 250 mL beaker labeled “A3” for Acid #3. Stir well with a clean glass stirring rod. Rinse the graduated cylinder well with water.
3. **Measure the pH of the full-strength lemon juice.**
	1. Predict the pH range of the lemon juice and record this in Table D.
	2. Dip an indicator strip into beaker A1, and record the resulting color in Table D.
	3. Compare the color to the indicator strip’s color key to determine the pH range (e.g., 5–6) and record.
4. **Test predictions about the pH of the other two acidic solutions.**
	1. Using what you know from your models and the information provided in Table B, predict what the pH will be for solution A2 and solution A3. Record these predictions in Table D.
	2. Dip an indicator strip into beaker A2, and record the resulting color in Table D.
	3. Compare the color to the indicator strip’s color key to determine the pH range (e.g., 5–6) and record.
	4. Repeat b and c for solution A3.
	5. Compare your data to your predictions in Table D and record any observations that might explain any differences.
5. **Make serial dilutions of a common base.**
	1. Pour 100 mL of bleach into a 250 mL beaker labeled “B1” for Base # 1. Be especially careful with full-strength bleach and clean up any drips.
	2. Using a 10 mL graduated cylinder, take 1 mL of the solution in beaker B1 and add it to 99 mL of distilled water in a 250 mL beaker labeled “B2” for Base #2. Stir well with a clean glass stirring rod. Rinse the graduated cylinder well with water.
	3. Using the 10 mL graduated cylinder, take 1 mL of the solution in beaker B2 and add it to 99 mL of distilled water in a 250 mL beaker labeled “B3” for Base #3. Stir well with a clean glass stirring rod. Rinse the graduated cylinder well with water.
6. **Measure the pH of the full-strength bleach.**
	1. Predict the pH range of the bleach and record this in Table D.
	2. Dip an indicator strip into beaker B1, and record the resulting color in Table D.
	3. Compare the color to the indicator strip’s color key to determine the pH range (e.g., 5–6) and record.
7. **Test predictions about the pH of the other two basic solutions.**
	1. Using what you know from your models and the information provided in Table B, predict what the pH will be for solution B2 and solution B3. Record these predictions in Table D.
	2. Dip an indicator strip into beaker B2, and record the resulting color in Table D.
	3. Compare the color to the indicator strip’s color key to determine the pH range (e.g., 5–6) and record.
	4. Repeat b and c for Solution B3.
	5. Compare your data to your predictions in Table D and record any observations that might explain any differences.

**PART III: Utilizing a Multi-Use Indicator to Determine pH**

1. **Cook up a cabbage pH indicator.**
	1. Measure approximately 100 mL of water into a 250 mL beaker.
	2. Weigh out approximately 20 g of red cabbage.
	3. Cut the cabbage into small pieces about the size of a penny.
	4. Put the cabbage into the beaker of water.
	5. Heat the cabbage and water on a hot plate, stirring occasionally, until the temperature reaches approximately 95°C. Remove beaker from the hot plate using an oven mitt or hot pad. Allow to cool.
2. **Finish preparation of the cabbage pH indicator solution.**
	1. Fold a piece of filter paper in half. Fold it again in quarters. Open it in the center to create a filter paper “cone,” and place it in a funnel. Place the funnel and filter into a 250 mL Erlenmeyer flask.
	2. Once the cabbage solution is cool, pour it slowly through the filter system into the flask, which will remove the remaining solid cabbage particles. You have created a natural pH indicator solution.
3. **Calibrate the cabbage pH indicator by retesting the solutions.**
	1. Label 7 small paper cups for the solutions A1, A2, A3, distilled water, B1, B2, and B3.
	2. Transfer approximately 10 mL of each solution to the appropriate paper cup. If using a graduated cylinder, be sure to thoroughly rinse out any previous solutions, and rinse between transfers.
	3. Pour approximately 5 mL of cabbage indicator into each of the 7 paper cups. Gently swirl the cups to mix the solutions.
	4. Record the colors of the solution/indicator mix in Table D to complete the cabbage pH indicator key. Describe the colors in words or use colored pencils or markers, and note the pH range for each color equivalent to what you determined using indicator strips. Save the cups for comparison in Step 17.

**PART IV: Examining the pH of Common Household Solutions**

1. **Predict the pH of common household solutions.**
	1. Using the information provided in Table E, make predictions about the pH of the common household solutions provided by your instructor, based on what you know of their physical and chemical properties. Arrange them from the most basic to most acidic in Table F.
2. **Determine the pH of each solution using the cabbage indicator solution.**
	1. Label 6 small paper cups: “Ammonia,” “Glass Cleaner,” “Vinegar,” “Milk,” “Baking soda solution,” and “Soda Pop.”
	2. Pour approximately 10 mL of each solution into the appropriate paper cup. If you use a graduated cylinder, be sure to rinse thoroughly between solutions. Be careful to avoid inhaling ammonia fumes.
	3. Test each solution with an indicator strip to measure the pH before any dilution with cabbage indicator. Compare the results with the indicator key, and record the pH values in Table F.

[If a pH meter is also being used, this is also a good time to test each solution using that method (rinse the probe well between solutions).]

* 1. Pour approximately 5 mL of cabbage indicator solution into each of the 6 paper cups. If you use a graduated cylinder, be sure to thoroughly rinse out any previous solutions. Gently swirl the cups to mix the solutions.
	2. Record the colors of the solutions in Table F and compare with the cabbage pH indicator key (the cups you prepared in Step 15) to determine the relative pH of each solution. Describe the colors in words or use colored pencils or markers.
	3. Once you have recorded your observation using the cabbage indicator, discard the solution/indicator mixes in a sink or as directed by your instructor.
	4. Compare the data from the indicator strips and the cabbage indicator to your predictions. How did the pHs determined by the cabbage solution compare with those determined by the indicator strips? Based on your results, reorder the solutions if necessary at the bottom of Table F.
1. **Clean up!**
	1. Dispose of all materials according to your teacher’s directions.

# Data

Record your data in your lab notebook or in the space below.

**Table A. Observations from Models of Relative H+ and OH**– **Concentrations**

|  |  |  |  |
| --- | --- | --- | --- |
| **pH** | **Sketch** | **Ratio****[H+]:[OH–]** | **Comparison of Ratios** |
| 6 |  | *Approximate:**Calculated:* |  |
| 7 |  | *Approximate:**Calculated:* |  |
| 8 |  | *Approximate:**Calculated:* |  |

**Table B. The pH Scale and Its Associated Ionic Concentrations**

|  |  |  |
| --- | --- | --- |
| **pH** | **Concentration of H+ Ions (mol/L)** | **Concentration of OH– Ions (mol/L)** |
| **Decimal Form** | **Scientific Notation** | **Decimal Form** | **Scientific Notation** |
| 14 | 0.00000000000001 | 1.0E-14 | 1 | 1.0E+00 |
| 13 | 0.0000000000001 | 1.0E-13 | 0.1 | 1.0E-01 |
| 12 | 0.000000000001 | 1.0E-12 | 0.01 | 1.0E-02 |
| 11 | 0.00000000001 | 1.0E-11 | 0.001 | 1.0E-03 |
| 10 | 0.0000000001 | 1.0E-10 | 0.0001 | 1.0E-04 |
| 9 | 0.000000001 | 1.0E-09 | 0.00001 | 1.0E-05 |
| 8 | 0.00000001 | 1.0E-08 | 0.000001 | 1.0E-06 |
| 7 | 0.0000001 | 1.0E-07 | 0.0000001 | 1.0E-07 |
| 6 | 0.000001 | 1.0E-06 | 0.00000001 | 1.0E-08 |
| 5 | 0.00001 | 1.0E-05 | 0.000000001 | 1.0E-09 |
| 4 | 0.0001 | 1.0E-04 | 0.0000000001 | 1.0E-10 |
| 3 | 0.001 | 1.0E-03 | 0.00000000001 | 1.0E-11 |
| 2 | 0.01 | 1.0E-02 | 0.000000000001 | 1.0E-12 |
| 1 | 0.1 | 1.0E-01 | 0.0000000000001 | 1.0E-13 |
| 0 | 1 | 1.0E+00 | 0.00000000000001 | 1.0E-14 |

**Table C. Calculation of H+/OH– Ratios at More Extreme pH**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **pH** | **Concentration of H+ (mol/L)** | **Concentration of OH– (mol/L)** | **Ratio****(H+:OH–)** | **Explanation** |
| 2 |  |  |  |  |
| 11 |  |  |  |  |

**Table D. pH of Serial Dilutions**

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Solution Name** | **A1** | **A2** | **A3** | **Distilled****Water** | **B3** | **B2** | **B1** |
| **Predicted pH Range** |  |  |  |  |  |  |  |
| **Indicator Strip Color** |  |  |  |  |  |  |  |
| **Indicator Strip pH Range** |  |  |  |  |  |  |  |
| **Cabbage Indicator Color** |  |  |  |  |  |  |  |
| **Observations** |  |

**Table E. Common Properties of Acids and Bases**

|  |  |  |
| --- | --- | --- |
| **Common Properties of Acids** |  | **Common Properties of Bases** |
| *Sour taste**Feels sticky* *Can corrode metal and skin**Usually gases or liquids* | *Bitter taste**Feels slippery* *Reacts with oils and greases**Frequently solids* |

**Table F. pH of Common Solutions**

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  | **Strongest Acid** |  | **Weakest Acid** | **Neutral** | **Weakest Base** |  | **Strongest Base** |
| **Predicted Order** |  |  |  | *Water* |  |  |  |
| **Indicator Strip Color** |  |  |  |  |  |  |  |
| **Indicator Strip pH Range** |  |  |  |  |  |  |  |
| **Cabbage Indicator Color** |  |  |  |  |  |  |  |
| **Cabbage Indicator pH Range** |  |  |  | *7* |  |  |  |
| **Correct Order** | *(lowest pH)* |  |  |  |  |  | *(highest pH)* |
| **Observations** |  |