Chem 112 – Experiment 5 – Simulation – pH Indicators

Background

Using pH Indicators

The pH of a substance (how acidic or basic it is) can be measured using a pH meter. It can also be measured, less accurately, using pH (acid-base) indicators — substances that change color depending on pH. Outside the lab, strips coated with pH indicators are used to measure the pH of swimming pools and fish tanks.

pH

Pure water contains equal concentrations of the hydroxide and hydronium ions and therefore is neutral. Water dissociates slightly into ions in an equilibrium process called autoionization as shown below.

$$2H_2O(l) \Leftrightarrow H_3O^+ + OH^-$$

The equilibrium constant for water is represented as **K_w**.

$$K_w = [OH^-][H_3O^+] = 1\times10^{-14} \text{ @ } 25^\circ C$$

A change in $[H_3O^+]$ causes an inverse change in $[OH^-]$. Therefore, all acidic solutions will contain a low $[OH^-]$, whereas all basic solutions will contain a low $[H_3O^+]$. The acidity or basicity of a solution is often expressed as pH calculated as the negative logarithm of the $[H_3O^+]$.

$$pH = -\log_{10}[H_3O^+]$$

This means that the higher the $[H_3O^+]$, the lower the pH. Therefore, an acidic solution will have a pH lower than 7, a neutral solution will have a pH of 7 (at 25^{o}c), and a basic solution will have a pH higher than 7. The general pH range is from 0 to 14. For example, a solution that is 1M in H_3O^+ has a pH of 0, whereas a solution that is 0.001 M in OH^- has a pH of 11. The pH can be measured in the laboratory by using either acid-base indicators or a pH meter.

Another useful relationship that can be derived from Kw is the relationship between pH and pOH. Since you will already, by this stage, have seen this in class I will not go into the derivation of it but simply give you the resultant relationship.

$$pH + pOH = 14 \text{ @ } 25^oC$$

Acid–Base Equilibrium

Strong Acids and Bases

Acids and bases are classified according to the degree to which they dissociate in water. Strong acids and bases dissociate completely, whereas weak acids and bases only partially dissociate into ions. For example, HCl, which is a strong acid, dissociates completely into hydronium ion (H_3O^+) and its conjugate base, Cl^- as shown in the reaction below.

$$HCl(aq) + H_2O(l) = H_3O+ + Cl^-$$

Similarly, a strong base such as NaOH completely dissociates into hydroxide ion (OH^-) and its conjugate acid (Na^+).

$$NaOH(aq) \longrightarrow OH^- + Na^+$$

Because strong acids completely dissociate into hydronium ions, the initial concentration of the acid can be used to determine the pH. Similarly, for strong bases the initial base concentration is used to determine the pOH, the pH subsequently determined by using the relationship between pH and pOH.

Weak Acids and Bases

However, a weak acid (HA) only partially dissociates into hydronium ion and its conjugate base (A^-), and equilibrium is established.

$$\mathbf{HA(aq) + H_2O(l) \Leftrightarrow H_3O^+ + A^-}$$

The equilibrium constant is called the acid dissociation constant (Ka) and is calculated using the formula below.

$$\mathbf{Ka = [H_3O^+][A^-] / [HA]}$$

where $[H_3O^+]$, $[A^-]$, and [HA] are the molar concentrations of the hydronium ion, conjugate base, and undissociated acid, respectively at equilibrium. Ka is a measure of the acid's strength. The smaller the Ka, the weaker the acid is.

This is better done via an example. Acetic acid (CH_3CO_2H) is a weak acid whose Ka = 1.8x10-5 @25^{o}C. How do you determine the pH of an 0.5M acetic acid solution?

a) Obvious first step is to write a balanced chemical equation for the equilibrium in question!

$$\mathbf{CH_3CO_2H(aq) + H_2O(l) \Leftrightarrow CH_3CO_2^- + H_3O^+}$$

b) Write the equilibrium expression.

$$K_a = \frac{[CH_3CO_2^-][H_3O^+]}{[CH_3CO_2H]}$$

(Remember: - pure liquids and solids do not appear in the expression, thus no $[H_2O(l)]$)

c) Now the fun part – setting up the ICE Table.

	$[CH_3CO_2H(aq)]$	+	$[H_2O(l)]$	⇔	$[CH_3CO_2^-]$	+	$[H_3O^+]$
I	0.5				0		0
C	-x				x		x
E	0.5-x				x		x

$$1.8x10^{-5} = \frac{(x)(x)}{0.5-x}$$

i) *Not your lucky day and you have to rearrange this so that we get the beloved quadratic equation.*

$$x^2 = (0.5-x)(1.8x10^{-5})$$
$$x^2 = 9.0x10^{-6} – (1.8x10^{-5})x$$
$$x^2 + (1.8x10^{-5})x – 9.0x10^{-6}$$

Now solve for x ☹. Just looking at the ICE Table, x, among other things is equal to the $[H_3O^+]$ and thus your way to determine the pH

ii) *It is your lucky day and you can use the approximation method. So much easier.*

Note: - *Since you have already met this in class and know the criteria needed in order to use the approximation method, I will forego repeating it here. All I will say is that whichever method you use,* ***approximation*** *or solving the* ***quadratic****, the answer will be the same. It is your choice as to which method you prefer.*

A similar approach can be used for weak bases.

pH (Acid–Base) Indicators

The pH of a substance can be measured using either pH (acid–base) indicators or a pH meter. Acid–base indicators are substances that change color as a function of pH. They can be used either in solution or as pH paper, a paper strip soaked with one or more indicators. Each acid–base indicator has a specific color change pH interval that is dependent on the indicator's pKa ($-\log_{10}K_a$). Within the color change interval, indicators are often a mixture of colors between the low and high pH colors. Generally, the color change starts and ends approximately one pH unit on either side of its PK_a. For example, the pK_a of bromocresol green is 4.7 and the color change takes place over a pH range of ~3.8 to 5.4. Given the color change pH interval, some indicators are better for measuring the pH of acidic solutions, whereas others are more suitable for measuring the pH of basic solutions.

About This Lab

In this lab, you will investigate different pH indicators with solutions of varying pH. Solutions of HCl and NaOH will be diluted with water to obtain a range of pH values. You will then have to identify the most effective pH indicator for each pH range.

Open the simulation by clicking on the virtual lab icon shown on the left on the Hayden-McNeil Web Site. The simulation will launch in a new window.

You may need to move or resize the window in order to view both the Procedure and the simulation at the same time.

Follow the instructions in the Procedure to complete each part of the simulation. When instructed to record your observations, record data, or complete calculations, record them for your own records in order to use them later to complete the post-lab assignment.

Procedure

Experiment 5a – Measure the pH of Acids

1. Take **seven small**, clean **test tubes** from the **Containers shelf** and place them on the workbench.

2. **Double-click** on the test tubes and **label them with numbers 1–7**.

3. Take **0.1 M hydrochloric acid** from the **Materials shelf** and **add 5 mL** to **test tube 1**.

4. Take **water** from the **Materials shelf** and **add 5 mL** to **test tube 1**.

5. Take **water** from the **Materials shelf** and **add 9 mL** to **test tubes 2–7**.

6. Create a **series of successively diluted acidic** solutions **as follows**:

a)	Pour 1 mL from 1 into 2.
b)	Pour 1 mL from 2 into 3.
c)	Pour 1 mL from 3 into 4.
d)	Pour 1 mL from 4 into 5.
e)	Pour 1 mL from 5 into 6.
f)	Pour 1 mL from 6 into 7.

7. Take **bromothymol blue** from the **Materials shelf** and **add 2 drops to each test tube**. **Observe** the **color** of the solutions. **Record your results** to reference later.

8. Clear your station by dragging all of the test tubes to the recycling bin beneath the workbench.

9. **Repeat steps 1–6**.

10. Take **methyl yellow** from the **Materials shelf** and **add 2 drops** to **each test tube**. **Observe** the **color** of the solutions. **Record your results**.

11. Clear your station by emptying the test tubes into the waste and then placing the test tubes in the sink.

12. **Repeat steps 1–6.**

13. Take **bromocresol green** from the **Materials shelf** and **add 2 drops** to **each test tube**. **Observe** the **color** of the solutions. **Record your results**.

14. Clear your station by emptying the test tubes into the waste and then placing the test tubes in the sink.

Experiment 5b – Measure the pH of Bases

1. Take **seven small**, clean **test tubes** from the **Containers shelf** and place them on the workbench.
2. **Double-click** on the **test tubes** and **label them with numbers 1–7**.
3. Take **0.100 M sodium hydroxide** from the **Materials shelf** and **add 5 mL** to **test tube 1**.
4. Take **water** from the **Materials shelf** and **add 5 mL** to **test tube 1**.
5. Take **water** from the **Materials shelf** and **add 9 mL** to **test tubes 2–7**.
6. Create a **series of successively diluted basic solutions as follows:**

a)	Pour 1 mL from 1 into 2.
b)	Pour 1 mL from 2 into 3.
c)	Pour 1 mL from 3 into 4.
d)	Pour 1 mL from 4 into 5.
e)	Pour 1 mL from 5 into 6.
f)	Pour 1 mL from 6 into 7.

7. Take **bromothymol blue** from the **Materials shelf** and **add 2 drops** to **each test tube**. **Observe** the **color** of the solutions. **Record your results**.
8. Clear your station by emptying the test tubes into the waste and then placing the test tubes in the sink.
9. **Repeat steps 1–6.**
10. Take **alizarin yellow** from the **Materials shelf** and **add 2 drops** to **each test tube**. **Observe** the **color** of the solutions. **Record your results**.
11. Clear your station by emptying the test tubes into the waste and then placing the test tubes in the sink.
12. **Repeat steps 1–6.**
13. Take **phenolphthalein** from the **Materials shelf** and **add 2 drops** to **each test tube**. **Observe** the **color** of the solutions. **Record your results.**
14. Clear the bench of all materials, containers, and instruments.