NAME___________________________________ **SECTION**____________

PARTNER(S) _______________________________ **DATE**________________

EXPLORING ACIDS AND BASES

Part I: Properties of Acids and Bases

As you already know, it is not advisable to taste chemicals in an effort to analyze or identify them. But you have had prior personal experience with the taste of many distinctive chemicals. Circle your response to the following.

From you general experience, you categorize the taste of lemon juice (it contains citric acid) as:
sour bitter sweet

When you last got soap in your mouth (it contains traces of base), it tasted:
sour bitter sweet

1. Obtain a spot plate and place a few drops of the dilute solutions in each well in the order listed in the chart below. Test each one as indicated and enter the result or final color in the box provided. Litmus and pH indicators are in the form of small papers. Just place a glass rod into the sample and touch it to the paper to get the result. Use a small piece of universal pH paper as possible. For phenolphthalein, add a drop from the dropper bottle into the well. Add just a small piece of Mg (cleaned with steel wool) and see what happens. The phenolphthalein will not affect the Mg results.

Sample	Affect on blue litmus	Affect on red litmus	pH (from chart)	Result with phenolphthalein	Result with Mg
HCl					
NaOH					
H_2SO_4					
H_2O or HOH					
KOH					
$HC_2H_3O_2$ or CH_3COOH					
NH_3 (aq) (NH_4OH)					

2. Based on the results of your tests divide the compounds into three groups. For example, which of the solutions gave the same test results with litmus and phenolphthalein?

Group 1	
Group 2	
Group 3	

Look at the chemical formulas for the members of each group. Do members of a group have any component in common?

The compound $HC_2H_3O_2$ (CH_3COOH) in water is vinegar and tastes sour. Can you supply a general name for the compounds in the group with $HC_2H_3O_2$ (CH_3COOH)?

An **acid** is a compound that produces H^+ ions in water solution. **Bases** produce OH^- ions under the same conditions. Compounds that do not form a surplus of either of these ions in aqueous solution are classified as **neutral**. This way of classifying compounds was developed by the Swedish chemist Arrhenius.

$$HCl \rightarrow H^+ + Cl^- \text{ (in water)}$$

$$NaOH \rightarrow Na^+ + OH^- \text{ (in water)}$$

Indicators such as litmus and phenolphthalein (HPth) change color because of structural and light absorbing properties that change when they lose a hydrogen ion.

$$\text{HPth (colorless)} \rightarrow \text{Pth}^- \text{ (pink)} + H^+$$

$$\text{HPth (colorless)} + OH^- \rightarrow \text{Pth}^- \text{ (pink)} + H_2O$$

Acids can react with active metals such as magnesium in a single replacement reaction that produced hydrogen gas.

As a review, write the reaction for Mg with H_2SO_4 here:

____

From an examination of the formulas for acids and the formulas for bases, and the tests you performed, identify your groups by their chemical classifications in the table below.

Group 1	
Group 2	
Group 3	

3. Based on your observations, circle the most appropriate answer for each statement below:

Acids in litmus are:	blue	red
Bases in phenophthalein are:	colorless	pink
With an active metal bases are:	reactive	unreactive
A pH above 7 indicates:	base	neutral acid
A pH below 7 indicates	base	neutral acid

4. It is important to note that even a sample of distilled water contains hydrogen (H^+) and hydroxide (OH^-) ions because of the reaction below:

$$H_2O \rightleftarrows H^+ + OH^-$$

However, in pure water the concentration of H^+ is equal to the concentration of OH^-.

Circle the correct response for each of the following statements:

In pure water:	$H^+ > OH^-$	$H^+ = OH^-$	$H^+ < OH^-$
Compared to water in acid:	$H^+ > OH^-$	$H^+ = OH^-$	$H^+ < OH^-$
Compared to water in bases:	$H^+ > OH^-$	$H^+ = OH^-$	$H^+ < OH^-$

5. Test the household products listed in the table below and fill in the results and your conclusions.

Sample	**Affect on litmus**	**pH (from chart)**	**Result with phenolphthalein**	**Acid or base?**	**More H^+ or OH^- ?**
cola drink					
cleaner					
juice					
bicarbonate solution					
vinegar					
hand soap					

anti-freeze					

6. From the lab on solutions, what is the criterion for determining whether or not a solution is a conductor of electricity?

Should acids and bases conduct electricity? Explain.

You will now determine how strong a conductor of electricity your acids and bases are. You will be using the Electrical Conductivity Apparatus (ECA) which permits a bulb to light when a conductor (metal strip) or conducting solution is placed between the electrodes. Recall from your study of solutions that compounds can be strong electrolytes, weak electrolytes or non-electrolytes.

Test each of the original solutions (in the wide mouth jars) in the ECA. REMEMBER TO PULL OUT THE PLUG BEFORE TOUCHING THE ELECTRODES and TO CLEAN THEM WITH DISTILLED WATER BETWEEN TESTS. Determine if the acid or base is a strong or weak electrolyte.

Sample	**Conductivity Result**	**Strong or Weak Electrolyte?**
HCl		
NaOH		
H_2SO_4		
H_2O or HOH (distilled)		
KOH		
$HC_2H_3O_2$ or CH_3COOH		
NH_3 (aq) (NH_4OH)		

Now here is a statement to ponder. Acids are not ionic compounds; they are covalent compounds and have no charged particles. Yet, their aqueous solutions do conduct electricity.

What is happening when these materials are placed in water?

Something must happen when they are placed in water that generates the H^{+}. The process is called **ionization**. Highly polar covalent bonds in acids behave like ionic bonds in the presence of water.

Explain how the amount of ionization compares for HCl and CH_3COOH based on the conductivity results.

7. Place about 1 mL of 0.1M HCl in a small test tube and add 1 drop of phenolphthalein.

   What color is the solution?____

   Now add NaOH dropwise slowly just until a visible change occurs.

   Do you think the solution still contains more H^+ than OH^-? Why or why not?

   Place the contents of the test tube in an evaporating dish and boil off the water.
   In the space provided below, describe what is left in the dish and write the chemical equation for the double replacement reaction that took place.

   As you were adding each drop of base to the acid, what must have been happening to the number of H^+ in solution? ______________________

8. Is there a way of knowing exactly how many H^+ are in a solution? Actually, the **pH scale** is based on numbers of moles of H^+ in solution.

So far you have determined the pH of solutions by using a paper indicator. The electronic apparatus, the pH meter, electronically measures the pH. Investigate the change in pH with changes in the concentration of the strong acid HCl.

The instructor will show you how to standardize and use the pH meter.

(You may find the following formula handy for doing dilution calculations)

First sample				Diluted sample		
Concentration	x	Volume	=	Concentration	x	Volume
1.0 M	x	1.0 mL	=	*.10 M*	x	10 mL

Prepare your own solution by starting with 10 mL of the available 1.0 M acid. Take the pH reading. Then use a 1.0 mL pipet to withdraw 1.0 mL of the 1.0 M acid and place it in a 10 mL graduated cylinder. Add distilled water to the 10 mL mark and mix well. This will produce a 0.10 M solution since you have diluted it by a factor of 10.

Looking at this in a slightly different way, the first sample contained a certain number of H^+ from the ionizing acid molecules; the second sample contains 1/10th as many H^+. When you finish taking the pH for the 0.10M sample, repeat the dilution process to make a 0.01M sample. What % of the H^+ ions that were in the original sample is now in this third sample?

Repeat the process until you make the 0.0001M solution.

Concentration of the HCl sample	**pH**
1.0M or 1×10^0	
0.10M or 1×10^{-1}	
0.01M or 1×10^{-2}	
0.001M or 1×10^{-3}	
0.0001M or 1×10^{-4}	

Make a statement about the relationship between the concentration and pH from the data table. Using the graphing calculator, plot a graph of pH as a function of the hydrogen ion concentration. What type of relationship do you find between these variables?

Part II: Analysis by Titration

We can use the fact that acids react with bases in precise amounts to determine the amount or concentration of an acid with a base or vice versa. **Titration** is a process of reacting two materials using measured quantities of each in order to determine solution concentrations or sample amounts. If the reaction involves acids and bases it is called an acid-base titration. By determining very precisely the volumes of acid and base solutions needed for neutralization, the concentration (molarity) of an acid or base can be determined by comparing it to an acid or base whose concentration is known accurately.

For the reaction below where all coefficients are the same:

$$HCl + NaOH \rightarrow NaCl + H_2O$$

at the neutralization point: **moles NaOH = moles HCl**

This point of complete neutralization is also called the **equivalence point**.

How would you know when complete neutralization occurred?

The number of moles of solute is given by the formula: **mol = M x V**

Therefore:

$$(\mathbf{M \times V})_{base} = \mathbf{mol}_{base} = \mathbf{mol}_{acid} = (\mathbf{M \times V})_{acid} \quad \text{or} \quad (\mathbf{M \times V})_{base} = (\mathbf{M \times V})_{acid}$$

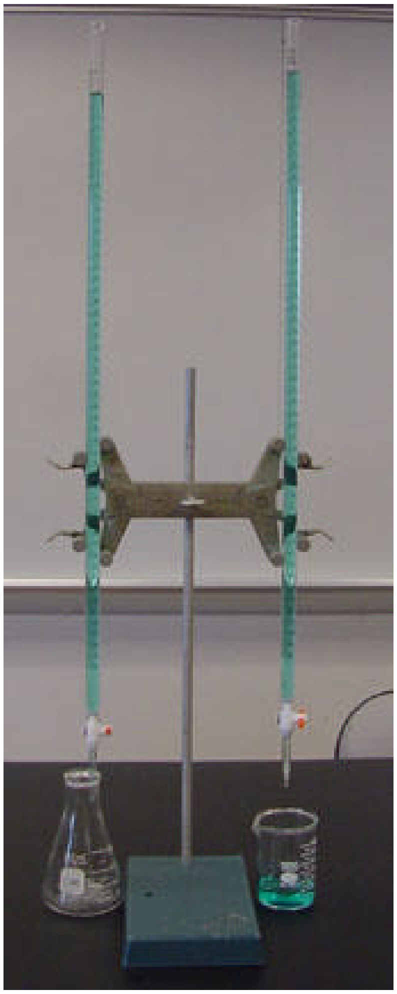
For example, if we wish to determine the concentration of a NaOH solution which is known to be about 0.5 M, we can titrate the solution against (using) an HCl solution of precisely known concentration. We use a **buret**, which is like a dispensing graduated cylinder calibrated in 0.1 mL (so you can read the volume to 0.01 mL). One solution is added to the other until the equivalence point is just reached. The **end point** is characterized by a change in the color of an added indicator such as phenolphthalein and should be very close to the neutralization or equivalence point if the indicator is chosen correctly. This is a visual detection method and is dependent on the analyst. If we know three of the variables in the equation above, we can determine the fourth mathematically. When we use this titration process with a solution of known or standard concentration to determine the concentration of a second the solution the process is known as standardization.

Note: An acid or base solution with a concentration recorded to three decimal places (such as 0.515 M) is generally considered a standard. Solutions with concentrations expressed to one decimal place

(0.5 M) are not standard.

Standardization

Ask your Instructor whether you will be doing this part individually or as a group. If you are instructed to work as a group, be sure all members of the group have an opportunity to do one titration.



1. Obtain a ring stand, buret clamp, and two 50 mL burets. Clean the burets if needed and rinse with distilled water. Label one buret for acid and the other for base.

2. In a clean beaker obtain approximately 100 mL of standardized HCl solution. Record the concentration on the data sheet.

3. Using a small amount (5-10 mL) of the acid, rinse the acid buret being sure to run some acid through the tip. Discard waste acid in the sink. Fill the acid buret with 45 to 50 mL being sure the liquid level is within the marked region and there are no air bubbles in the tip. Read the initial volume to the nearest 0.01 mL and record in the data table. Remember to read the lowest point of the curved liquid or **meniscus**. Repeat the procedure with the second buret and the NaOH.

4. Place a 250 mL Erlenmeyer flask beneath the buret containing the HCL and allow 25-30 mL of acid to flow into the flask. Record the final level to the nearest 0.01 mL. Add 3-4 drops of phenolphthalein to the flask and swirl the flask to mix.

5. Place the flask with the acid and indicator under the buret containing the NaOH. Partially open the stopcock and allow the NaOH to flow SLOWLY into the flask. As the NaOH flows in swirl or agitate the flask to mix the solutions thoroughly without splashing. As you get close to the equivalence point, a pink color will appear where the NaOH enters the flask solution and then disappears on mixing. At this point, add the NaOH drop-by-drop until the palest pink color remains for a period of 30 seconds. At this point the solution is just past the neutralization point.

6. To determine the end point more precisely (or to compensate for overshooting the end point) add a few drops more of the HCl solution until the solution just turns colorless with good mixing. Add NaOH dropwise until the palest pink color remains. Record the final volumes for the HCl and NaOH on the data sheet.
7. Repeat the titration two more times.

8. Complete the calculations to determine the molarity of the NaOH and find the average. Determine the range of concentrations (that is, highest minus lowest value). Record all results on the data sheet.

Titration Data/ Results Sheet

HCl Concentration ______________ **M**

	Trial 1	**Trial 2**	**Trial 3**
HCl			
Initial reading (I)			
Final reading (F)			
Volume used (F-I)			
NaOH			
Initial reading (I)			
Final reading (F)			
Volume used (F-I)			
$M_{NaOH} = (M_{HCl} \times V_{HCl})/V_{NaOH}$			
M_{NaOH}			
Average M_{NaOH}			
Range of M_{NaOH}			

Compare your average and range with the results from another student or group. Support your comparison with data.

Post-Lab Questions

1. You are handed a colorless liquid sample and asked to determine if it is an acid or base using at least two simple tests. Explain what you would do and expect to observe for each possibility.

2. Write the reaction for the titration performed in Part II.

   What should the pH at the neutralization point of this titration be? ________________

   An indicator changes color at the end point but this may not be the same as the neutralization or equivalence point. Indicators vary in the pH range where a color change occurs. What criterion would you use to select an indicator for a neutralization reaction?

   How can we ever know if we have reached the exact neutralization or equivalence point?

3. In the Part II titration you used a buret. You did the same titration with a dropper in Part I. Which titration was more precise, the one using the dropper or the one using the buret? Explain.

4. In the titration of HCl with NaOH we could use the formula $[(M \times V)_{base} = (M \times V)_{acid}]$ because the number of moles of acid and base at the neutralization point were equal. Could we use this equation for the titration of H_2SO_4 with NaOH? If not, what would you have to do to arrive at the M_{acid}?

Using your response and the data from your titration in Part II, calculate the molarity as if the acid was H_2SO_4 rather than HCl.

5. Now, how can aqueous ammonia, NH_3 (aq) be a base? The Arrhenius concept cannot handle ammonia, which behaves like a base. A better definition was formulated by Bronsted and Lowry. They defined acids as proton (H^+) donors and bases as proton (H^+) acceptors.

Based on the Bronsted-Lowry definition, write an equation that shows how aqueous ammonia is a base.

6. Aqueous solution of HCl, HBr, and HI are very conductive, while HF (aq) is not. Explain the difference in behavior.