

EXPERIMENT E1

Chemistry of the Kitchen - Acids and Bases

Prof. T. Hamade, UM-SJTU JI & SJTU Chemistry Department

(Modified version of D. Van Dinh and University of Michigan General Chemistry Laboratory Manual)

I. OBJECTIVES

- Classify common household chemicals as either acids or bases using a homemade indicator solution.
- Develop an understanding of pH scale.
- Examine the difference between strong and weak acids.
- Titrate a sample of vinegar to determine the concentration of acid.

II. INTRODUCTION

Take a moment and look around your house. Many every day chemicals used in your home are acids or bases. Some of most common acids and you may recognize and use includes vinegar (acetic acid), lemon juice (citric acid), and ammonia (some cleaners). Acids and are essential substances in home, industry and the environment. For example, the vast quantity of sulfuric acid manufactured in the United States each year is needed to produce fertilizers, polymers, steel, and many other materials. The influence of the acids on the environment can be seen through acid rain, which has caused numerous historic buildings and monuments to erode. If you have ever had a goldfish, you know how important it is to monitor and control the acidity of the water in aquarium (Figure 1). A characteristic that acids and bases share is their ability to turn certain organic compounds, such as vegetables materials, distinctive colors. Utilizing this knowledge, what are some methods in which you could test the acidity of the water in your goldfish aquarium?



Figure 1. A healthy fish tank requires careful control of acidity levels.

III. BACKGROUND

A. Properties of Acids and Bases

Since the 17th century, acids and bases have been characterized by their sour and bitter taste, respectively. In modern chemistry, these concepts have taken on considerably more precise meaning. In fact, there are three definitions of acids and bases, the **classical (Arrhenius)**, the **Brønsted – Lowry**, and the **Lewis**, which greatly expand our knowledge in these chemicals. Acids and bases differ greatly in their strength in water, that is, in the amount of H_3O^+ or OH^- produced per mole of substance dissolved. They are generally as either strong or weak, depending on their dissociation into ions in water. Acids and bases are **electrolytes** in water, so the classification of acid and base strength correlates with the classification of the electrolyte strength: strong electrolyte such as strong acids dissociate completely (Figure 2), and weak electrolytes such as the weak acids undergo partial dissociation (Figure 3).

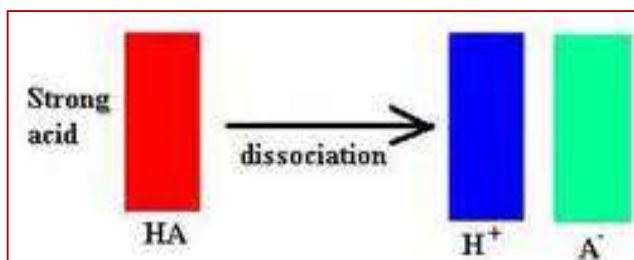
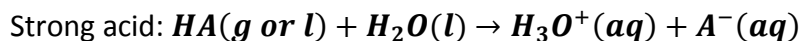
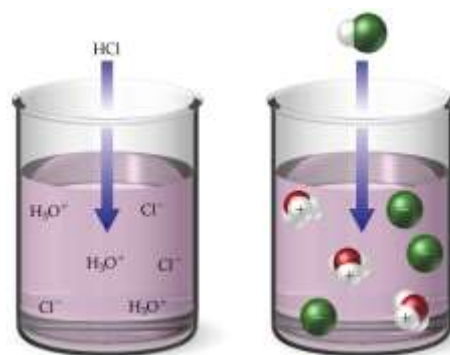
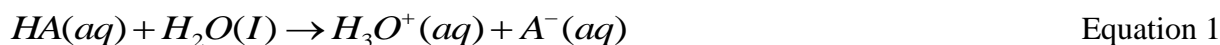


Figure 2. Behavior of a strong acid in water



In an aqueous solution of a strong acid, virtually no HA molecules are present (Equation 1).



At equilibrium, the $[\text{H}_3\text{O}^+] = [\text{A}^-] = [\text{HA}]_{\text{initial}}$ and the $[\text{HA}]_{\text{equilibrium}} = 0$. Then the equilibrium constant for strong acid is essentially very large ($K_a = [\text{H}_3\text{O}^+][\text{A}^-]/0$, very large). Since this process essentially goes to completion, a single arrow is used (\rightarrow) in the case of strong acids.

The situation is different in the case of a weak acid. From Figure 3, the majority of HA molecules in a solution of a weak acid are undissociated.

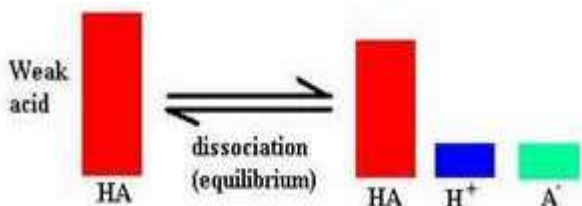
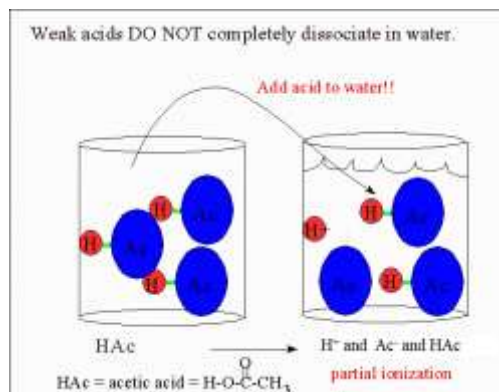
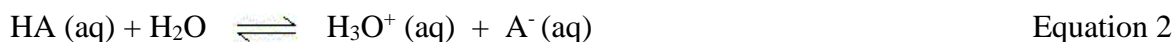


Figure 3. Behavior of a weak acid in water



Weak acid: $\text{HA} (\text{g or l}) + \text{H}_2\text{O} (\text{l}) \rightarrow \text{H}_3\text{O}^+ (\text{aq}) + \text{A}^- (\text{aq})$

Thus, $[\text{H}_3\text{O}^+] \ll [\text{HA}]_{\text{initial}}$ and $[\text{HA}]_{\text{equilibrium}} < [\text{HA}]_{\text{initial}}$ because the amount of acid dissociated is insignificant compared to initial concentration. Notice in Equation 2 below that the reaction is expressed as an equilibrium (\rightleftharpoons) indicating that the reaction does not necessary produce 100% products. In fact in the case of most weak acids, less than 5% of original HA molecules will actually dissociate to produce $\text{H}_3\text{O}^+(\text{aq})$ and $\text{A}^-(\text{aq})$ ions.



The percentage of acid that is actually dissociated can be quantified in terms of the acid dissociation constant, K_a , whose expression for the dissociation of a general weak acid, HA, in water is shown in Equation 3. Notice $[\text{H}_2\text{O}]$ does not show up in the form of K_a .

At equilibrium:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \quad \text{Equation 3}$$

$K_a = x \cdot x / ([\text{HA}]_{\text{init}} - x) = \text{small}$. Since x is insignificant compared to $[\text{HA}]_{\text{init}}$, then $K_a = x^2 / [\text{HA}]_{\text{init}}$.

Like any equilibrium constant, the magnitude of K_a tells how far to the right the reaction has preceded when equilibrium is reached. Thus, the stronger the acid, the higher the $[\text{H}_3\text{O}^+]$ at equilibrium, the larger the K_a :

Stronger acid \Rightarrow higher % HA dissociation \Rightarrow higher $[\text{H}_3\text{O}^+]$ \Rightarrow larger K_a

Likewise, the weaker the acid, the smaller the K_a :

Weaker acid \Rightarrow lower % HA dissociation \Rightarrow lower $[\text{H}_3\text{O}^+]$ \Rightarrow smaller K_a

B. The pH Scale

In aqueous solutions, $[\text{H}_3\text{O}^+]$ can vary over an enormous range: from about 10 M to 10^{-15} M. to handle numbers with negative exponents more conveniently in calculations, we convert them to positive numbers using numerical system called p-scale, the negative of the common (base 10) logarithm of the number. Applying this numerical system to $[\text{H}_3\text{O}^+]$ gives pH, the negative logarithm of $[\text{H}_3\text{O}^+]$ or $[\text{H}^+]$ (Equation 4).

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{Equation 4}$$

For example, a solution with a $[H_3O^+]$ of $5.4 \times 10^{-4} M$ has a pH of 3.27 (Equation 4a).

$$pH = -\log[H_3O^+] = -\log(5.4 \times 10^{-4}) = 3.27 \quad \text{Equation 4a}$$

Likewise, the pH of a $1.0 \times 10^{-12} M$ H_3O^+ solution is 12.00 on the pH scale. Note that the higher the pH, the lower the $[H_3O^+]$. Therefore, an acidic solution has a lower pH (higher $[H_3O^+]$) than a basic solution (see for example in Figure 4 how abundance ratio of CH_3COOH drops with pH while the CH_3COO^- increases with pH). At $25^\circ C$ the $[H_3O^+]$ in pure water is $1.0 \times 10^{-7} M$, so the pH of pure water at $25^\circ C$ is 7.00 and aqueous solutions typically fall within range 0 to 14. This information is summarized in Table 1.

Type of solution	pH
Basic	> 7.00
Neutral	$= 7.00$
Acidic	< 7.00

Table 1. pH of acids & bases

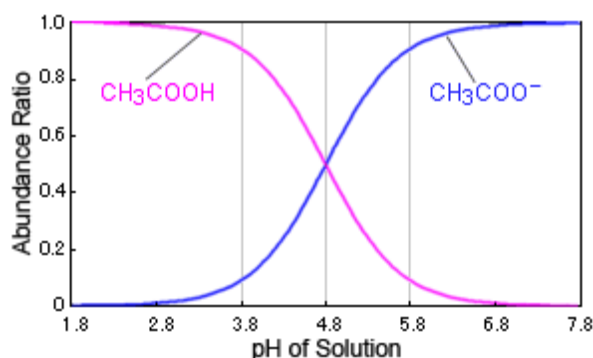


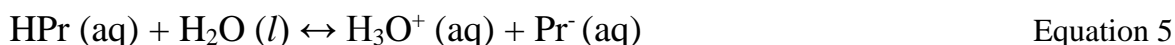
Figure 4. Acetic acid abundance ratio with pH

In the laboratory, pH values are usually obtained with an acid-base indicator or with a more precise instrument called a pH meter.

C. Calculating the pH

It is possible using what we have learned so far to calculate the theoretical pH of any aqueous acidic solution without using a pH meter. In the case of a strong acid, our job is simple since we can assume that the concentration of $[H_3O^+]$ is the same as the starting concentration of the strong acid and we simply use the calculation from Equation 4. However for weak acids the amount of acid partially dissociated to $[H_3O^+]$ is much smaller than the amount of acid itself and can be calculated just from knowing the K_a and the initial concentrations of the weak acid. Giving the K_a and $[HA]_{initial}$ of the weak acid to calculate $[H_3O^+]$, we can calculate the pH. As an example, we will calculate the pH of 0.10 M solution of propanoic acid (C_2H_5COOH , which we symbolize as HPr). A good first step in doing these problems is to look up the value of the K_a . In the case of propanoic acid, $K_a = 1.3 \times 10^{-5}$ at $25^\circ C$.

In order to determine the $[H_3O^+]$ of this solution, we must next write the balanced equation and K_a expression, as in Equation 2 (Equation 5) and 3 (Equation 5a), respectively.



$$K_a = 1.3 \times 10^{-5} = \frac{[H_3O^+][Pr^-]}{[HPr]} \quad \text{Equation 5a}$$

From Equation 5, we see for each mole of HPr that dissociate one mole of H_3O^+ and Pr^- each produced. Knowing this and given $[HPr]_{initial}$ (0.10 M) we can set up a reaction table (Table 2), commonly called an I.C.E. table (Initial Change Equilibrium table).

	HPr (aq)	+ H ₂ O (l)	↔	H ₃ O ⁺ (aq)	+ Pr ⁻ (aq)
Initial	0.10 M	-		0	0
Change	- x	-		+ x	+ x
Equilibrium	0.10M- x	-		x	x

Table 2. Weak acid concentration calculations at equilibrium

In the “Initial” row we include the starting concentration of the weak acid. Note that the concentration of H_3O^+ and the conjugate base Pr^- are both initially equal to zero. Because we don’t yet know how much of the HPr dissociates, we express the changes in terms of variable, x. we know that how much HPr is lost (-x) in the “Change” row, we must gain the same amount: (+x) of H_3O^+ and Pr^- . The “Equilibrium” row is the sum of the “initial” and “Change” rows.

We are now ready to substitute information from our “Equilibrium” row into the K_a expression and solve for x (Equation 5b).

$$K_a = \frac{[H_3O^+][Pr^-]}{[HPr]} = 1.3 \times 10^{-5} = \frac{(x)(x)}{(0.10 - x)} \quad \text{Equation 5b}$$

Since K_a is small for HPr, it dissociates very little so the value of x is negligibly small compared to 0.10 M; therefore, we avoid solving the quadratic equation by approximating $0.10 \text{ M} - x = 0.10 \text{ M}$ (Equation 5c)

$$1.3 \times 10^{-5} = \frac{(x)(x)}{(0.10)} \quad \text{Equation 5c}$$

$$x = \sqrt{(0.10)(1.3 \times 10^{-5})} = 1.1 \times 10^{-3} \text{ M} = [H_3O^+] \quad \text{Equation 5d}$$

Equation 6 shows that now that we have found the $[H_3O^+]$ in our solution of 0.10 M propanoic acid, we can calculate its pH using Equation 4.

$$pH = -\log[H_3O^+] = -\log(1.1 \times 10^{-3}) = 2.96 \quad \text{Equation 6}$$

Handling log significant figures (SF): $\log 1.1$ (2SF all) = 2.96 (2SF beyond decimal point). $\text{Anti } \log 2.96$ (2SF beyond decimal point) = 1.1 (2SF all).

This mathematical process can also be reversed. For example, an environmental chemist can measure the experimental pH and use it to determine the concentration of acid in a sample of acid rain.

D. Titration

In addition to the method described above, the concentration of an acid can also be determined experimentally by an acid – base titration. In an acid – base titration, a measured volume of acidic solution of unknown is placed in a flask beneath a burette containing the known (standardized) base solution (Figure 5). A few drops of acid – base indicator, usually phenolphthalein, are added to the flask containing the acidic solution. The standardized solution of base is then added slowly to the flask until the end point is reached. The end point of the titration occurs when the indicator changes color permanently due to the presence of excess OH^- ions (phenolphthalein is colorless in acid and pink in base). Knowing the stoichiometry of the acid – base reaction and the amount of base used to reach the end point, scientists can discover the unknown $[H_3O^+]$ of the acidic solution.



Figure 5. Experimental setup for an acid-base titration

E. Overview

In this series of experiments, you will examine several of various methods in which the pH of a solution can be measured. Knowing that acids and bases possess the ability to change the color of certain indicator solutions, you will use red cabbage extract and a universal indicator in Part A as acid – base indicators. You will develop a pH scale based on the fact that red cabbage extract changes colors when mixed with different household chemicals. In Part B, you will study the differences between strong and weak acids using pH meter. Using this knowledge, you will then determine an unknown concentration of a weak acid. In Part C, you will calculate the concentration of acid in vinegar by performing an acid – base titration.

IV. EXPERIMENTAL PROCEDURES: “Make sure you take photos of your favorite lab work for use in your final PPT presentation assigned by your TA about one of the experiments E1-E5”

Part A. Relative Acidity/ Basicity of Common Household Products

Chemicals used	Materials used
Universal indicator papers: Strips of universal pH indicator paper. Various household chemicals: (example ammonia, vinegar, Shampoo (colorless), soda (a colorless variety), milk, lemon juice, liquid detergent (without dyes), milk of magnesia, tap H ₂ O, deionized H ₂ O, Bleach)	Several beakers of various sizes 10-mL Graduated cylinder Glass stirring rod Test tube w/rubber stopper Test tube rack Funnel Hot plate & Knife

1. Working alone, select one or two household products to test from the available samples. Rinse a freshly washed beaker with 5 mL of your selected household chemical(s).
2. Pour 10 mL of the same household product into the beaker(s). Immerse partially a strip of universal pH indicator paper, then let the indicator paper(s) dry and label it with the sample that you tested. Repeat step 1 once again.
3. Label your samples and bring the indicator paper to the front of the lab for observation by the class. Be sure to record results for other household products (color of indicator paper and corresponding pH from the indicator paper box or from the chart posted on laboratory benches).
4. Repeat procedures (2.&3.) to test the pH of a stock solution of 1M NH₃.H₂O (inside eye-dropper vial) but use directly only 1-2 drops.

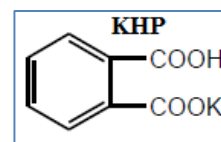
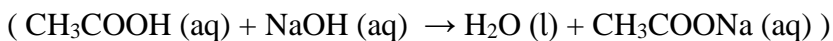
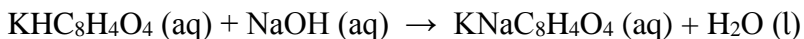
For the remaining procedures parts B & C, students work in groups as assigned by instructor where each two students in a group works together.

Part B. Concentration of unknown molarity of NaOH solution using KHP titration

Introduction

Potassium hydrogen phthalate (KHP, KHC₈H₄O₄) can be used to titrate the concentration of NaOH solution. (Acetic Acid (HAc): CH₃COOH)

$$C_{\text{NaOH}} V_{\text{NaOH}} = m_{\text{KHP}} / MW_{\text{KHP}}$$



Chemicals used	Materials used
Unknown molarity NaOH solution inside bottle labelled as 0.1M NaOH. KHP pH indicator (phenolphthalein)	Provided glassware: beakers, Erlenmeyer flasks, & graduation cylinders Weighing papers & desiccators

Procedure for Part B

1. At a designated area, wear protective gloves and wash the provided glassware of your group using detergent, brush and tap water and then rinse with de-ionized water, 3 times each. Be careful not to break or crack the glassware and not to injure yourself. Any broken or cracked glassware must be reported to the instructor and the glassware must be disposed carefully in the provided glass disposal box and not into the trash can or sink.
2. Using the Mettler balance, carefully weigh out 0.4000-0.6000g KHP into 3 labeled freshly washed Erlenmeyer flasks, record each weight. Add no more than 30-40 mL of de-ionized water to entirely solve KHP and 2-3 drops of phenolphthalein (to one of the flasks).
3. Run about 25.0 mL of deionized water through a 50-mL burette and then rinse the burette with about 5.0 mL of the labelled 0.1 M NaOH solution. Mount the burette in the burette stand.
4. Use a funnel to add roughly 30.0 mL of the labelled 0.10 M NaOH solution into the burette. Drain any air bubbles from the tip of the burette into a waste beaker. Record the initial volume V_0 of the burette (± 0.02 mL).
5. Slowly add the NaOH solution from the burette drop-wise to the KHP solution in the flask with the indicator, swirling the flask after each addition. Continue until the endpoint is reached. Record the final volume V_1 .
6. Repeat steps 4-5 for a second and third trial using the other flasks with KHP solution.

Part C. Acid – Base Titration of Vinegar Solution

$$\text{Calculation: } 10 C_{\text{NaOH}} V_{\text{NaOH}} = C_{\text{HAc}} V_{\text{HAc}}$$

Chemicals Used	Materials Used
Same solution bottle of NaOH of Part B. Vinegar with unknown concentration of acetic acid (CH_3COOH) Phenolphthalein	Burette and burette stand 250-mL Erlenmeyer flask 100-mL Beaker (2) 25-mL Pipet w/bulb & graduated cylinder

1. Run about 25.0 mL of deionized water through a 50-mL burette and then rinse the burette with about 5.0 mL of the labelled 0.1 M NaOH solution. Mount the burette in the burette stand. (no need to rinse if you are using the same burette and stock solution bottle of NaOH from Part B)
2. Use a funnel to add the labelled 0.1 M NaOH solution (from Part B) into the burette and fill it up to near below the top mark. Drain any air bubbles from the tip of the burette into a waste beaker. Record the initial volume V_1 (± 0.02 mL).
3. Pipet 25.00 mL of the vinegar into a 250.00 mL volumetric flask, add de-ionized water to the calibration line and shake upside down several times.
4. Pipet 25.00 mL of the solution in step 3, into 3 freshly washed & rinsed Erlenmeyer flasks and add 2 – 3 drops of phenolphthalein to the flask.
5. Slowly add the NaOH solution, drop wise from the burette to the vinegar, swirling the flask after each addition. Continue until the endpoint is reached. Record the final volume V_2 (± 0.02 mL).
6. Repeat steps 2 – 4 for a second trial and third trial.

Part D. Closing Your Lab Session

1. Wear protective gloves and wash the used glassware using detergent, brush and tap water and then rinse with de-ionized water. Be careful not to break or crack the glassware. Any broken or cracked glassware must only be disposed carefully into the provided glass disposal box
2. Clean up your designated working area thoroughly so the next lab session can use safely.
3. A designated team from the groups will be instructed just before dismissal of the laboratory to inspect and insure the general cleaning and safety of the laboratory.

From now and on, you must follow Part D thoroughly when conducting any experimental work inside the lab area.

E1: Chemistry of the Kitchen: Acids and Bases

Name:	Lab instructor:
Date:	Lab Section:

V. PRE-LABORATORY EXERCISES (PLE)

1. Define the underlined words in the **BACKGROUND** section.
2. Determine the concentration of H_3O^+ in a solution with a pH of 9.78.
3. Given solutions of the same concentration, which would you expect to have a lower pH, a strong or weak acid? Explain.
4. In this laboratory experiment you will be calculating the concentration of an unknown sample of acetic acid using the pH of the sample and the K_a for acetic acid. Use your textbook to find the K_a for acetic acid. Be sure to record this value in your notebook so you will have it available during the experiment.
5. You will be performing a titration of vinegar (an aqueous solution of acetic acid, $HC_2H_3O_2$, also more commonly known as CH_3COOH) in this laboratory experiment. To prepare you for this titration, please read the section on acid – base titrations in your textbook and then do the following:

- a) Write the balanced molecular equation and the net ionic equation for the neutralization reaction between aqueous acetic acid and aqueous sodium hydroxide.
- b) You place 10.00 mL of a $\text{HC}_2\text{H}_3\text{O}_2$ solution of unknown concentration in a flask and add a few drops of indicator. You then titrate the acid with 0.24 M NaOH. If the initial reading on the burette was 0.19 mL and the final reading was 26.50 mL, what is the concentration of the $\text{HC}_2\text{H}_3\text{O}_2$ solution?
- c) Some solutions (such as vinegar) are commonly reported in terms of percent by mass. Assuming the density of the acetic acid solution you found in question 6b is the same as the density of pure water at 25°C , determine the percent by mass of the vinegar in the sample.
6. Using your notes from CHM110 and literature, write the equation that predicts the effect of temperature on equilibrium constant, then estimate to the equilibrium constant of acetic acid at 12°C by using its known value at 25°C .

E1: Chemistry of the kitchen: Acids and Bases

Name:	Lab instructor:
Date:	Lab section:

VI. RESULTS AND POST – LABORATORY QUESTIONS (PLQ)

Part A. Relative Acidity/Basicity of Common Household Products

Attach a table summarizing the colors of the universal indicator paper that you used to test the pH of the selected one or two household products. Put a star next to the household indicator product(s) that you personally tested. Your table should include a column that correlates the color of the universal indicator paper with the pH of the solution (this information is found on the box of the universal indicator paper or on a chart placed at the work benches).

Part B. Concentration of Unknown Molarity of NaOH Solution

Trial	1	2	3
m_{KHP} (g)			
Total Vol. of NaOH (mL)			
C_{NaOH} (mol/L) = $m_{\text{KHP}} \times 1000 / (MW_{\text{KHP}} \times V_{\text{NaOH}})$			
Average C_{NaOH} (mol/L)			

*Show your work for the calculation of the concentration of the unknown NaOH and compare to the value on the label of the used NaOH bottle.

Part C. Acid – Base Titration of Vinegar: $MW_{\text{KHP}} = 204 \text{ g/mol}$

	Initial Vol. of NaOH (mL)	Final Vol. of NaOH (mL)	Total Vol. of NaOH (mL)
Trial 1			
Trial 2			
Trial 3			
Average			

Show your work for the calculation of the molarity of acetic acid vinegar.

Show your work for the calculation of the % by mass of acetic acid in vinegar. How does this value compare to the value on the bottle of vinegar?

SAMPLE DATASHEET FOR A LAB SECTION
(FOR REFERENCE ONLY, STUDENTS TO OMIT THIS PAGE)

	A	B	C	D	E	F	G	H	I	J	K	L	M	N	O	P
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