Lesson 21: Acids & Bases - Far From Basic

Lesson Objectives:

- Students will identify acids and bases by the Lewis, Bronsted-Lowry, and Arrhenius models.
- Students will calculate the pH of given solutions.

Getting Curious

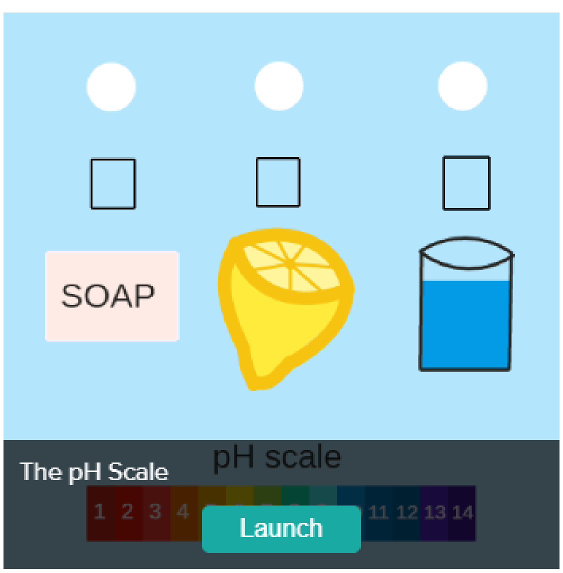
You've probably heard of pH before. Many personal hygiene products make claims about pH that are sometimes based on true science, but frequently are not. pH is the measurement of $[H^+]$ ion concentration in any solution. Generally, it can tell us about the acidity or alkaline (basicness) of a solution.

Click on the **CK-12 PLIX Interactive** below for an introduction to acids, bases, and pH, and then answer the questions below.

Directions: ***Log into CK-12 as follows:***

Username - your Whitmore School Username

Password - whitmore2018



Questions: Copy and paste questions 1-3 in the Submit Box at the bottom of this page, and answer the questions **before** going any further in the lesson:

1. After dragging the small white circles (in this simulation, the indicator papers) onto each substance, what do you observe about the pH of each substance?

2. If you were to taste each substance (highly inadvisable in the chemistry laboratory!), how would you imagine they would taste?

3. Of the three substances, which would you characterize as acid? Which as basic? Which as neutral? Use the observed pH levels to support your hypotheses.

Chemistry Time

Properties of Acids and Bases

Acids and bases are versatile and useful materials in many chemical reactions. Some properties that are common to aqueous solutions of acids and bases are listed in the table below.

Acids	Bases
conduct electricity in solution	conduct electricity in solution
turn blue litmus paper red	turn red litmus paper blue
have a sour taste	have a slippery feeling
react with bases to create a neutral solution	react with acids to create a neutral solution
react with active metals to produce hydrogen gas	

Note: Litmus paper is a type of treated paper that changes color based on the acidity of the solution it comes in contact with.

Defining Acids and Bases

An early way of classifying acids and bases was proposed by Svante Arrhenius, a Swedish chemist, in 1894. An Arrhenius acid is any compound that releases H^+ ions when dissolved in water. An Arrhenius base is a compound that generates hydroxide ions (OH^-) when dissolved in water. Some representative examples are given in the table below:

Acids	Bases
hydrochloric acid: $HCl \rightarrow H^+ +$ Cl-	sodium hydroxide: $NaOH \rightarrow Na^+ + OH^-$
nitric acid: HNO3 $\rightarrow H^+ + NO_3^-$	potassium hydroxide: $KOH \rightarrow K^+ + OH^-$
hydrobromic acid: $HBr \rightarrow H^+ + Br^-$	calcium hydroxide: $Ca(OH)_2 \rightarrow Ca^{2+} + 2\ OH^-$

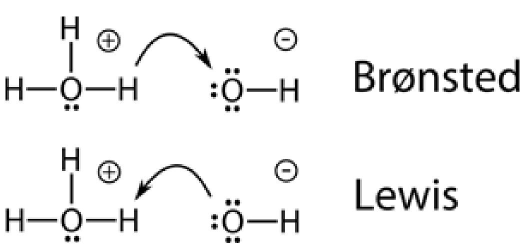
Many strong acids and bases can be identified based on the Arrhenius model. However, there are many compounds that share a number of common characteristics with acids and bases but do not fit the Arrhenius definitions. In the early 1920s, the Danish scientist Johannes Brønsted and the English researcher Thomas Lowry each published ideas that expanded the Arrhenius concept. According to this newer definition, a Brønsted-Lowry acid is any compound that can donate a proton (an H+ion) to an appropriate acceptor. A Brønsted-Lowry base is a compound that can remove (or accept) a proton from a relatively Brønsted-Lowry acid.

Overall, the Brønsted-Lowry model suggested that any acid-base reaction could be reduced to the transfer of a proton from an acid to a base. For example, the reaction between hydrochloric acid (HCl) and sodium bicarbonate ($NaHCO_3$) involves the transfer of an H^+ ion from the acid (HCl) to the base (the bicarbonate ion). The resulting carbonic acid (H_2CO_3) is unstable and breaks down to form carbon dioxide and water:

$$HCl + NaHCO_3 \rightarrow NaCl + H_2CO_3$$

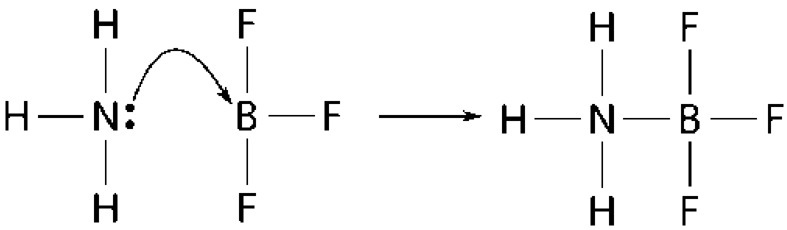
$$H_2CO_3 \xrightarrow{\text{breaks down to}} CO_2 + H_2O$$

Another more general definition of acids and bases was offered by the American chemist G.N. Lewis. A Lewis acid is any chemical species that accepts a pair of electrons, and a Lewis base is a chemical species that donates a pair of electrons. This is the broadest most commonly used definition, and all compounds that qualify as an acid or base under the previous definitions are also Lewis acids and bases.



In the diagram above, we see the same process illustrated multiple times, highlighting the ways in which the compounds are acting as an acid or a base according to each definition. According to the Brønsted-Lowry model, the protonated water molecule (the acid) is donating a proton to the OH^- ion (the base). According to the Lewis model, the hydroxide ion has a pair of electrons (indicated by the black bar) that it donates to the protonated water molecule. "Pronated" refers to the extra hydrogen atom in the molecule, which increases its charge from neutral to +1. In both instances, the hydroxide ion serves as a base and the protonated water molecule is the acid.

An example of a Lewis acid-base reaction that would not fit the other definitions of acid and base is the formation of an adduct between boron trifluoride and ammonia:



CK-12 Foundation

The ammonia serves as a Lewis base by donating its lone pair of electrons to make a new bond with boron (the Lewis acid), which has an empty orbital (indicated by ∘) that can accept two electrons.

Unless otherwise indicated, we will be using the Brønsted-Lowry model of acids and bases for the remainder of the chapter.

Monoprotic and Polyprotic Acids

Acids can further be categorized based on how many acidic hydrogen atoms they contain. Acidic hydrogen atoms are those which will be transferred to a base.
A monoprotic acid has only one acidic hydrogen that would be transferred to a strong base, whereas a polyprotic acid has two or more. Common monoprotic acids include HCl, HBr, and HNO_3. A common diprotic acid is sulfuric acid (H_2SO_4), and phosphoric acid (H_3PO_4) provides an example of a triprotic acid. In each case, all hydrogens are available to participate in acid-base reactions. However, that is not the case for all acidic molecules. For example, in acetic acid (CH_3COOH), only the hydrogen bonded to the oxygen atom is acidic. The other three hydrogens are covalently bonded to carbon and cannot be removed by any of the bases that we will consider in this chapter.

Name	Structure
hydrobromic acid	HBr
hydrochloric acid	HCl
hydrofluoric acid	HF
hydroiodic acid	HI
nitric acid	HNO_3
perchloric acid	$HClO_4$
acetic acid	CH_3COOH

Name	Structure
carbonic acid	H_2CO_3
sulfuric acid	H_2SO_4
sulfurous acid	H_2SO_3

Name	Structure
phosphoric acid	H_3PO_4

Questions: Copy and paste question 4 in the Submit Box at the bottom of this page, and answer the question before going any further in the lesson:

4.Of the three ways in which scientists define acids and bases, which makes the most sense to you? Explain that definition of acids and bases in your own words.

Self-Ionization of Water

Water is an example of what is known as an amphoteric substance, which means that it can act as both an acid and a base. In the presence of a strong acid, water can be a proton acceptor (a base), producing the hydronium ion (H_3O^+):

$Cl(aq)+H_2O(l)\rightarrow Cl-(aq)+H_3O^+(aq)$

However, water can also donate a proton (acting as an acid) when combined with a strong enough base, producing the hydroxide ion (OH^-):

$CH_3O^-(aq)+H_2O(l)\rightleftarrows CH_3OH(aq)+OH^-(aq)$

Overall, water is a weak acid and a weak base. Because it has both of these properties, any sample of liquid water undergoes the following acid-based reaction in which both hydronium and hydroxide ions are produced to a very small extent:

$H_2O(l)+H2O(l)\rightleftarrows H_3O^+(aq)+OH^-(aq)$

Because the reactants are in the liquid phase, they are not included in the equilibrium constant expression. As a result, the value of Keq for this reaction can be calculated as follows:

$Keq=[H_3O^+][OH^-]$

Water's ability to act as an acid or a base is relatively weak, so we would expect the reactants to be heavily favored in this equilibrium. Indeed, at 25°C, this equilibrium constant has a value of only 1.0×10^{-14}. However, despite the minimal extent of self-ionization, this is a fundamentally important equilibrium for any reactions that take place in water, which includes essentially all biochemical reactions that occur inside any living organism. Because of its particular importance, this equilibrium constant is given the special symbol K_w.

In a pure sample of water, there are no external sources of H_3O^+ or OH^- (no additional acids or bases), so for each H_3O^+ ion that is formed by the self-ionization of water, an OH^- ion will be formed as well. As a result, $[H_3O^+] = [OH^-]$ in pure water. Because both of these concentrations are the same, we can solve for the equilibrium concentrations using the value of K_w. Let x be the concentration of H_3O^+ (and therefore also the concentration of OH^-):

$$\begin{aligned} \mathrm{K_w} &= [\mathrm{H_3O^+}][\mathrm{OH^-}] \\ 1.0 \times 10^{-14} &= [x][x] \\ 1.0 \times 10^{-14} &= x^2 \\ 1.0 \times 10^{-7} &= x \end{aligned}$$

At equilibrium, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ in a sample of pure water. An aqueous solution in which $[H_3O^+] = [OH^-]$ is referred to as a **neutral solution**. However, the addition of an external acid or base will shift the relative amounts of these two ions. Adding an acid will increase the amount of H_3O^+. As a consequence, the amount of OH^- will need to decrease in order to re-establish equilibrium (at which point the equilibrium expression for K_w will once again have the correct value). In an **acidic solution**, $[H_3O^+] > [OH^-]$. Similarly, in a **basic solution**, the amount of OH^- will increase and the amount of H_3O^+ will decrease, so $[H_3O^+] < [OH^-]$.

Shorthand Notation for Aqueous Acids

It is very common for chemists to write H^+ instead of H_3O^+ when talking about aqueous solutions of acids and bases. However, H^+ will not exist as an isolated ion if dissolved in water. Instead, it will be closely associated with (at least) one molecule of the solvent. It is generally acceptable to use H^+, but it should be understood that this is just a shorthand notation for H_3O^+. Consequently, the expression for K_w is often written as follows:

$Kw=[H^+][OH^-]$

Acid-base reactions are also sometimes written in a way that makes use of this shorthand. For example, you might see the following acid-dissociation reaction:

$HNO_3(aq) \rightarrow H^+(aq) + NO{-}3(aq)$

However, a better description of this process would be the following:

$HNO_3(aq) + H_2O(l) \rightarrow H_3O+(aq) + NO^-_3(aq)$

H^+ is not just "falling off" of nitric acid. Instead, it is being pulled off by water, which is a better base than the resulting nitrate anion. Again, the shorthand version is acceptable to use, but keep in mind that a more accurate description would include water as a base whenever "H^+" is being generated in an aqueous solution.

pH of Aqueous Solutions

Because we are dealing with such small concentrations of H^+ and OH^-, a system was invented in order to talk about the acidity or basicity of a solution that uses more manageable numbers. The Danish chemist Søren Sørenson proposed a new quantity that he called pH, which is defined as follows:

$pH = -log[H^+]$

One of the confusing consequences of pH values as defined by this equation is that the higher the hydrogen ion concentration, the lower the pH. A solution with a hydrogen ion concentration of 1×10^{-3} would have a pH of 3, and a solution with a hydrogen ion concentration of 1×10^{-5} would have a pH value of 5.

Example

What is the pH of a neutral solution (at 25°C)?

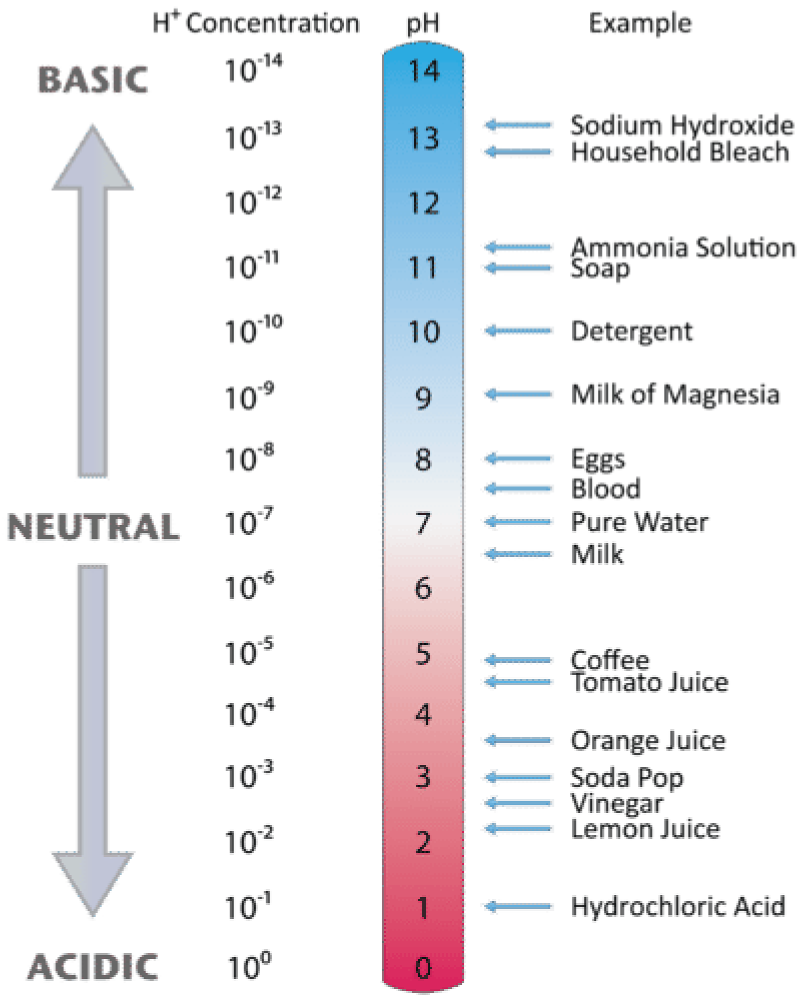
Answer:

As we saw in the previous section, a neutral solution has a hydrogen ion concentration of 1.0×10^{-7}.

$$pH = -log[H^+]$$
$$pH = -log[1.0 \times 10^{-7}]$$
$$pH = 7.00$$

A neutral solution has a pH of 7. Acidic solutions have higher concentrations of H^+, so they have pH values that are less than 7. Conversely, basic solutions have lower concentrations of H^+ and pH values greater than 7.

Here is a list of the pH values for some common acidic and basic solutions:



CK-12 Foundation

Example

What is the concentration of hydrogen ions in a solution that has a pH value of 4.67?

Answer:

Start with the definition of pH and plug in the known value:

$$\mathrm{pH} = -\log[\mathrm{H}^+]$$
$$4.67 = -\log[\mathrm{H}^+]$$

The logarithm function can be inverted as follows:

The hydrogen ion concentration in this solution is 2.1×10^{-5} M.

$$
\begin{aligned}
-\log[\mathrm{H^+}] &= 4.67 \\
\log[\mathrm{H^+}] &= -4.67 \\
10^{\log[\mathrm{H^+}]} &= 10^{-4.67} \\
[\mathrm{H^+}] &= 10^{-4.67} \\
[\mathrm{H^+}] &= 2.1 \times 10^{-5}
\end{aligned}
$$

Questions: Copy and paste question 5 in the Submit Box at the bottom of this page, and answer the question before going any further in the lesson:

5. How do the concentrations of H_3O^+ and OH^- compare in an acid? In a base?

Strong vs. Weak Acids

So far, we have primarily been defining acids by their ability to donate an H^+ ion and bases by their ability to accept an H^+ ion. However, acids and bases vary in their relative ability to undergo these processes.

In general, acids can be classified as strong or weak based on the extent to which they produce H_3O^+ when dissolved in water. For a generic acid, we can write the following equilibrium reaction:

$HA(aq)+H_2O(l)\rightleftarrows H_3O^+(aq)+A^-(aq)$

Using the usual shorthand notation, this equation can also be written as follows:

$HA(aq)\rightleftarrows H^+(aq)+A^-(aq)$

This type of equilibrium, in which a proton is being transferred to water, is often indicated by writing the equilibrium constant as Ka. The relative position of this equilibrium for a given acid determines whether it will be considered strong or weak. When dissolved in water, a strong acid will completely transfer its proton to the solvent. In terms of the equilibrium above, the products will be heavily favored ($K_a >> 1$). In fact, the products are so heavily favored that the reverse reaction is often not even considered, and the proton transfer is written as unidirectional. For example, the strong acid HCl can dissociate in water according to the following reaction:

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

$(HCl(aq) \rightarrow H^+(aq)+Cl^-(aq))$

At equilibrium, essentially no intact HCl molecules are still present in solution.

In contrast, the equilibrium for a weak acid favors the reactants. A particularly common type of weak acid is an organic molecule that contains a carboxyl group (CO_2H, sometimes written as COOH). For example, acetic acid (the acidic component of vinegar) has the formula CH_3CO_2H. Its dissociation equation can be written as follows:

$CH_3CO_2H(aq)+H_2O(l) \rightleftarrows H_3O+(aq)+CH_3CO^-{}_2(aq)$

$(CH_3CO_2H(aq) \rightleftarrows H^+(aq)+CH_3CO^-{}_2(aq))$

Because we are dealing with a weak acid, Ka for this equilibrium is much less than 1. At equilibrium, most of the acetic acid molecules are still intact, and only a small percentage have transferred their protons to the solvent. The Ka values for some weak acids are listed in the table below:

Acid Name	Structure	K_a
hydrofluoric acid	H-F	7.1×10^{-4}
nitrous acid	O=N-O-H	4.5×10^{-4}
formic acid	HCOOH	1.7×10^{-4}
acetic acid	CH_3COOH	1.8×10^{-5}
hydrocyanic acid	H^-CN	4.9×10^{-10}

Strong vs. Weak Bases

Analogous to the acid dissociation reaction from the previous section, we can write the reaction between a generic base and water as follows:

$B(aq)+H_2O(l) \rightleftarrows BH^+(aq)+OH^-(aq)$

The equilibrium constant for a reaction in which a base is deprotonating water (taking water's hydrogen atom) is often given the symbol Kb. Strong bases and weak bases can then be defined based on the position of this equilibrium. A weak base would have a very small K_b value (much less than 1), indicating that most molecules of the base do not remove a proton from water. Conversely, a strong base would have a K_b value greater than or equal to 1.

Nitrogen-containing compounds are a common type of weak base. The lone pair on the nitrogen atom can accept a proton from water as follows:

$NH_3(aq)+H_2O(l) \rightleftarrows NH^+_4(aq)+OH^-(aq)$

The equilibrium constant for this reaction is quite low, so most of the NH3molecules will not remove a proton from water. Kb values for a few nitrogen-containing bases are listed in the table below:

Base	K_b
ethylamine ($CH_3CH_2NH_2$)	5.6×10^{-4}
methylamine (CH_3NH_2)	4.4×10^{-4}
ammonia (NH_3)	1.8×10^{-5}

The only strong bases that are commonly used in general chemistry courses are ionic compounds composed of metal cations and hydroxide anions, such as NaOH, KOH, or $Ba(OH)_2$.

Calculating pH for Acidic and Basic Solutions

Strong Acids and Bases

In the case of strong acids and bases, the pH for a solution of known concentration is relatively easy to calculate. For example, the strong acid HCl will dissociate completely, so we assume that the amount of acid added to the solution is equal to the amount of H^+ present at equilibrium.

Example

What is the pH of a 0.150 M aqueous solution of HCl?

Answer:

Because HCl is a strong acid, we can assume that all of its acidic hydrogens are transferred to the solvent molecules. Therefore, $[H^+] = 0.150$ M, so pH can be calculated as follows:

$pH = - \log [H^+]$

$pH = - \log [0.150]$

$pH = 0.82$

As expected for an acidic solution, the pH is much lower than 7.

Because the strong bases that you will encounter are all ionic compounds that contain the hydroxide anion, you can assume complete dissociation in water, which would tell you the concentration of hydroxide. Then, the concentration of H^+ can be calculated using the expression for K_w.

Example

What is the pH of a 0.245 M aqueous solution of NaOH?

Answer:

Because NaOH is a soluble ionic compound, we can assume that it fully dissociates in water. After dissociation, $[OH^-]$ = 0.245 M. We can then use the value of K_w to determine $[H^+]$.

$K_w=[H^+][OH^-]$

$1.0\times10^{-14}=[H^+][0.245]$

$[H^+]=4.08\times10^{-14}$

Then, use the definition of pH to determine the pH of this solution:

pH = - log $[H^+]$

pH = - log [4.08 × 10^{-14}]

pH =13.39

As expected for a basic solution, this value is significantly larger than 7.

Weak Acids and Bases

Because only a small portion of any available weak acid or base molecules undergo a proton transfer to form either H^+ or OH^- ions, calculating the pH of one of these solutions is slightly more complicated. The following example problem outlines the general strategy for answering this type of question.

Example

The K_a of acetic acid (CH_3CO_2H) is 1.8×10^{-5}. Calculate the pH of a 0.50 M solution of acetic acid.

Answer:

First, set up an ICE table for the acid dissociation equation. Before any proton transfers occur, we have acetic acid at a concentration of 0.50 M, no acetate anion, and an H+ concentration of 1.0×10^{-7} M. Once the reaction begins, some of the acetic acid will be converted to acetate as it transfers an H+ ion to the solvent. The amount of acetic acid that dissociates in order to reach equilibrium (the quantity you are trying to find) is represented by x. A corresponding increase then occurs in the concentrations of the products. The equilibrium concentrations of each species can then be written in terms of x.

	CH_3CO_2H	⇄	H^+	+	$CH_3CO_2^-$
Initial concentration (M)	0.50		1.0×10^{-7}		0.00
Change (M)	-x		+x		+x
At equilibrium (M)	0.50 – x		1.0×10^{-7} + x		x

Then, write the equilibrium constant expression and plug in the equilibrium values from the ICE table.

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

$$1.8 \times 10^{-5} = \frac{[1.0 \times 10^{-7} + x][x]}{[0.50 - x]}$$

Solving this equation by hand would be quite difficult. Fortunately, for essentially all of the weak acid problems that you will be expected to solve, we can make two simplifying assumptions. First, we assume that only a very small percentage of the acetic acid molecules will be ionized, which means that the drop in concentration (x) is much smaller than the original concentration (0.50 M). This is reasonable based on the small value of K_a. Mathematically, this assumption means that (0.50 - x) is approximately equal to 0.50. For example, say we found that x had a value of 0.0012. If we subtract this value from 0.50, we get 0.4988, and rounding to the correct number of significant figures would give us 0.50. Based on this assumption, we can simplify the above equation as follows:

$$1.8 \times 10^{-5} = \frac{[1.0 \times 10^{-7} + x][x]}{[0.50]}$$

The second assumption is that the amount of H^+ produced by this reaction is much larger than the amount already present. This is reasonable because even a weak acid tends to be much more acidic than pure water (unless it is present in extremely low concentrations). Mathematically, this means that we can ignore the 1.0×10^{-7} M H+ already present. For example, if x has a value of 0.0012, the quantity (1.0×10^{-7} + x) is equal to x after rounding to the correct number of significant figures. This second assumption further simplifies the equation above, and we can now solve for x:

$$1.8 \times 10^{-5} = \frac{[x][x]}{[0.50]}$$

$$9.0 \times 10^{-6} = x^2$$

$$x = 0.0030$$

The concentration of H+ at equilibrium is equal to 1.0×10^{-7} + 0.0030 M, which after rounding is simply 0.0030 M. Then, find the pH as usual:

$$\text{pH} = -\log[\text{H}^+]$$

$$\text{pH} = -\log[0.0030]$$

$$\text{pH} = 2.52$$

A 0.50 M solution of acetic acid would have a pH of 2.52.

Grading Rubric:

***Note:** Your grade will be calculated by the sum of the points earned for each question. Points are earned according to the chart below.

To get a 10: A total score of 10 upon the first submission or after the first revision.

To get a 9: A total score of 9 after the first revision.

To get an 8: A total score of 8 after the first revision.

To get a 7: A total score of 7 after the first revision.

To get a 6: A total score of 6 after the first revision.

To get a 5: Any score lower than a 6; Plagiarism – purposeful or mistaken, which will lower your final grade for the course (So, be very careful when posting your work!); lack of effort, disrespect, or attitude. Lesson requirements have not been met.

Stop and Think (Total content points possible = 2.5)	Answer is clearly written and accurate. Answer is based on the lesson content. **.5 point each**		Answer is not accurate or is not based on lesson content. **0 points each**
Short Answer (Total content points possible = 1)	Answer is clearly written and accurate. Answer is based on lesson content. **.5point each**	Answer is clearly written and but may have 1 factual omission or error. **.25 point each**	Answer is not clearly written. There are several factual omissions. **0 points each**
Calculate and Show Your Work (Total points possible = 3	Answer is correct and calculations are complete and accurate. **.5 point each**	Answer is correct, but calculations are incomplete or Answer is incorrect, but calculations are complete and accurate. **.25 point each**	Answer is incorrect and calculations are incomplete. **0 points each**
Apply Your Knowledge Total points possible 1.5 points	Answer is clearly written and accurate. **1.5 points**	Answer is clearly written and accurate. May have 1 factual omission or error. **1 point**	Answer is not clearly written. There are several factual omissions. or **0 points**
Closure Total points possible 2 points	Answer is clearly written and accurate. **2 points**	Answer is clearly written and but may have 1 factual omission or error. **1 point**	Answer is not clearly written. There are several factual omissions. **0 points**

Assignment:

Stop and Think (Questions 1-5)

Copy and paste the Stop and Think Questions found throughout the lesson content and answer them in the submit Box below. Stop and Think questions should be based on the lesson content. You will not do outside research for these questions.

Short Answer Answers can be found in the lesson content. You will not conduct outside research.

6. What definition is most widely applicable in defining acids and bases?

7. Describe the difference between a monoprotic acid and a polyprotic acid.

Calculate and Show your steps:

8. Calculate the pH of the following solutions. Show all your calculations. Assume complete dissociation of each strong acid ([Acid] = [H^+]):

a. 4.5×10^{-3}M HBr

b. 1.34×10^{-4}M HCl

c. 7.98×10^{-2}M HNO_3

9. Calculate the hydrogen ion concentrations in each of the following solutions. Show all your calculations.

a. pH = 1.05

b. pH = 5.65

c. pH = 2.42

Apply Your Knowledge

10. A 0.10 M solution of a weak acid gave a pH reading of 4.2. Calculate the hydrogen ion concentration, then determine K_a for the unknown acid. Show all of your calculations for full credit!

Closure

9. Based on the pH measurements for each substance in the PLIX, calculate the [H^+] ion concentration for lemon juice, soap, and water.