**Magnet Biochem pH and Molarity [concentration]**

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. Although the topic of ‘acid and base equilibrium’ is not covered until AP Chem, learning a little in this course will help with your understanding of solutions in terms of molarity, so that you will be able to understand the origin of the pH scale for describing acids and what the pH value says about an acidic or basic solution.

According to the Arrhenius theory of acids and bases, when an acid is added to water, it donates an H+ ion to water to form H3O+ (often represented by H+). The higher the concentration of H3O+ (or H+) in a solution, the more acidic the solution is. An Arrhenius base is a substance that generates hydroxide ions (OH-) in water. The higher the concentration of OH- in a solution, the more basic the solution is.

Pure water undergoes a reversible reaction in which both H+ and OH- are generated.

H2O(l)    http://preparatorychemistry.com/images/arrow_double_CS.gif    H+(aq)  +  OH-(aq) \*note: aq=aqueous (in solution)

The equilibrium constant for this reaction, called the water dissociation constant, Kw, is 1.01 × 10-14 at 25 °C. (\*In AP Chem, you learn how to derive this, but we will not in this course. I’m just using it to make a point)

Kw = [H+][OH-]  =  1.01 × 10-14   at 25 °C

Because every H+ (H3O+) 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.

Kw  =  [H+][OH-]  =  (x)(x)  =  1.01 × 10-14 [\*note: (x)(x)= x2, so take the square root of both sides]

x = [H+] =  [OH-]  =  1.01 × 10-7 M      (\*Note: M=Molarity=mol/L)   
                           (1.005 × 10-7 M before rounding)

What should you take from this? Note that both exponents are -7. Now think about what you know regarding the pH of water, and hopefully you will see where I’m going with this.

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 concentrations equal to 1.01 × 10-14 (at 25 °C). If an acid, like hydrochloric acid (a strong acid), is added to water, the concentration of the H+ goes up, and the concentration of the OH- goes down, but the product of those concentrations remains the same. An acidic solution can be defined as a solution where the [H+] > [OH-].

Typical solutions of dilute acid or base have concentrations of H+ and OH- between 10-14 M and 1 M. The table below shows the relationship between the H+ and OH- concentrations in this range.

**Concentrations of H+ and OH- in Dilute Acid and Base Solutions at 25 °C**

|  |  |
| --- | --- |
| **[H+]** | **[OH-**] |
| 1.0 M (1.0 x 100) | 1.0 × 10-14 M |
| 1.0 × 10-3 M | 1.0 × 10-11 M |
| 1.0 × 10-7 M | 1.0 × 10-7 M |
| 1.0 × 10-10 M | 1.0 × 10-4 M |
| 1.0 × 10-14 M | 1.0 M (1.0 x 100) |

 \*Notice that if we added the exponents across each column, they all = 14 (disrespective of the – sign)

We could describe the relative strengths of dilute solutions of acids and bases by listing the molarity of H+ for acidic solutions and the molarity of OH- for basic solutions. There are two reasons why we use the pH scale instead. The first reason is that instead of describing acidic solutions with [H+] and basic solutions with [OH-], chemists prefer to have one scale for describing both acidic and basic solutions. Because the product of the H+ and OH- concentrations in such solutions is always 1.01 × 10-14 at 25 °C, when we give the concentration of H+, we are indirectly also giving the concentration of OH-. For example, when we say that the concentration of H+ in an acidic solution at 25 °C is 10-3 M, we are indirectly saying that the concentration of OH- in this same solution is 10-11 M. When we say that the concentration of H+ in a basic solution at 25 °C is 10-10 M, we are indirectly saying that the OH- concentration is 10-4 M. The pH concept makes use of this relationship to describe both dilute acid and dilute base solutions on a single scale.

The next reason for using the pH scale instead of H+ and OH- concentrations is that in dilute solutions, the concentration of H+ is small, leading to the inconvenience of measurements with many decimal places, such as 0.000001 M H+, or to the potential confusion associated with scientific notation, as with 1 × 10-6 M H+. In order to avoid such inconvenience and possible confusion, pH is defined as the negative logarithm of the H+ concentration. A logarithm is an exponent, and that’s all you need to know about it for doing this.

pH  =  -log[H+]

Instead of saying that a solution is 0.0000010 M H+ (or 1.0 × 10-6 M H+) and 0.000000010 M OH- (or 1.0 × 10-8 M OH-), we can indirectly convey the same information by saying that the pH is 6.00.

pH =  -log[H+]  =  -log(1.0 × 10-6)  = 6.00

When taking the logarithm of a number, report the same number of decimal positions in the answer as you had significant figures in the original value. Because 1.0 × 10-6 has two significant figures, we report 6.00 as the pH for a solution with 1.0 × 10-6 M H+. The table below shows a range of pH values for dilute solutions of acid and base.

**pH of Dilute Solutions of Acids and Bases at 25 °C**

|  |  |  |
| --- | --- | --- |
| **[H+]** | **[OH-**] | **pH** |
| 1.0 (100) | 1.0 × 10-14 | 0.00 |
| 1.0 × 10-1 | 1.0 × 10-13 | 1.00 |
| 1.0 × 10-2 | 1.0 × 10-12 | 2.00 |
| 1.0 × 10-3 | 1.0 × 10-11 | 3.00 |
| 1.0 × 10-4 | 1.0 × 10-10 | 4.00 |
| 1.0 × 10-5 | 1.0 × 10-9 | 5.00 |
| 1.0 × 10-6 | 1.0 × 10-8 | 6.00 |
| 1.0 × 10-7 | 1.0 × 10-7 | 7.00 |
| 1.0 × 10-8 | 1.0 × 10-6 | 8.00 |
| 1.0 × 10-9 | 1.0 × 10-5 | 9.00 |
| 1.0 × 10-10 | 1.0 × 10-4 | 10.00 |
| 1.0 × 10-11 | 1.0 × 10-3 | 11.00 |
| 1.0 × 10-12 | 1.0 × 10-2 | 12.00 |
| 1.0 × 10-13 | 1.0 × 10-1 | 13.00 |
| 1.0 × 10-14 | 1.0 | 14.00 |

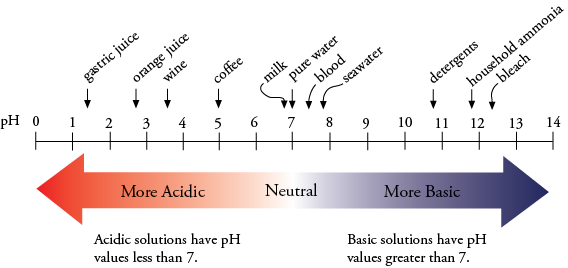
This table illustrates several important points about pH. Notice that

* When the solution is acidic ([H+] > [OH-), the pH is less than 7.
* When the solution is basic ([OH-] > [H+]), the pH is greater than 7.
* When the solution is neutral ([H+] = [OH-]), the pH is 7. (Solutions with pH's between 6 and 8 are often considered essentially neutral.)

Also notice that

* As a solution gets more acidic (as [H+] increases), the pH decreases.
* As a solution gets more basic (higher [OH-]), the pH increases.
* As the pH of a solution decreases by one pH unit, the concentration of H+ increases by ten times.
* As the pH of a solution increases by one pH unit, the concentration of OH- increases by ten times.

The pH, [H+], and [OH-] of some common solutions are listed in the figure below. Notice that gastric juice in our stomach has a pH of about 1.4, and orange juice has a pH of about 2.8. Thus, gastric juice is **more than ten times** more concentrated in H+ than orange juice. **A change of 1 on the pH scale is actually 10 fold!**



**pH of Common Substances**   Acidic solutions have pH values less than 7, and basic solutions have pH values greater than 7. The more acidic the solution is, the lower its pH. The more basic a solution is, the higher the pH. The corresponding H+ and OH- concentrations are shown in units of molarity. Notice that a decrease of one pH unit corresponds to a ten-fold increase in [H+], and an increase of one pH unit for a basic solution corresponds to a ten-fold increase in [OH-].

pH + pOH= 14

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

**Expect a test question or two from the above information**.

I will **not** give you problems (like the ones below) to solve on the upcoming test, but for those who plan to take AP Chem, I included a few problems below for reference:

**EXAMPLE 1**– pH Calculations:  The H+ concentration of a 0.025 M HCl solution is 0.025 M H+. What is its pH?

**Solution**:

pH = -log[H+]  =  -log(0.025) =  1.60

**EXAMPLE 2** – pH Calculations:  In Example 2, we found that the H+ concentration of a 2.9 × 10-3 NaOH solution was 5.1× 10-12 M H+. What is its pH?

**Solution:**

pH = -log[H+]  =  -log(7.5 × 10-12) =  11.29

We can convert from pH to [H+] and [OH-] using the following equations, as demonstrated in Examples 5 and 6.

[H+]  =  10-pH

http://preparatorychemistry.com/images/pH_web_3.png

**EXAMPLE 3** – pH Calculations:  What is the [H+] in a glass of lemon juice with a pH of 2.12?

**Solution:**

[H+]  =  10-pH  =  10-2.12  =  7.6 × 10-3 M H+

**EXAMPLE 4**  – pH Calculations: What is the [OH-] in a container of household ammonia at 25 °C with a pH of 11.900?

**Solution:**

[H+]  =  10-pH  =  10-11.900  =  1.26 × 10-12 M H+

http://preparatorychemistry.com/images/pH_web_4.png

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