Chemistry Hour\_\_\_\_\_ Name\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
Dr. Wexler
Comparing the Strengths of Acids (HS-PS1-4; HS-PS1-5)
Date \_\_\_\_\_

**Name of Lab Activity:**Comparing the Strengths of Acids

**Background:**
A. Chemical Equilibria
In some chemical reactions, the reactants are not entirely converted to products. This is because as the products form, they react to re-form the reactants in a reverse reaction.

When the rate of the forward reaction is equal to the rate of the reverse reaction, the system is said to be at equilibrium. At equilibrium, the concentrations of the reactants and products do not change, and these concentrations are determined by the equilibrium constant (K) of the reaction.

B. Ionization of Monoprotic Acids: Strong vs. Weak Acids
Acids ionize and produce hydrogen ions in aqueous solution. The strength of an acid depends on how completely it ionizes. Strong acids undergo essentially complete ionization, whereas weak acids are only partly ionized at equilibrium. Like other reversible reactions, the ionization of a weak acid can be represented by an equilibrium constant expression. The value of this expression, called the acid ionization constant, Ka, is a measure of the strength of a weak acid. Ka is the equilibrium constant for the reaction in which a weak acid is in equilibrium with its conjugate base in aqueous solution. For monoprotic acids like acetic acid: (the acid releases a single proton when it dissociates)
CH3COOH ⇌ CH3COO- + H+
Ka = [CH3COO-][H+] / [CH3COOH]

If you know both the molarity of the acid solution as well as its pH (pH = -log[H+], then you can easily calculate Ka since [CH3COO-] and [H+] are equal:
 Ka = [CH3COO-][H+] / [CH3COOH] = [H+][H+] / [CH3COOH] = [H+]2 / [CH3COOH]

The formula for Ka shows that strong acids have a higher Ka than weak acids.

C. What is pH?
We are surrounded by dilute solutions of acids and bases, inside and out. The orange juice and coffee that help us start our day are acidic, and the gastric juices they mix with in our stomach are also naturally acidic. Many of our household cleaning liquids, like the detergent solution in our washing machines and the bleach we add to whiten our clothes, are basic.
 

D. Dissociation of Water and pH
Pure water undergoes a reversible reaction in which both H+ and OH- are generated.
H2O(l) H+(aq) + OH-(aq)

The equilibrium constant for this reaction, called the water dissociation constant, Kw, is 1.0 x 10-14 at 25°C.
Kw = [H+][OH-] = 1.0 x 10-14 at 25°C

Because every H+ ion that forms is accompanied by the formation of an OH- ion, the concentrations of these ions in pure water are the same and can be calculated from Kw = 1.0 x 10-14 M2

Kw = [H+][OH-] = 1.0 x 10-14 M2

[H+] = [OH-] = 1.0 x 10-7 M

The equilibrium constant expression shows that the concentrations of H+ and OH- in water are linked. As one increases, the other must decrease to keep the product of the two concentrations equal to 1.0 x 10-14. If an acid, like hydrochloric acid, is added to water, the concentration of the H+ ion goes up, and the concentration of the OH- ion goes down (by forming H2O), but the product of those concentrations remains the same. An acidic solution is defined as a solution in which the [H+] > [OH-]. A basic solution has [H+] < [OH-].

E. **pH is defined** as the negative logarithm of the H+ concentration: **pH = -log[H+]**.
If a solution is 1.0 x 10-6 M H+, we say that the pH is -log(1.0 x 10-6) = 6.00.
**The higher the [H+], the lower the pH.** From **[H+] = 10-pH** we see that 10-6 = pH 6; 10-5 = pH 5; 10-4 = pH 4, etc.

We can do three types of calculations:
1. find pH given molarity and Ka
2. find Ka given molarity and pH
3. find molarity given Ka and pH

1. We can calculate the pH of an acid if we know both its molarity and its Ka

To derive [H+]: Ka = [H+]2 /[acid] 🡪 [H+]2 = Ka [acid] 🡪**[H+] = (Ka [acid])1/2**

A. Calculate [H+] if Ka = 1.75 x 10-5 and the acid concentration is 1.5M
 [H+] = (Ka [acid])1/2 = (1.75 x 10-5 x 1.5)1/2 = (2.625 x 10-5)1/2 = (26.25 x 10-6)1/2 = 5.12 x 10-3M
B. pH = -log[H+]
 pH = -log(5.12 x 10-3) = 2.29

2. We can calculate the Ka of an acid if we know both its molarity and its pH.
For example, if a 1.5M solution of acetic acid has a pH of 2.29, then:
A. **Ka = [H+]2 /[acid]** = [H+]2 /1.5
B. pH = 2.29, so [H+] = 10-pH = 10-2.29 = 5.129 x 10-3
C. Solving: Ka = (5.129 x 10-3)2 /1.5 = (2.631 x 10-5)/1.5 = 1.75 x 10-5

3. We can calculate molarity of an acid if we know both its Ka and pH
**[acid] = [H+]2/Ka**
[H+] = 10-pH
For example, if the pH =2.29 and the acid Ka = 1.75 x 10-5, then
 [acid] = [H+]2/Ka = (10-2.29)2/1.75 x 10-5 = 10-4.58/1.75 x 10-5 = 2.63 x 10-5/1.75 x 10-5 = 1.50M

**Special Materials:**1.0M acetic acid
0.1M acetic acid
0.1M hydrochloric acid
Digital pH meter
pH 4.5 calibration solution
Distilled water

**Pre-Lab:**1. Explain the difference between the terms “weak acid” and “dilute acid”.

2. Formic acid has a Ka of 1.80 x 10-4. Acetic acid has a Ka of 1.75 x 10-5. Which acid is stronger? Explain.

3. Define pH.

4. For a given acid, what happens to the pH and the hydrogen ion concentration as the acid becomes more concentrated?
A. pH

B. [H+]

5. For a given acid, what happens to the pH and the hydrogen ion concentration as the acid becomes more dilute?

A. pH

B. [H+]

4. What is the molar concentration of H+ ions in a solution having a pH of 4? What is the molar concentration of OH- ions? (for [OH-]: use the product rule for exponents, where [H+][OH-] must equal 10-14)

A. [H+]

B. [OH-]

5. What is the concentration of H+ ions in a solution having a pH of 10? What is the concentration of OH- ions?
A. [H+]

B. [OH-]

**6. What is the pH of a 1.0M solution of acetic acid? Show all calculations.**
Given: Ka acetic acid = 1.75 x 10-5
Ka = [H+]2 /[acid] 🡪 [H+]2 = Ka [acid] 🡪[H+] = (Ka [acid])1/2
pH = -log[H+].

A. Solve for [H+]:

B. Solve for pH:
Hint: calculate the log of [H+], then change the negative sign to a positive.

**7. What is the pH of a 0.1M solution of acetic acid? Show all calculations.**Given: Ka acetic acid = 1.75 x 10-5
Ka = [H+]2 /[acid] 🡪 [H+]2 = Ka [acid] 🡪[H+] = (Ka [acid])1/2
pH = -log[H+].

A. Solve for [H+]:

B. Solve for pH:
Hint: calculate the log of [H+], then change the negative sign to a positive.

C. Is this pH higher or lower than the pH for the 1.0M acetic acid solution?

**Procedure and Results:**
**Part 1: Effect of acid concentration on pH
A. Determine the pH of a 1.0M solution of acetic acid**
Instructions: You have already calculated the predicted pH for a 1.0M solution of acetic acid (prelab Q6). Use a digital pH meter to determine the pH of this solution. Take three measurements and calculate the average. Rinse off the pH meter electrode with dH2O and recalibrate before each reading.
Trial 1
Trial 2
Trial 3

Experimental pH = \_\_\_\_\_\_\_\_

Predicted pH = \_\_\_\_\_\_\_\_

**B. Determine the pH of a 0.1M solution of acetic acid**
Instructions: You have already calculated the predicted pH for a 0.1M solution of acetic acid (prelab Q7). Use a digital pH meter to determine the pH of this solution. Take three measurements and calculate the average. Rinse off the pH meter electrode with dH2O and recalibrate before each reading.
Trial 1
Trial 2
Trial 3

Experimental pH = \_\_\_\_\_\_\_\_

Predicted pH = \_\_\_\_\_\_\_\_

**C. Fill in the comparison table below.**

|  |  |
| --- | --- |
| Acetic acid concentration | pH |
| 1.0M  |  |
| 0.1M |  |

Are your results consistent with your predictions (prelab Q6 and Q7)?

**Part 2: Which is a stronger acid: acetic acid or hydrochloric acid? Compare their pH.**Instructions: compare the pH of a 0.1M solution of HCl and a 0.1M solution of CH3COOH

Measure the pH of a 0.1M solution of hydrochloric acid (HCl)

Trial 1
Trial 2
Trial 3

Experimental pH of 0.1M HCl = \_\_\_\_\_\_\_\_\_

Experimental pH of 0.1M CH3COOH (from Part 1) = \_\_\_\_\_\_\_\_\_\_\_

**Conclusion:** Which is a stronger acid? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Research questions:**
1. Acid rain is rainwater that has a pH lower than the normal value of 5.5 for rainwater. It has been blamed for the corrosion of buildings and statues and for widespread environmental damage, including damage to fish, forests, and crops. Government regulations aimed at reducing acid rain (and acid precipitation in general, both wet and dry) focus mainly on limiting the release from automobiles, power plants and factories of sulfur oxides (especially sulfur dioxide) and nitrogen oxides into the air. Explain why controlling emissions of these two chemicals reduces acid precipitation.

2. Explain why the global increase in CO2 levels is damaging to ocean life such as corals, shellfish, and fish populations.