

Chapter 11: Acids and Bases

For our purposes, an **acid** is a substance that produces hydrogen ion (H^+) when dissolved in water. A **base** is a substance that produces hydroxide ion (OH^-) when dissolved in water.

These are the Arrhenius definitions of acids and bases, named for Swedish chemist Svante August Arrhenius (1859 – 1927), who first defined them. Arrhenius proposed this definition of acids and bases in his 150-page doctoral thesis presented in 1884. Among the many ideas he proposed, the most important was that neither salts nor pure water conducted electricity, but that salt solutions did conduct electricity. The explanation was that salt dissociated into anions and cations. While we accept this idea as obvious today, in 1884 it was outrageous, and Arrhenius barely received a passing mark on his work. In 1903, Arrhenius won the Nobel Prize in Chemistry (the third Nobel Prize ever given) for the work reported in his thesis.

There is another, complimentary, definition of acids/bases – Brønsted-Lowry acids/bases. This definition is named for Danish chemist Johannes Nicholar Brønsted (1879 – 1947) and English chemist Thomas Martin Lowry (1874 – 1936). Working independently, these two scientists developed essentially the same definition of acids and bases. A Brønsted-Lowry acid donates hydrogen ion, while a Brønsted-Lowry base accepts hydrogen ion. Brønsted-Lowry acid-base theory is more general than Arrhenius, and all Arrhenius acids and bases are also Brønsted-Lowry bases. However, the reverse is not true.

There are several different ways of classifying acids and bases. Two of these classifications are: is the acid (or base) **strong** or **weak**, and is the acid (or base) organic or inorganic?

An acid (or base) is strong if ~100% of the acid (or base) molecules ionize to form hydrogen ion (or hydroxide ion) in water. If the acid (or base) is weak, then << 100% of the acid (or base) molecules ionize to form hydrogen ion (or hydroxide ion) in water. Normally, the concentration of acid or base is set at 0.10M when this comparison is made.

Acids and bases tend to be either strong or *very* weak. One of the strongest of the weak acids is acetic acid, $\text{CH}_3\text{CO}_2\text{H}$. In water (0.10M concentration), it ionizes ~1%.

Organic acids are composed of carbon, hydrogen, and oxygen. The acid part of the organic acid molecule is called a carboxylic acid group (CO_2H or COOH). When organic acids ionize, they lose hydrogen ion from the carboxylic acid group – not from any other location in the molecule! Organic acids are covered in more detail in organic chemistry.

Inorganic acids are generally composed of nonmetallic elements. The polyatomic ions you memorized (or should have) are portions of the inorganic acids. Adding hydrogen ion to the polyatomic ions produces the inorganic acids. This subject is also covered in Chapter 4.

All organic acids, and almost all organic bases are classified as weak. Many inorganic acid and bases are also weak. This reduces the total number of strong acids and bases to a manageable number. **MEMORIZE THESE STRONG ACIDS AND BASES** (Table 11.1). Any acid (or base) not classified as strong is a weak acid (or base) by default.

Strong Acids:		Strong Base:	
<u>Formula</u>	<u>Name</u>	<u>Formula</u>	<u>Name</u>
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO ₃	Nitric acid	RbOH	Rubidium hydroxide
HClO ₄	Perchloric acid	CsOH	Cesium hydroxide
H ₂ SO ₄	Sulfuric acid	Ca(OH) ₂	Calcium hydroxide

Table 11.1. Strong acids and strong bases.

Properties of acids and bases.

Acids share a set of common properties. Probably the first of these properties observed was taste – acids tend to have a sour taste. This property was likely first observed with acetic acid. When people made wine, over time the wine would sour. The alcohol slowly oxidized to acetic acid causing souring. It was this same slow oxidation of alcohol to acetic acid that was used to make vinegar (which is just a dilute solution of acetic acid in water, with some trace substances).

Let me be completely clear about the “sour taste of acids” property:

**NEVER TASTE ANYTHING IN THE CHEMISTRY
LAB UNLESS YOU WANT TO DIE!**

The history of chemistry is full of scientists who have died from tasting their experiments. This is not an experiment that needs to be replicated, and you don't need to imitate them.

Acids tend to change the colors of certain kinds of dyes used as *indicators*: a substance that changes color with acids or bases. One widely used indicator is litmus. Litmus is blue with bases, and red with acids.

Acids react with metals to produce hydrogen gas, and bubbles of this flammable gas can be seen as this reaction proceeds.

Bases tend to have a bitter taste and a slippery, soapy feel on the skin. In addition to the warning I gave you about tasting things in the chemistry lab, let me add another warning:

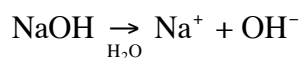
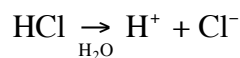
**NEVER RUB ANYTHING ON YOUR SKIN IN THE
CHEMISTRY LAB UNLESS YOU WANT TO BE
PERMANENTLY SCARRED AND MAYBE DIE!**

The slippery, soapy feel of bases was probably the first base property observed, and was noticed when people discovered how to clean wool before dyeing it different colors. Wool “fresh off of the sheep” is dirty and oily and has to be cleaned before colored dyes will stick to the wool fibers. Soaking wool in a mixture of wood ashes in water removed the oil, allowing the dyes to stick to the wool. Eventually, people learned to make *soap* by adding wood ashes (as a source of base) to melted animal fat. It was many years before people routinely used soap and water for personal hygiene, but soap was routinely used for cleaning wool and other fibers before dyeing.

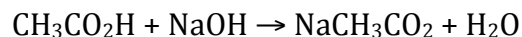
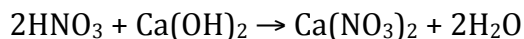
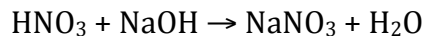
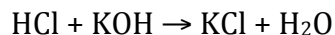
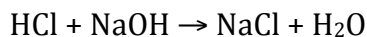
Bases also change indicator colors, and if red litmus is combined with base the litmus become blue.

Blue litmus or red litmus is sometimes soaked into strips of paper to make *litmus paper*, which is more convenient to use than the liquid solutions. Other dyes are sometimes put on paper to make pH paper, allowing us to estimate the actual pH of a solution instead of simply saying, “It’s acid” or “It’s base”.

Acids make water acidic by increasing the amount of hydrogen ion present in solution; bases make water basic by increasing the amount of hydroxide ion present in solution:

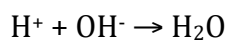


When acids and bases are combined, hydrogen ion reacts with hydroxide ion to form water. Consider the following reactions between various acids and bases:



In all of these reactions, an acid reacts with a base to produce a salt and water. This is the general pattern of acid – base reactions: acids and bases combine to form salts and water.

The reaction between acids and bases is called **neutralization**, represented by the reaction:



When acids and bases react, the acids lose their acidic properties and the bases lose their basic properties.

Water and pH.

Danish chemist Søren Peder Lauritz Sørensen (1868 – 1939) introduced the concept of pH at Carlsberg Laboratory in 1909. According to the Carlsberg Foundation, pH stands for “power of hydrogen”. By strict definition, pH is related to the activity of hydrogen ion:

$$\text{pH} = -\log(\alpha_{\text{H}^+})$$

However, it is more common to define pH as the negative log of the molar concentration of hydrogen ion:

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

We will use this more common definition in this chapter.

In water, the concentration of hydrogen ion and hydroxide ion are related by the equation:

$$K_w = 1.00 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$$

Where the notation $[H^+]$ and $[OH^-]$ indicates the equilibrium molar concentrations (molarity) of hydrogen ion and hydroxide ion. In water, hydrogen ion and hydroxide ion concentrations are inversely related; if one increases the other must decrease. Table 11.2 illustrates this relationship over the pH range of 1 to 13.

Look closely at the values given in Table 11.2. At any pH value, the product of hydrogen ion and hydroxide ion concentration equals 1.00×10^{-14} . For any given pH, we can readily calculate hydrogen ion concentration:

$$[H^+] = 10^{-\text{pH}}$$

In water, the pH defines not only the concentration of hydrogen ion but also the concentration of hydroxide ion.

pH	$[H^+]$	$[OH^-]$
1.00	1.0×10^{-1}	1.0×10^{-13}
2.00	1.0×10^{-2}	1.0×10^{-12}
3.00	1.0×10^{-3}	1.0×10^{-11}
4.00	1.0×10^{-4}	1.0×10^{-10}
5.00	1.0×10^{-5}	1.0×10^{-9}
6.00	1.0×10^{-6}	1.0×10^{-8}
7.00	1.0×10^{-7}	1.0×10^{-7}
8.00	1.0×10^{-8}	1.0×10^{-6}
9.00	1.0×10^{-9}	1.0×10^{-5}
10.00	1.0×10^{-10}	1.0×10^{-4}
11.00	1.0×10^{-11}	1.0×10^{-3}
12.00	1.0×10^{-12}	1.0×10^{-2}
13.00	1.0×10^{-13}	1.0×10^{-1}

Table 11.2. Molar concentrations of hydrogen ion and hydroxide ion, in water, as a function of pH.

We define **pOH** similarly to pH; $\text{pOH} = -\log[OH^-]$. For water solutions, all we need is one value of pH, or pOH, or one concentration for either hydrogen ion or hydroxide ion, to calculate all other values. For example, imagine I have a water solution with $\text{pH} = 4.5$:

$$[\text{H}^+] = 10^{-4.5} = 0.000\ 032 = 3.2 \times 10^{-5} M$$

$$[\text{OH}^-] = \frac{1.00 \times 10^{-14}}{3.2 \times 10^{-5}} = 3.1 \times 10^{-10} M$$

$$\text{pOH} = -\log[3.1 \times 10^{-10}] = 9.5$$

We can also use a simpler formula to relate pH and pOH:

$$\text{pH} + \text{pOH} = 14.00$$

At $\text{pH} < 7$, hydrogen ion concentration is greater than hydroxide ion concentration. Aqueous solutions at $\text{pH} < 7$ are described as “acidic”. At $\text{pH} > 7$, hydroxide ion concentration is greater than hydrogen ion concentration, and these solutions are described as “basic”. A pH of 7.00 indicates that hydrogen ion and hydroxide ion concentrations are equal; $\text{pH} = 7.00$ is described as “neutral”.

Be extremely careful in how you use and think about the term “neutral”. A neutral pH ($\text{pH} = 7.00$) means that the concentrations of hydroxide and hydrogen ion are equal to each other; it absolutely positively does NOT mean that there isn't any hydrogen ion or hydroxide ion! Pure water spontaneously breaks apart into hydrogen ions and hydroxide ions:



This is an equilibrium reaction, and the equilibrium constant is 1.00×10^{-14} . This equilibrium is sufficiently important to have it's own distinctive equilibrium constant symbol: K_w .

If stoichiometric amounts of strong acids (HCl , HNO_3 , H_2SO_4) are combined with strong bases (NaOH , KOH), then the solution pH is very close to 7.0. We make the general statement that “Strong acids neutralize strong bases producing a neutral pH ($\text{pH} \sim 7.00$).” However, the situation is very different for the reaction of a weak acid with a strong base, or a strong acid with a weak base, or a weak acid with a weak base. When stoichiometric amounts of strong/weak acids and bases are combined, the solution pH is not 7.00. Conversely, if strong/weak acids and bases are combined until the pH equals 7.00, then stoichiometry hasn't been achieved. The type of acid and base combined tells us the resulting pH of the solution, and the type of salt produced (Table 11.3).

Acid type	Base type	Salt type	Solution pH (stoichiometric)
Strong	Strong	Neutral	~ 7.0
Strong	Weak	Acid salt	< 7.0
Weak	Strong	Base salt	>7.0
Weak	Weak	?????????	?????????

Table 11.3. Salt types and solution pH for stoichiometric combinations of various acid-base types.

The salts produced from combining strong acids and bases (NaCl, NaNO₃, Na₂SO₄, KCl, KNO₃, K₂SO₄) are **neutral salts**. If these neutral salts are dissolved in water, the solution pH is close to 7.0.

Strong acids react with weak bases, such as ammonia (ammonium hydroxide). When stoichiometric amounts of strong acids and weak bases are combined, the solution pH is less than 7.0! Salts such as NH₄Cl, NH₄NO₃, (NH₄)₂SO₄ are **acid(ic) salts**, and the solution pH of these salts is less than 7.0.

Weak acids react with strong bases, and when stoichiometric amounts are combined the resulting solution pH is greater than 7.0. The salts produced, such as CH₃CO₂Na, are **base (basic) salts**, and the solution pH of these salts is greater than 7.0.

For combinations of weak acids with weak bases, no general result can be predicted.

Acid-Base Titrations.

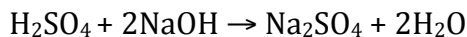
An acid-base **titration** is a laboratory procedure involving the addition of measured volumes of an acid (or base) of known concentration to a known volume of base (or acid) for the purpose of determining the concentration of the base (or acid).

The most useful equation for acid-base calculations is a version of the dilution equation:

$$n_a M_a V_a = n_b M_b V_b$$

Where M_a and M_b are the concentrations of the acid and base respectively, V_a and V_b are the volumes of the acid and base, and n_a and n_b are the number of hydrogen ions per molecule of acid, and the number of hydroxide ions per molecule of base.

Consider a titration of sulfuric acid (H_2SO_4) with sodium hydroxide. If 25.00 mL of sodium hydroxide requires 18.35 mL of 0.09345M sulfuric acid for neutralization, what is the concentration of sodium hydroxide? The chemical reaction is:



Notice that sulfuric acid provides 2 hydrogen ions for every molecule of sulfuric acid, while sodium hydroxide provides one hydroxide per “molecule” of sodium hydroxide. Substituting the known values into our equation gives us:

$$2 \times 0.09345\text{M} \times 18.35 \text{ mL} = 1 \times M_b \times 25.00 \text{ mL}$$

$$M_b = 0.1372\text{M}.$$

When stoichiometry is achieved between the acid and the base, then the moles of acid must equal the moles of base (moles of H^+ = moles OH^-). If we chose to do so, we could solve acid-base calculations using the same techniques used for any other stoichiometry calculation. For this reaction, we can readily calculate that we have 0.001714808 moles of sulfuric acid (0.09345 moles/liter \times 0.01835 L). From the reaction stoichiometry, 2 moles of sodium hydroxide are required for each mole of sulfuric acid. Our 0.001714808 moles of sulfuric acid require 0.003429616 moles of sodium hydroxide. The concentration of sodium hydroxide is 0.003429616 moles of sodium hydroxide/0.02500 L = 0.1372M.

One common titration procedure is **standardization**. This process involves titrating a known amount of a pure, standard material with an acid (or base) for the purpose of determining the concentration of the acid (or base). Once the concentration of the acid (or base) has been determined, it can be used to accurately titrate other substances. Sodium hydroxide solution is typically standardized by titrating a pure form of potassium hydrogen phthalate (abbreviated KHP, Figure 11.1).

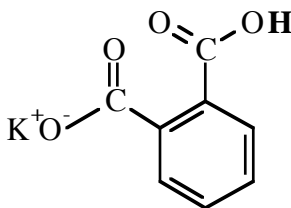


Figure 11.1. Potassium hydrogen phthalate, KHP. The bold-faced hydrogen is the hydrogen ion released by this acid.

KHP has a formula weight of 204.23 grams/mole, and there is one hydrogen ion produced (shown in bold) for each molecule of the acid.

Imagine we have a sodium hydroxide solution of unknown concentration. We weigh a portion of KHP (0.2503 grams), dissolve it in some deionized water, and titrate it with the sodium hydroxide solution until the acid is neutralized, requiring 22.39 mL of sodium hydroxide. What is the concentration of sodium hydroxide solution?

$$\begin{aligned} \text{moles KHP} &= \frac{\text{grams KHP}}{F.W. \text{ KHP}} \\ &= \frac{0.2503\text{g}}{204.23\text{g / mole}} \\ &= 0.001\ 225\ 579 \text{ moles KHP} \end{aligned}$$

When neutralization is achieved, the moles of sodium hydroxide must equal the moles of KHP. The concentration of sodium hydroxide is:

$$[\text{NaOH}] = \frac{0.001\ 225\ 579 \text{ moles NaOH}}{0.02239\text{L}} = 0.05474\text{M}$$

There are several methods used to indicate the end of a titration, but one common method involves using an *indicator*. An indicator is a dye that changes color at specific pH values. Phenolphthalein (Figure 11.2) is a useful indicator for acid-base titrations because of its distinctive color change; in acid solutions, phenolphthalein is colorless while in base solutions it is pink. Generally, 2 or 3 drops of 1% phenolphthalein solution is added to the acid solution. When the solution turns pink, all of the acid has been neutralized and we can stop titrating.

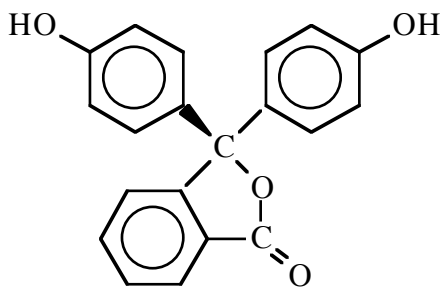
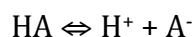


Figure 11.2. Phenolphthalein indicator.

Buffers.

A **buffer** is a solution containing a weak acid and a salt of the **same** weak acid. An example of a buffer solution is an aqueous mixture of acetic acid and sodium acetate ($\text{CH}_3\text{CO}_2\text{H}$ and $\text{CH}_3\text{CO}_2\text{Na}$). A mixture of acetic acid and sodium chloride is **NOT** a buffer solution. Buffer solutions have the ability to **resist** pH changes when small amounts of strong acids or bases are added to the solution.

A weak acid does not ionize 100 % as does a strong acid. When a weak acid dissolves in water, equilibrium between the ionized and unionized forms of the acid is rapidly reached, described by the reaction:



Where HA is the unionized acid and A^- is the ionized form of the acid. The equilibrium expression for this reaction is:

$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

Where K_a is the equilibrium constant, and $[\text{H}^+]$, $[\text{A}^-]$, and $[\text{HA}]$ are the equilibrium concentrations of hydrogen ion, ionized acid, and unionized acid respectively. The above expression can be rearranged, producing the **Henderson-Hasselbach equation**:

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

Where $\text{p}K_a = -\log K_a$.

The equation was first written in 1908 by American physiologist and chemist, Lawrence Joseph Henderson (1878 – 1942). Henderson is considered by many as one of the leading biochemists of the 20th century, and he wrote the equation to describe the carbonate buffer system. In 1916, the equation was re-written in its logarithmic form by Danish physician and chemist Karl Albert Hasselbach (1874 – 1962). Hasselbach was a pioneer in the medical use of pH measurements, and used to equation to study a condition called *metabolic acidosis*, where the body produces too much acid, reducing the bloods pH.

Consider a solution containing 0.0500M acetic acid and 0.0250M sodium acetate. The K_a of acetic acid is 1.8×10^{-5} ; therefore the $\text{p}K_a$ is 4.74. Plugging these values into the Henderson-Hasselbach equation we get:

$$\begin{aligned} \text{pH} &= 4.74 + \log \frac{[0.0250M]}{[0.0500M]} \\ &= 4.74 + (-0.30) \\ &= 4.44 \end{aligned}$$

What if we added a small amount of strong base (NaOH) to the buffer described above? When sodium hydroxide is added to the buffer, some of the unionized acid will be neutralized and converted to the ionized form. The concentration of HA will decrease, and the concentration of A⁻ will increase. The absolute value of the increase and decrease will be the same as the total moles of sodium hydroxide added.

Let's say that we started with 100.0 mL of the buffer. Before adding any sodium hydroxide we have:

$$\begin{aligned} 0.0250M \times 0.1000L &= 0.00250 \text{ moles A}^- \\ 0.0500M \times 0.1000L &= 0.00500 \text{ moles HA} \end{aligned}$$

Imagine we add 5.00 mL of 0.1000M sodium hydroxide. This volume of sodium hydroxide solution corresponds to:

$$0.00500L \times 0.1000M = 0.000500 \text{ moles sodium hydroxide.}$$

The A⁻ concentration will increase by 0.000500 moles, and the HA concentration will decrease by the same amount. The total volume of the solution increases by 5.00 mL. The new concentrations of A⁻ and HA are:

$$\begin{aligned} [A^-] &= (0.00250 + 0.000500)/0.1050L = 0.0286M \\ [HA] &= (0.00500 - 0.000500)/0.1050L = 0.0429M \end{aligned}$$

The resulting pH is:

$$\begin{aligned} \text{pH} &= 4.74 + \log \frac{[0.0286M]}{[0.0429M]} \\ &= 4.74 + (-0.18) \\ &= 4.56 \end{aligned}$$

The pH has increased by 0.12 pH units.

Compare this pH change with a similar situation using plain water instead of buffer solution. Plain water has a pH very close to 7. Adding 5.00 mL of 0.1000M sodium hydroxide solution results in a final sodium hydroxide concentration of $0.000\ 500\ \text{moles}/0.1050\text{L} = 0.00476\text{M}$. The pOH is 2.32 so the pH is 11.68. This is a change of 4.68 units. Clearly, pH changes are smaller in buffer solutions compared to pH changes in ordinary water.

A similar calculation shows that addition of small amounts of strong acid changes the buffer pH only slightly. When adding acid, the concentration of unionized weak acid increases, while the concentration of the ionized form decreases; hydrogen ion combines with the ionized acid (A^-) to form more of the unionized acid (HA). Using the same starting buffer information from above:

$$\begin{aligned} 0.0250\text{M} \times 0.1000\text{L} &= 0.00250\ \text{moles } \text{A}^- \\ 0.0500\text{M} \times 0.1000\text{L} &= 0.00500\ \text{moles HA} \end{aligned}$$

Imagine we add 5.00 mL of 0.1000M HCl. This volume of HCl corresponds to:

$$0.00500\text{L} \times 0.1000\text{M} = 0.000\ 500\ \text{moles hydrochloric acid.}$$

The new concentrations of A^- and HA will be:

$$\begin{aligned} [\text{A}^-] &= (0.002\ 50 - 0.000\ 500)/0.1050\text{L} = 0.0190\text{M} \\ [\text{HA}] &= (0.005\ 00 + 0.000\ 500)/0.1050\text{L} = 0.0524\text{M} \end{aligned}$$

The resulting pH is:

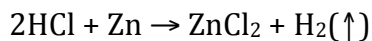
$$\begin{aligned} \text{pH} &= 4.74 + \log \frac{[0.0190]}{[0.0524\text{M}]} \\ &= 4.74 + (-0.44) \\ &= 4.30 \end{aligned}$$

The pH decreased by 0.14 pH units compared to the starting value of 4.44.

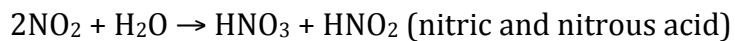
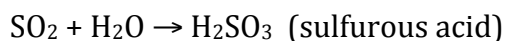
By comparison, adding 5.00 mL of 0.1000M HCl to 100.0 mL of plain water produces a hydrogen ion concentration of 0.00476M. The pH of this solution is 2.32, much lower than the starting pH, 7.0.

Other general reactions.

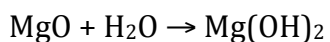
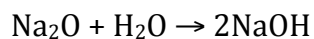
Acids react with many metals to produce hydrogen gas. The reaction below is representative of these types of reactions.



Acid anhydrides, which are typically non-metal oxides such as CO_2 , SO_2 , NO_2 , etc., react with water to form acids:



Metal oxides generally react with water to produce metal hydroxides (bases). For example, sodium oxide or magnesium oxide react with water to produce sodium hydroxide or magnesium hydroxide respectively:



Chapter 11 Homework:

Vocabulary. The following terms are defined and explained in the text. Make sure that you are familiar with the meanings of the terms as used in chemistry. Understand that you may have been given incomplete or mistaken meanings for these terms in earlier courses. The meanings given in the text are correct and proper.

Acid	Base	Strong
Weak	Indicator	Litmus paper
Buffer	pH	pOH
Neutral salts	Acid(ic) salts	Base(ic) salts
Titration	Standardization	Henderson-Hasselbach equation
Neutralization		

Perform the indicated acid-base calculations.

1. How much phosphoric acid, H_3PO_4 , solution (0.05534M) must be added 25.00 mL of 0.01095M calcium hydroxide, $\text{Ca}(\text{OH})_2$, to completely neutralize the solution?
2. A sample of benzoic acid, $\text{C}_7\text{H}_6\text{O}_2$, weighing 0.1993 g and containing one hydrogen ion per molecule of acid, required 18.72 mL of sodium hydroxide for neutralization. What is the concentration of sodium hydroxide?
3. Which solution requires more 0.05534M potassium hydroxide, KOH, for neutralization: a) 25.00 mL of 0.0878M nitric acid, HNO_3 or b) 7.50 mL of 0.0250M phosphoric acid?
4. Which solution requires more 0.1003M acetic acid ($\text{CH}_3\text{CO}_2\text{H}$, one hydrogen ion per molecule) for neutralization: a) 25.00 mL of 0.09938M potassium hydroxide or b) 10.00 mL of 0.05535M calcium hydroxide?
5. A 15.00 mL sample of 0.1001M sodium hydroxide required 22.37 mL of hydrochloric acid, HCl, for neutralization. What is the concentration of HCl?

Perform the indicated pH calculations.

1. The concentration of hydrogen ion is $3.6 \times 10^{-5}\text{M}$. Calculate the concentration of hydroxide ion, the pH, and the pOH.
2. The pH of a solution is 5.53. What is $[\text{H}^+]$ and $[\text{OH}^-]$?

3. A solution of nitric acid is prepared by diluting 10.00 mL of 0.02500M nitric acid to a final volume of 50.00 mL. What is the expected pH of the diluted solution?
4. 25.00 mL of 0.01534M hydrochloric acid is combined with 35.00 mL of 0.01135M sodium hydroxide. What is the pH of the final solution?
5. A buffer solution containing acetic acid ($\text{CH}_3\text{CO}_2\text{H}$) and sodium acetate ($\text{CH}_3\text{CO}_2\text{Na}$) is prepared using 10.69 grams of acetic acid and 11.73 grams of sodium acetate dissolved in enough water to produce 250.0 mL of solution. What is the expected pH of this buffer? The pK_a for acetic acid is 4.74.
6. A buffer made from benzoic acid (a monoprotic acid) and sodium benzoate contains 0.0250M sodium benzoate and 0.0150M benzoic acid. A 5.00 mL volume of 0.1334M sodium hydroxide is added to 50.00 mL of this buffer. By how much will the pH change, i.e. what is the difference between the original pH and the pH after adding the sodium hydroxide? The pK_a for benzoic acid is 4.20.
7. How much sodium formate (NaCO_2 , the salt of formic acid) must be added to 10.00 grams of formic acid (HCO_2H , monoprotic acid), to produce a buffer solution having a $\text{pH} = 4.00$? The pK_a for formic acid is 3.74.

Acid-base calculations answers:

1. $n_a M_a V_a = n_b M_b V_b$

$$3 \times 0.05534\text{M} \times V_a = 2 \times 0.01095\text{M} \times 25.00 \text{ mL}$$

$$V_a = 3.298 \text{ mL}$$

2. Formula weight of benzoic acid is 122.12 g/mole.

$$0.1993 \text{ g} / 122.12 \text{ g/mole} = 0.001632 \text{ moles benzoic acid.}$$

At neutralization, moles benzoic acid = moles sodium hydroxide.

$$0.001632 \text{ moles sodium hydroxide} / 0.01872 \text{ L} = 0.08718\text{M}$$

3. $n_a M_a V_a = n_b M_b V_b$

a) $1 \times 0.0878\text{M} \times 25.00 \text{ mL} = 1 \times 0.05534\text{M} \times V_b$

$$V_b = 39.7 \text{ mL}$$

b) $3 \times 0.0250\text{M} \times 7.50 \text{ mL} = 1 \times 0.05534\text{M} \times V_b$

$$V_b = 10.2 \text{ mL}$$

4. $n_a M_a V_a = n_b M_b V_b$

a) $1 \times 0.1003\text{M} \times V_a = 1 \times 0.09938\text{M} \times 25.00 \text{ mL}$

$$V_a = 24.82 \text{ mL}$$

b) $1 \times 0.1003\text{M} \times V_a = 2 \times 0.05535\text{M} \times 10.00 \text{ mL}$

$$V_a = 11.04 \text{ mL}$$

5. $n_a M_a V_a = n_b M_b V_b$

$$1 \times M_a \times 22.37 \text{ mL} = 1 \times 0.1001\text{M} \times 15.00 \text{ mL}$$

$$M_a = 0.06712\text{M}$$

pH calculation answers:

$$1. \quad K_w = 1.00 \times 10^{-14} = [H^+][OH^-]$$

$$1.00 \times 10^{-14} = [3.6 \times 10^{-5}][OH^-]$$

$$[OH^-] = 2.8 \times 10^{-10}$$

$$pH = -\log[H^+] = -\log[3.6 \times 10^{-5}] = 4.44$$

$$14.00 = pH + pOH$$

$$14.00 - 4.44 = pOH = 9.56$$

$$2. \quad [H^+] = 10^{-pH}$$

$$[H^+] = 10^{-5.53} = 3.0 \times 10^{-6}$$

$$[OH^-] = 1.00 \times 10^{-14} / 3.0 \times 10^{-6} = 3.3 \times 10^{-9}$$

$$3. \quad M_i V_i = M_f V_f$$

$$0.02500M \times 10.00 \text{ mL} = M_f \times 50.00 \text{ mL}$$

$$M_f = 0.0050000M = [H^+]$$

$$pH = -\log[0.0050000] = 2.30$$

$$4. \quad 0.02500 \text{ L} \times 0.01534M = 0.0003835 \text{ moles HCl}$$

$$0.03500 \text{ L} \times 0.01135M = 0.00039725 \text{ moles NaOH}$$

NaOH is in excess. Moles of NaOH after reacting with HCl is

$$0.00039725 \text{ moles} - 0.0003835 \text{ moles} = 0.00001375 \text{ moles NaOH}$$

$$0.00001375 \text{ moles} / 0.06000L = 0.0002292M \text{ NaOH} = [OH^-]$$

$$pOH = -\log[0.0002292] = 3.64$$

$$pH = 14.00 - 3.64 = 10.36$$

$$5. \quad MW \text{ CH}_3\text{CO}_2\text{H} = 60.05 \text{ g/mole}$$

$$FW \text{ CH}_3\text{CO}_2\text{Na} = 82.03 \text{ g/mole}$$

$$[\text{CH}_3\text{CO}_2\text{H}] = (10.69\text{g} / 60.05 \text{ g/mole}) \div 0.2500\text{L} = 0.7121\text{M}$$

$$[\text{CH}_3\text{CO}_2\text{Na}] = (11.73\text{g} / 82.03 \text{ g/mole}) \div 0.2500\text{L} = 0.5720\text{M}$$

$$\text{pH} = \text{pK}_a + \log(0.5720\text{M}/0.7121\text{M})$$

$$\text{pH} = 4.74 + (-0.095) = 4.64$$

6. Starting pH of buffer:

$$\text{pH} = 4.20 + \log(0.0250\text{M}/0.0150\text{M})$$

$$\text{pH} = 4.20 + (0.22) = 4.42$$

$$\text{Initial moles of salt} = 0.05000\text{L} \times 0.0250\text{M} = 0.00125 \text{ moles}$$

$$\text{Initial moles of acid} = 0.05000\text{L} \times 0.0150\text{M} = 0.000750 \text{ moles}$$

$$\text{Moles NaOH added} = 0.00500 \text{ L} \times 0.1334\text{M} = 0.000667 \text{ moles}$$

$$\text{Final moles of salt} = 0.00125 + 0.000667 = 0.001917 \text{ moles}$$

$$\text{Final moles of acid} = 0.000750 - 0.000667 = 0.0000830 \text{ moles}$$

$$[\text{salt}] = 0.001917 \text{ moles} / 0.05500\text{L} = 0.0349\text{M}$$

$$[\text{acid}] = 0.0000830 \text{ moles} / 0.05500\text{L} = 0.00151\text{M}$$

$$\text{pH} = 4.20 + \log(0.0349\text{M}/0.00151\text{M})$$

$$\text{pH} = 4.20 + (1.36) = 5.56$$

$$\text{change in pH} = 5.56 - 4.42 = 1.14 \text{ pH units.}$$

7. F.W. $\text{HCO}_2\text{H} = 30.03 \text{ g/mole}$

$$\text{F.W. NaCO}_2 = 52.01 \text{ g/mole}$$

$$10.00 \text{ g HCO}_2\text{H} = 0.3330 \text{ moles}$$

$$4.00 = 3.74 + \log ([\text{A}^-]/[0.3330])$$

$$10^{0.26} = [\text{A}^-]/[0.3330]$$

$$1.182 = [\text{A}^-]/[0.3330]$$

0.606 moles = [A⁻] = sodium formate

0.606 moles x 52.01 g/mole = 31.52 g sodium formate.