Chem in the Kitchen: Natural Acid-Base Indicators , Buffers and Salts

Organic acids are found in many fermented foods. A carboxylic acid is the functional group found in an organic acid. For example, vinegar is a solution of acetic acid. Acetic acid is an organic **monoprotic acid**.

Other organic acids found in the kitchen include citric acid and malic acid.

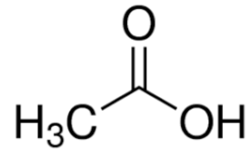

Figure 1 : Acetic acid structure

Acid-Base Indicators

An acid-base indicator is a molecule whose color responds to changes in the hydronium ion concentration (**recall that pH = -log [H3O$^+$]).** The equilibrium for an indicator is shown below with HIn, the weak acid form of the indicator, and In$^-$, the conjugate base.

$$\mathrm{HIn(aq)} \quad + \quad \mathrm{H2O(l)} \quad \Leftrightarrow \quad \mathrm{In^-\ (aq)} \quad + \quad \mathrm{H3O^+(aq)} \qquad (1)$$

(**weak acid form**) (**conjugate base)**

One indicator used for analyzing acids and bases is phenolpthalein. The weak acid form of phenolphthalein is colorless and its conjugate base is red. Phenolphthalein changes color between pH 8 and 10. Phenolphthalein is colorless in an acidic solution and magenta in a basic solution (Figure 2).

(a) **(b)**

Figure 2 phenolphthalein as an acid-base indicator

a) Vinegar sample, with indicator, is a pale yellow b) the indicator is magenta in a basic solution

Natural Acid-Base Indicators

Molecules from plants and flowers can also be used to determine if a solution is acidic, basic or neutral. Anthocyanins are one class of compounds that are natural acid-base indicators. Anthocyanins are responsible for the dark red, purple and blue pigments in fruits and vegetables. In contrast to phenolphthalein, red cabbage anthocyanins exhibit a broad range of colors over the pH scale. This indicator can be used to estimate the pH of a solution.

In the first part of this lab, you will explore the color changes of the anthocyanins present in butterfly pea flower over a broad range of pH standards (pH 1-12). You will determine the approximate pH ranges where each indicator changes colors.

Buffers

A **buffer** is a solution whose pH does not change significantly when small amounts of a strong base or acid are added. The action of a buffer is due to the presence of both a weak acid and a weak base, usually present in roughly equal molar amounts. Consider a sodium acetate-acetic acid buffer system as an example. In this buffer, there are approximately equal amounts of a weak acid (acetic acid) and a weak base (acetate from sodium acetate)). If a strong acid were added to the buffer, the acetate ion would react with it

$$H^+(aq) + H_3C_2OO^-(aq) \rightarrow H_3CCOOH(aq) \quad (2)$$

Likewise, added OH^- would be consumed by the weak acid

$$OH^-(aq) + H_3CCOOH(aq) \rightarrow H_2O(l) + H_3C_2OO^-(aq) \quad (3)$$

Since both H^+ and OH^- can be removed by reaction with the components of the buffer, the pH of the solution will remain fairly stable. Only when enough H^+ or OH^- has been added to almost exhaust the supply of $HCOO^-$ or HCOOH will the pH change significantly (will the **buffer capacity** be exhausted).

In the second part of this lab you will prepare a buffer using vinegar (a solution of acetic acid) and sodium acetate. You will prepare a sodium acetate-acetic acid buffer and measure the pH. Finally, you will analyze the effect addition of a strong acid or base has on the pH of a buffer.

Recognizing Salts That Hydrolyze

The acidity or basicity of a salt depends on the parent acid and base used to produce the salt. **Neutral salts** are formed from the neutralization of strong acids with strong bases. Thus KCl, the product of the reaction between the strong acid HCl and the strong base KOH, is a neutral salt. Because it contains only K^+ and Cl^-, neither capable of hydrolysis,[1] its presence has no effect on pH. **Basic salts** are formed from the neutralization of a weak acid and a strong base, *e.g.,* KOOCH can be thought of as the product of the weak acid HCOOH and the strong base KOH. The cation in this basic salt does not affect pH, but the $HCOO^-$ anion, which is the conjugate base of a weak acid, will hydrolyze to produce OH^- (equation (5c)). The weaker the parent acid, the stronger the basicity of the anion in the salt, and the more this salt raises pH.[2] **Acidic salts** are formed from neutralization of a strong acid and a weak base, *e.g.,* CH_3NH_3Cl is the product of the reaction between the strong acid HCl and the weak base CH_3NH_2. Such salts lower pH because they contain the conjugate acid of a weak base (in CH_3NH_3Cl, the acid is $CH_3NH_3^+$).

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[1] Why don't neutral salts hydrolyze and affect pH? Consider the possible products if the ions from KCl *did* hydrolyze. The products of the hydrolysis of K^+ would be KOH (a strong base) and H^+. Immediately upon forming they would react with one another to re-form K^+ and H_2O, and no net reaction would occur. Similarly for chloride ion, the products would be HCl (a strong acid) and OH^-. These, too, would react with one another to re-form Cl^- and H_2O and no net reaction would occur. So we conclude that neutral salts like KCl, NaCl, *etc.* do not hydrolyze:

$$K^+ + H_2O \rightarrow \text{no reaction} \quad (a)$$

$$Cl^- + H_2O \rightarrow \text{no reaction} \quad (b)$$

[2] Within each acid-base *conjugate pair*, the relationship is: The conjugate base of an acid increases in strength as the acid becomes weaker. Thus the conjugate base of a strong acid has virtually no basic nature while the conjugate base of a very weak acid is very strong. Similarly, the conjugate acid of a base increases in strength as the base strength decreases. So, the conjugate acid of a strong base has virtually no acidic nature, but the conjugate acid of a very weak base is very strong.

PRELAB	• Write your prelab (***Purpose***) in your lab notebook calling the entry "Prelab for Chemistry in the Kitchen Part I."
LAB NOTEBOOK PREP	. None all data will be recorded on the report sheets

Part A Natural Acid-Base Indicators

Heat 25.0 mL of water in a beaker to 80.0 °C. Weigh ~0.5 gram of butterfly pea flowers and add to a 250 mL beaker. Carefully add the warm water and stir the flowers in the water for five minutes. Filter the flowers by gravity filtration. The filtrate is the pea flower indicator solution. Allow to cool. Choose 5 pH standards over the range of 1-12. You want to make sure you cover the whole range. Add about 1 mL of each standard to a test tube. Next, add about 10 drops of the butterfly pea flower to each pH standards. Record the color changes in your notebook. Make sure to record your observations for color vs pH for each indicator.

Part B Exploring a buffer

Calibrating the pH Electrode - Offset Error Measurement

Connect the pH meter to a LabQuest device. Calibrate the Vernier combination pH electrode (determine its offset error) using pH 4, 7, and 10 buffers as described in Appendix H. Record the details of the offset error measurements and calculations in your notebook. *Time saving tip: Calibration can easily be done by one person. Other group members should move on to other tasks.*

Measure 10.0 mL of 0.100 M sodium acetate. Record the pH value in your notebook. Correct the pH using the offset error Measure 10.0 mL of 0.80 M acetic acid (aka vinegar). Record the pH value in your notebook. Correct the pH using the offset error. Mix thoroughly 10.00 mL of vinegar with 10.00 mL of sodium acetate solution. Record the pH of the buffer. To 5.00 mL of distilled water add 5-10 drops of pea flower indicator. Add 1.00 M HCl dropwise until a color change observed. Record the number of drops of HCl. Repeat, this time using 5.00 mL of the buffer (use two separate test tubes) . Record the number of drops of HCl.

Part C pH of salt solutions

Aqueous 1.0<u>M</u> solutions of $NaC_2H_3O_2$, $NaHSO_4$, Na_2CO_3, NH_4Cl are available in small beakers. Measure the pH of each solution. Correct the pH values using the offset error.

Report sheets (turn in one/group)

Part A: (4 pts)

pH

	Color

Part B: (2 pts)

# of drops for color change with water

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# of drops for color change with buffer

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Part C: (4 pts)

Salt Solution	**pH**

Part C continued (8 pts)

Write a net ionic equation for each salt solution that did not have a pH of 7.00 in Part C

Questions (3pts/question)

1. Is acetic acid a weak or strong acid? Explain your answer in 1-2 sentences.

2. Write a net ionic equation for the reaction of acetic acid with sodium hydroxide.

3. Phenolpthalein is a synthetic indicator used in some acid-base titrations that changes color between pH 8-9. Given your observations with the color changes for butterfly pea flower, would this natural indicator be a good choice for a titration of a weak acid with a strong base with a pH of ~8.5 at the equivalence point.

4. Write two net ionic equations to show how the sodium acetate/acetic acid buffer neutralizes the addition of a strong acid (such as HCl) or strong base (such as NaOH).