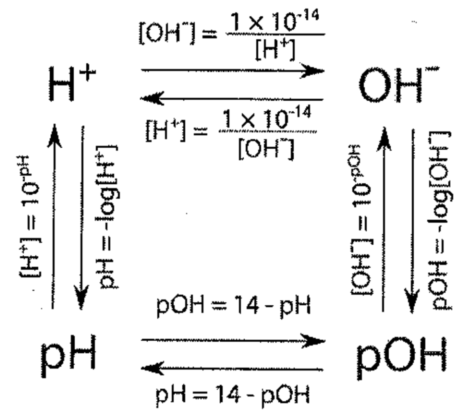
Name Key Date____________ Period__________

Acid Bases WebLab PhET Chromebook Version

https://phet.colorado.edu/en/simulation/acid-base-solutions

Click <**Introduction**> to begin.



Part 1: Procedure

1. The lab has 2 tools that allow you to test for pH values: A probe, and pH paper. Use each one by dipping it into the solution to be tested. Try all the given types of solutions and fill in the Data Chart with the pH value 0-14.

2. The circuit with a battery and bulb as shown: is the tool used to test for conduction of a solution. By dipping the wire leads into the solution, the bulb with either **remain unlit**, be **dimly lit**, be **somewhat bright** or **very bright**. Test each solution and record your observation for the bulbs brightness in the chart below.

Part 1: Data	pH Value from Probe	Color + pH Value from pH Paper	Observations from Circuit Tool Describe the brightness
Water (H_2O)	7.00	yellow 7	Dimly Lit
Strong Acid (*HA*)	2.00	red 2	Very Bright
Weak Acid (*A*)	4.50	orange 4-6	Somewhat Bright
Strong Base (*MOH*)	12.00	Blue 12-13	Very Bright
Weak Base (*B*)	9.50	Blue-Green 9-10	Somewhat Bright

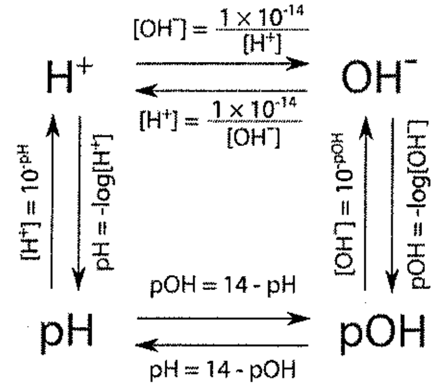
Part 1: Analysis

1. What pH value range is observed: a. for acids? Approx 2-6 b. for bases? Approx 9-13

2. *Why* are some solutions better conductors of electricity?

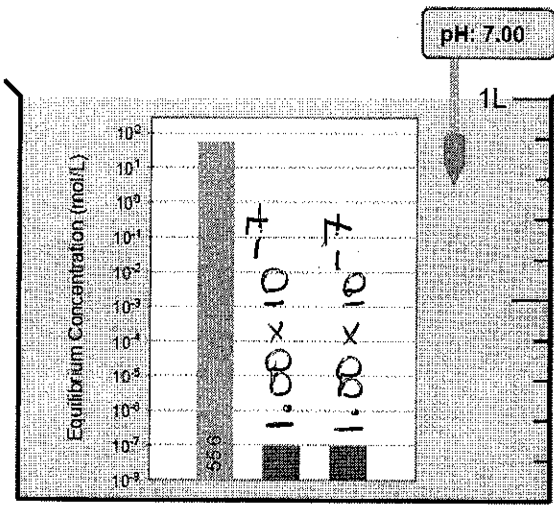
Because they ionize more completely in water

Part 2 Procedure, Data & Analysis:



Recall: The amount of ionization or dissociation of ions determines the strength of an acid or base. The concentration of $[H_3O^+]$, hydronium and $[OH^-]$, hydroxide ions can be used to calculate pH and pOH as shown on the diagram here:

1. Click on Water Solution, Graph View, Probe Tool. Insert the probe in the water. Notice that the initial concentration of the solution is given before any ionization or dissociation takes place. Fill in the missing concentration values for the hydronium and hydroxide ions on the chart here:



2. Use the concentration *value for $[H_3O^+]$* to calculate the pH. Show work:

$pH = -\log[H_3O^+] =$
$-\log(1.00 \times 10^{-7}) = 7$

3. Use the concentration *value for $[OH^-]$* to calculate the pOH. Show work:

$pOH = -\log[OH^-]$
$pOH = -\log(1.00 \times 10^{-7}) = 7$

$2\,H_2O \rightleftharpoons H_3O^+ + OH^-$
or H^+

4. Did your answer to #2 match the pH given in the simulation? yes, 7.

5. Is the answer to #3 equal to: (14 - pH)? yes show work: 14 - 7 = 7

6. Is the solution an acid, a base or neutral based upon the calculated pH? neutral

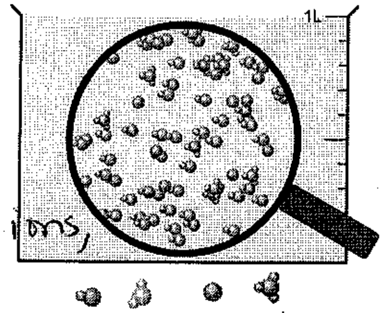
HONORS ONLY: ~~Repeat #2-6 for the four other solutions and attach notebook paper to show work.~~ Calculate the pH and pOH for the other solutions & show work on attached paper.

Part 3 Procedure, Analysis, Conclusion: My Solution

Across the bottom of the screen, click the  My Solution button. The default setting shows a weak acid with a concentration of 0.010 M. Insert the pH probe to show an initial pH of 4.50. The beaker is shown below:

1. Slide the initial concentration bar to the right to increase the number of solute molecules and then slide it to the left. What effect does changing the concentration have on the pH value? (Be specific)

Increasing the initial concentration means there is more solute (acid) in the beaker & therefore more $[H^+]$ ions, causing a lower pH.



$HA + H_2O \rightleftharpoons A^- + H_3O^+$

2. Return to your default setting and insert the probe. Now slide the strength to the right to make the acid stronger.

   a. As you increase the strength, describe the change in the number of blue A^- ions and orange H_3O^+ ions.

   Many extra A^- and H_3O^+ ions appeared and the HA began to disappear.

   b. As you increase the strength, describe the change in the concentrations of the ions in the solution? Hint: Click <Graph> to see how the concentrations rise and fall.

   Concentrations of A^- and H_3O^+ ions increased as strength of the acid was increased.

3. Yes or No? Does the pH seem to depend upon the concentration of $[H_3O^+]$ ions?

4. We always assume that strong acids will 100% ionize in water. Click reset and move the slider to strength: strong. Insert the probe. Record pH. Observe the number of ions in the beaker and click <Graph> to observe the concentrations.

   a. pH Value = 2.00

   b. YES or NO? Does the beaker contain particles that have not ionized and have 0% concentration? If so, what particle seems missing? HA. Why is it likely missing? Because it completely ionized (broke apart). There was none of the HA molecule remaining.

5. Click reset and change to a base. Repeat #1-4 above and describe any different results or simply write, "Same results for bases."

#1 Increasing concentration of a base caused a higher pH.

Conclusions: If the answer is no, explain why not. #2 - 4 Same Result for Bases

6. YES or NO? Can a weak acid be concentrated?

7. YES or NO? Can a strong acid be dilute?

8. YES or NO? For acids, can increasing the initial concentration increase the pH?

   It decreases the pH

9. YES or NO? For Bases, can increasing the initial concentration increase the pH?

Extension: In <my solution> Try different combinations of strength & initial concentration. Dip the probe and look in <graph> to record concentration of ion. Calculate pH to confirm results

Pt 2: Honors Only

Strong Acid $[H_3O^+] = 1.00 \times 10^{-2}$ m

$pH = -\log(1.00 \times 10^{-2}) = 2$ (pH)

$pOH = 14 - 2 = 12$ (pOH)

Weak Acid $[H_3O^+] = 3.16 \times 10^{-5}$ m

$pH = -\log(3.16 \times 10^{-5}) = 4.5$ (pH)

$pOH = 14 - 4.5 = 9.5$ (pOH)

Strong Base $[OH^-] = 1.00 \times 10^{-2}$ m

$pOH = -\log(1.00 \times 10^{-2} m) = 2$ (pOH)

$pH = 14 - 2 = 12$ (pH)

Weak Base $[OH^-] = 3.16 \times 10^{-5}$ m

$pOH = -\log(3.16 \times 10^{-5} m) = 4.5$ (pOH)

$pH = 14 - 4.5 = 9.5$ (pH)

note: The strong & weak have pH & pOH reversed.