

UNIT (6) ACIDS AND BASES

6.1 | Arrhenius Definition of Acids and Bases

Definitions for acids and bases were proposed by the Swedish chemist *Savante Arrhenius* in 1884.

Acids were defined as compounds that contain hydrogen and can dissolve in water to release hydrogen ions into solution. For example, HCl dissolves in water as follows:



Bases are defined as substances that dissolve in water to release hydroxide ions (OH⁻) into solution. For example, a typical base according to the Arrhenius definition is sodium hydroxide, NaOH:



The process that converts a molecule such as HCl into ions is called ionization.

Ionization is the production of ions from molecular compounds. No ions were initially present.

The process of converting a solid ionic compound, such as NaOH, into ions in solution is called dissociation.

Dissociation is the production of ions from ionic compounds. Ions were initially present, water just breaks apart ions.

Some common acidic and basic substances:

Acidic	Basic
citrus fruits	soap
vinegar	detergents
aspirin	milk of magnesia

Some characteristic properties associated with acids and bases:

Acids	Bases
sour taste	bitter taste
change litmus paper to red	change litmus paper to blue
react with hydroxide bases to produce salt	slippery, soapy feeling

6.2 | Naming Acids

Most acids are **oxyacids**; that is, they contain oxygen in addition to hydrogen and another element. The names of these acids come from the root name of the polyatomic ion.

Suffixes are used based on the ending of the original name of the polyatomic anion. If the name of the polyatomic anion ends with *-ate*, change it to *-ic* for the acid and if it ends with *-ite*, change it to *-ous* in the acid.

HNO_3 , which contains the nitrate (NO_3^-), is called nitric acid.

HNO_2 , which contains the nitrite (NO_2^-), is called nitrous acid.

In addition to the oxyacids, there are a small number of acids that don't contain oxygen, ie: HCl . To name these acids, the prefix *hydro-* is used before the name of the nonmetal ion, and its *ide* ending is changed to *-ic acid*.

HCl , the nonmetal ion is **chloride**, the *ide* is changed to **-ic acid**, so the name is hydrochloric acid.

Worked Example 6-1

Name the following acids:



Solution



The anion name is *-fluoride-* so the name is *hydrofluoric acid*.



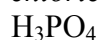
The oxyanion is carbonate, which ends in *-ate*; therefore the name of the acid is *carbonic acid*.



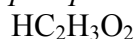
The anion name is *-iodide-* so the name is *hydroiodic acid*.



The oxyanion is chlorate, which ends in *-ate*; therefore the name of the acid is *chloric acid*.



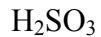
The oxyanion is phosphate, which ends in *-ate*; therefore the name of the acid is *phosphoric acid*.



The oxyanion is acetate, which ends in *-ate*; therefore the name of the acid is *acetic acid*.

Practice 6-1

Name the following acids:



Answer

6.3 | The pH Scale

The **pH** of a solution is a measure of the solution's acidity. Mathematically, pH is defined in terms of the negative logarithm of the H^+ concentration.

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

Square brackets around H^+ indicate the concentration in moles per liter (M).

pH less than 7.00 is an acidic solution

pH = 7.00 is a neutral solution

pH greater than 7.00 is a basic solution

Calculating pH From $[H^+]$

Case (I)

Calculating the pH values for which the coefficient in the exponential expression is 1.0.

This is an easy case and you **don't need a calculator**.

The pH is the negative of the exponent value of the power of 10.

$$[H^+] = 1.0 \times 10^{-4} \leftarrow pH = \text{this number without sign} \quad (pH = 4.00)$$

If this number is exactly

	[H ⁺]	pH	
	1.0	0	
	1.0 x 10 ⁻¹	1.00	
	1.0 x 10 ⁻²	2.00	
	1.0 x 10 ⁻³	3.00	
	1.0 x 10 ⁻⁴	4.00	
	1.0 x 10 ⁻⁵	5.00	
↑	Acidic	1.0 x 10 ⁻⁶	6.00
	Neutral	1.0 x 10 ⁻⁷	7.00
	Basic	1.0 x 10 ⁻⁸	8.00
		1.0 x 10 ⁻⁹	9.00
		1.0 x 10 ⁻¹⁰	10.00
		1.0 x 10 ⁻¹¹	11.00
		1.0 x 10 ⁻¹²	12.00
		1.0 x 10 ⁻¹³	13.00
↓		1.0 x 10 ⁻¹⁴	14.00

The Rule for Significant Figures in Logarithm Calculations

The rules for logs are different from other mathematical operations. The number of decimal places for a log must be equal to the number of significant figures in the original number.

Number of significant figures in [H⁺] = number of digits after decimal places in pH.

Case (II)

If the coefficient in the exponential expression is not 1.0, you **need a calculator**.

Worked Example 6-2

A solution has a [H⁺] = 3.7 x 10⁻⁵ M. What is the pH?

Solution

Steps for calculating pH on a TYPICAL calculator.

	Operation	Display
1.	Press 3.7	3.7
2.	Press EE or EXP	3.7 00
3.	Press 5	3.7 05
4.	Press +/- (change-of-sign)	3.7 -05
5.	Press LOG key	-4.431798276
6.	Press +/- (change-of-sign)	4.431798276

After rounding, the reported value is 4.43.

Two significant figures in H⁺ two digits to the right of the decimal point.

Practice 6-2

Calculate the pH of each of the following solutions and classify the solution as acidic or basic:

Lemon juice: $[H^+] = 1.1 \times 10^{-2} \text{ M}$

Urine: $[H^+] = 1.2 \times 10^{-6} \text{ M}$

Antiseptic: $[H^+] = 1.2 \times 10^{-10} \text{ M}$

Coffee: $[H^+] = 7.9 \times 10^{-6} \text{ M}$

Soft drink: $[H^+] = 2.5 \times 10^{-4} \text{ M}$

Blood: $[H^+] = 3.9 \times 10^{-8} \text{ M}$

Beer: $[H^+] = 5.0 \times 10^{-3} \text{ M}$

Detergent: $[H^+] = 3.2 \times 10^{-11} \text{ M}$

Answer

6.4 | Brønsted-Lowry Acids and Bases

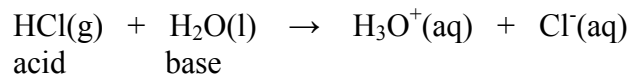
A broader definition of an acid and base was introduced (independently) in 1923 by the Danish chemist **Johannes Brønsted** and the English chemist **Thomas Lowry**. The definition is called Brønsted-Lowry Acids and Bases, honoring the names of its founders. The Brønsted-Lowry definition of an acid is very similar to the Arrhenius definition. Any substance that can donate a hydrogen ion is an acid according to Brønsted-Lowry. Acids are often referred to as **proton** donors because an H^+ ion (hydrogen without its electron) is simply a proton.

The Brønsted-Lowry definition of a **base** is quite different from the Arrhenius definition. The Brønsted-Lowry base is defined as any substance that can accept a hydrogen ion. In essence, a base is the opposite of an **acid**.

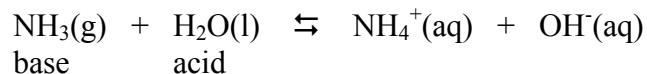
Acid: An acid is a proton (H^+ ion) donor

Base: A base is a proton (H^+ ion) acceptor

According to the Brønsted-Lowry definition, HCl is an acid because it donates H^+ to a water molecule.



According to the Brønsted-Lowry definition, NH_3 is a base because it accepts H^+ from the water molecule.



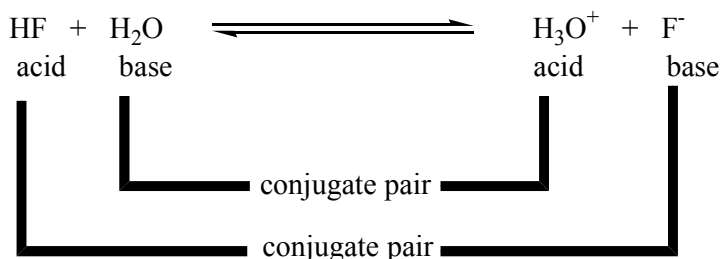
6.5 | Conjugate Acid-Base Pairs

Most Brønsted-Lowry acid-base reactions are reversible reactions in equilibrium.

Consider the reaction: $\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{F}^- + \text{H}_3\text{O}^+$

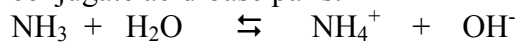
HF is the acid in the forward reaction because it donates H^+ to become F^- , while H_2O is the base because it accepts H^+ to become H_3O^+ . For the reverse reaction, F^- is the base because it accepts H^+ to become HF, and H_3O^+ is the acid because it donates H^+ to become H_2O .

The products of an acid-base reaction are themselves acids and bases. HF and F^- or H_2O and H_3O^+ are called conjugate acid-base pairs because their formulas differ by **one H^+** .



Worked Example 6-3

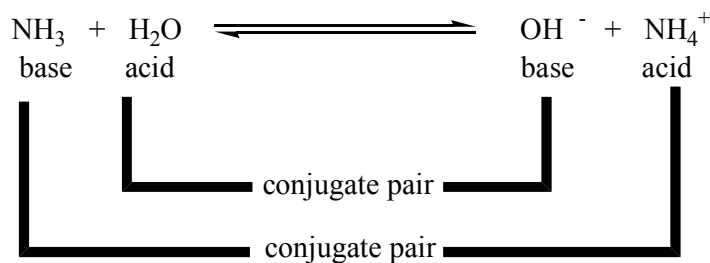
For the following reaction, label each substance as an acid or a base and show the conjugate acid-base pairs.



Solution

In the forward reaction, NH_3 is the base (it accepts H^+ to become NH_4^+) and H_2O is the acid (it donates H^+ to become OH^-). For the reverse reaction, NH_4^+ is the acid (it donates H^+ to become NH_3) and OH^- is the base (it accepts H^+ to become H_2O).

NH_3 and NH_4^+ or H_2O and OH^- differ by one H^+ and therefore conjugate acid-base pairs.



Practice 6-3

For the following reaction label each substance as an acid or a base and show the conjugate acid-base pairs.

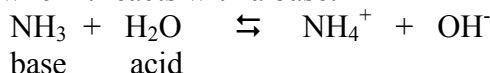


Answer

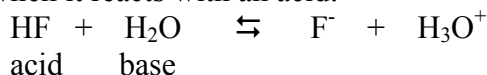
6.6 | Water: an Acid and a Base

Compounds or ions that can act as an acid or a base are called **amphoteric** substances. Water is the most common example of an amphoteric substance; others include hydrogen carbonate (HCO_3^-) and the amino acids (building blocks of proteins).

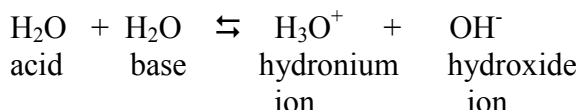
Water acts as an acid when it reacts with a base:



Water acts as a base when it reacts with an acid:

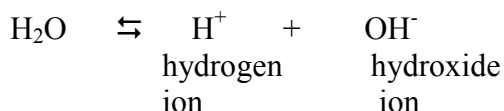


Water can act as both an acid and a base in the same reaction:



The above reaction illustrates the **self-ionization of water**. It does not proceed to the right to a large extent because water is weak acid and base.

A more simplified version of the self-ionization reaction is:



There is a common practice of substituting H^+ for H_3O^+ to simplify the equation.

Whenever you see the symbol H^+ (hydrogen ion), you should realize that we are actually referring to H_3O^+ (hydronium ion).

In pure water H^+ and OH^- form in equal amounts, therefore their concentrations are equal. The concentrations of H^+ and OH^- at 25°C are each $1.0 \times 10^{-7} \text{ mol/L (M)}$.

In pure water:

$$\begin{aligned} [\text{H}^+] &= 1.0 \times 10^{-7} \text{ M} \\ [\text{OH}^-] &= 1.0 \times 10^{-7} \text{ M} \end{aligned}$$

The square brackets, [], are used to indicate concentration in moles per liter (M).

Multiplying $[\text{H}^+]$ and $[\text{OH}^-]$ concentrations gives a constant called **ion product of water**, symbol K_w .

$$K_w = [\text{H}^+] \times [\text{OH}^-] = (1.0 \times 10^{-7}) \times (1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

$K_w = [\text{H}^+] [\text{OH}^-]$ applies, not only to pure water, but to any aqueous solution.

Notice that $[\text{H}^+]$ and $[\text{OH}^-]$ are inversely proportional. This means that if $[\text{H}^+]$ goes up, the $[\text{OH}^-]$ must come down so that the product of the two is still 1.0×10^{-14} .

Acidic solution: $[\text{H}^+] > 10^{-7} \text{ M}$ and $[\text{OH}^-] < 10^{-7} \text{ M}$
Neutral solution: $[\text{H}^+] = 10^{-7} \text{ M}$ and $[\text{OH}^-] = 10^{-7} \text{ M}$
Basic solution: $[\text{H}^+] < 10^{-7} \text{ M}$ and $[\text{OH}^-] > 10^{-7} \text{ M}$

Worked Example 6-4

Calculate the concentration of OH^- in an aqueous solution if the H^+ concentration is 2.8×10^{-3} .

Solution

We know that $K_w = [\text{H}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14}$
The $[\text{H}^+]$ is given and we solve for $[\text{OH}^-]$.

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{2.8 \times 10^{-3}} = 3.6 \times 10^{-12} \text{ M}$$

Practice 6-4

A) Calculate the $[\text{H}^+]$ given the $[\text{OH}^-]$ in each of the following aqueous solutions and classify each solution as acidic or basic.

$$[\text{OH}^-] = 6.2 \times 10^{-3} \text{ M}$$

$$[\text{OH}^-] = 1.3 \times 10^{-11} \text{ M}$$

B) Calculate the $[\text{OH}^-]$ given the $[\text{H}^+]$ in each of the following aqueous solutions and classify each solution as acidic or basic.

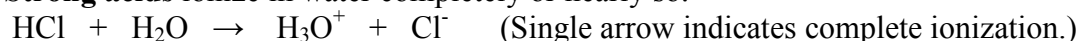
$$[\text{H}^+] = 7.4 \times 10^{-5} \text{ M}$$

$$[\text{H}^+] = 2.3 \times 10^{-9} \text{ M}$$

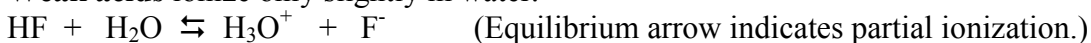
Answer

6.7 | Strengths of Acids and Bases

Strong acids ionize in water completely or nearly so.



Weak acids ionize only slightly in water.



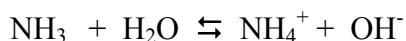
Common strong and weak acids:

Strong Acids		Weak Acids	
HCl	hydrochloric acid	HF	hydrofluoric acid
HBr	hydrobromic acid	HC ₂ H ₃ O ₂	acetic acid
HI	hydroiodic acid	HCHO ₂	formic acid
HNO ₃	nitric acid	H ₂ SO ₃	sulfurous acid
HClO ₄	perchloric acid	H ₂ CO ₃	carbonic acid
H ₂ SO ₄	sulfuric acid	H ₃ PO ₄	phosphoric acid

Strong bases dissociate in water completely or nearly so.



Weak bases ionize in water only slightly.



Common strong and weak bases:

Strong Bases		Weak Bases	
NaOH	sodium hydroxide	NH ₃	ammonia
KOH	potassium hydroxide	CH ₃ NH ₂	methylamine
Ca(OH) ₂	calcium hydroxide	C ₂ H ₅ NH ₂	ethylamine

Worked Example 6-5

For a 6.8×10^{-4} M KOH solution, determine the following:

a) $[\text{OH}^-]$ b) $[\text{H}^+]$, c) the pH

Solution

a) Because KOH is a strong base, it completely dissociates:



So the concentration of hydroxide ion of a 6.8×10^{-4} M KOH solution is 6.8×10^{-4} M.

$$[\text{OH}^-] = 6.8 \times 10^{-4}$$

b) Substituting the $[\text{OH}^-]$ value in $K_w = [\text{H}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14}$ yields:

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{6.8 \times 10^{-4}} = 1.5 \times 10^{-11} \text{ M}$$

c) $\text{pH} = -\log [\text{H}^+]$ $\text{pH} = -\log(1.5 \times 10^{-11})$ $\text{pH} = 10.82$

Practice 6-5

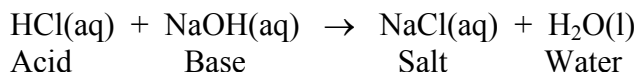
For a 7.4×10^{-3} M HCl solution, determine the following:

a) $[\text{H}^+]$ b) $[\text{OH}^-]$ c) The pH

Answer

6.8 | Reactions of Acids and Bases “Neutralization Reactions”

A **neutralization** reaction is when an acid reacts with a base to form a *salt* and water.



This reaction is a double-replacement reaction where the H^+ from the acid reacts with the OH^- from the base to produce the molecular compound water (H-OH or H_2O).

Salt is an ionic compound formed by combination of a cation and an anion.

The salt contains the cation from the base and the anion from the acid.

For any neutralization reaction we can write the **molecular equation**, **ionic equation**, and **net ionic equation**. See unit (4).

Worked Example 6-6

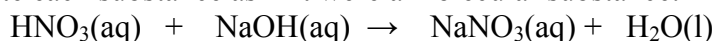
Write molecular, ionic, and net ionic equation for the reaction of HNO_3 and NaOH .

Solution

Molecular equation:

Neutralization is a double replacement reaction. See *unit 4*.

We write each substance as if it were a molecular substance:



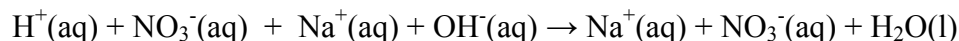
Ionic equation:

The formula $\text{HNO}_3(\text{aq})$ actually means $\text{H}^+(\text{aq})$ and $\text{NO}_3^-(\text{aq})$.

The formula $\text{NaOH}(\text{aq})$ actually means $\text{Na}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$.

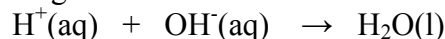
The formula $\text{NaNO}_3(\text{aq})$ actually means $\text{Na}^+(\text{aq})$ and $\text{NO}_3^-(\text{aq})$.

We can replace the molecular equation with the following ionic equation:



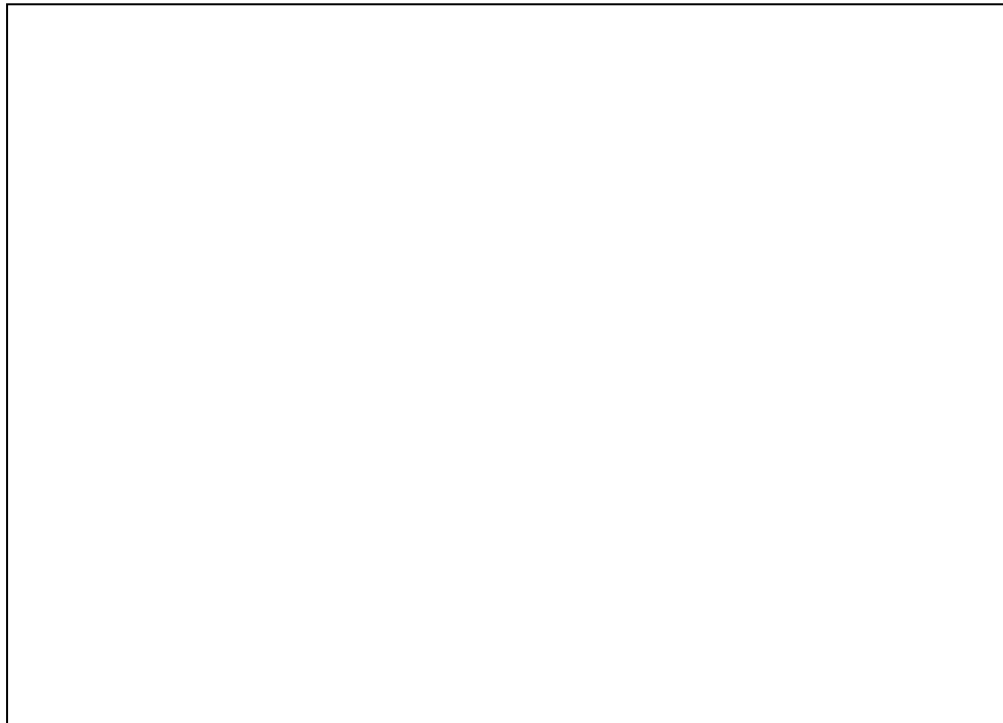
Net ionic equation:

Removing spectator ions (ions that have not changed) from the ionic equation gives the following:



Practice 6-6

Write molecular, ionic, and net ionic equation for the reaction of HCl and KOH.

Answer

6.9 | Acid-Base Titrations

The theory and principles of acid-base neutralization are combined with stoichiometry (*unit 4*) in a laboratory procedure called a **titration**. A titration can be used to determine the concentration of an unknown acid or base. The core concept is quite simple, if the concentration of one (acid, or base) is known, the concentration of the other can be determined.

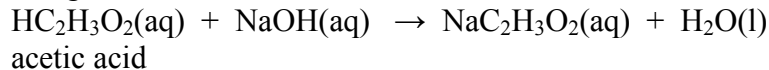
For example, consider an acid-base titration involving acetic acid (in vinegar) and NaOH (base). The concentration of the base is known and our task is to determine the concentration of the acid.

In the titration, the solution of unknown concentration (in this case vinegar) is placed in a flask. The solution of known concentration (in this case NaOH) is placed in a buret. The NaOH (from the buret) is carefully added to the vinegar (in the flask). The addition continues in small portions until an observed “color change” signals that the titration is complete. This is called the *end point* or *equivalence point* of the titration. The end point (color change) is signaled by an *indicator*, a dye that changes color depending on the acidity or basicity of the solution.

When the equivalence point is reached, the volume of the base is determined from the buret. We can calculate the number of moles of reacted base from volume and concentration of the base. Next, we use stoichiometry to calculate the moles of the unknown acid which is required to calculate the molarity of the acid.

Worked Example 6-7

A 5.00-mL sample of vinegar requires 7.20 mL of a 0.584 M solution of NaOH for titration to the equivalence point. A) Calculate the molarity of the acetic acid in the vinegar solution. B) Calculate the percent (m/v) of the acetic acid in the vinegar solution.

**Solution**

Step 1 The amount of base reacted (the moles of NaOH).

The volume of the base (in L) multiplied by the molarity of the base, gives the number of moles of the base.

$$0.00720 \text{ L NaOH} \times \frac{0.584 \text{ mol NaOH}}{1 \text{ L NaOH}} = 0.00420 \text{ mol NaOH}$$

Step 2 Since the stoichiometry factor is 1 (see balanced equation above - 1 mol $\text{HC}_2\text{H}_3\text{O}_2$ requires 1 mol NaOH), therefore the number of moles of the acid equals the number of moles of the base.

$$\text{Moles acid} = 0.00420 \text{ mol}$$

Step 3 Molarity of the acid is moles of acid (0.00420) divided by the volume of acid in L (0.00500 L).

$$\text{Molarity} = \frac{0.00420 \text{ mol}}{0.00500 \text{ L}} = 0.840 \text{ M}$$

Grams of acid:

The molar mass (from periodic table) of $\text{HC}_2\text{H}_3\text{O}_2$ is 60.06 g/mol.

$$\text{Grams acid} = 0.00420 \text{ mol acid} \times \frac{60.06 \text{ g acid}}{1 \text{ mol acid}} = 0.252 \text{ g acid}$$

$$\text{percent (m/v)} = \frac{\text{mass of acid}}{\text{mL of solution}} \times 100\%$$

$$\text{percent (m/v)} = \frac{0.252 \text{ g}}{5.00 \text{ mL}} \times 100\% = 5.05\%$$

6.10 | Buffer Solutions

The pH of water and most solutions changes drastically when you add a small amount of acid or bases to them.

A **buffer** is a solution that resists changes in pH when small amounts of acid or base are added.

Buffers contain two chemicals:

- 1) A substance to react with and remove the ADDED acid.
- 2) A substance to react with and remove the ADDED base.

Generally a buffer is a solution containing nearly equal concentrations of a weak acid and the salt of its conjugate base (recall Brønsted-Lowry acid-base).

The following table lists some weak acids and their conjugate bases.

Weak Acid	Conjugate Base
Acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$)	Acetate ion ($\text{C}_2\text{H}_3\text{O}_2^-$)
Carbonic acid (H_2CO_3)	Hydrogen carbonate ion (HCO_3^-)
Ammonium ion (NH_4^+)	Ammonia (NH_3)

How Does a Buffer Work?

A common buffer solution may be prepared from acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) and sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$).

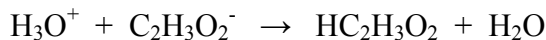
Remember that $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid so in solution it exists primarily in its molecular form $\text{HC}_2\text{H}_3\text{O}_2$. The $\text{NaC}_2\text{H}_3\text{O}_2$ however, is an ionic substance that will ionize in solution to the $\text{C}_2\text{H}_3\text{O}_2^-$ and Na^+ ions.

A) The Response of a $\text{HC}_2\text{H}_3\text{O}_2/\text{C}_2\text{H}_3\text{O}_2^-$ Buffer to the Addition of H_3O^+ Ions

What happens if a small amount of a **strong acid** such as HCl is added to the buffer solution?

The buffer must remove most of the added H_3O^+ ions from the strong acid, otherwise the pH would drop dramatically.

The H_3O^+ ions combine with the acetate ions and are removed according to the equation below:

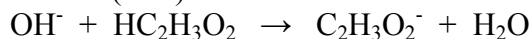


Since most of the new H_3O^+ ions are removed, the pH remains relatively constant.

B) The Response of a $\text{HC}_2\text{H}_3\text{O}_2/\text{C}_2\text{H}_3\text{O}_2^-$ Buffer to the Addition of OH^- Ions

What happens if a small amount of a **strong base** such as NaOH is added to the buffer solution?

The newly added base (OH^-) reacts with the acetic acid and is removed from solution.



Again, since most of the new OH^- ions are removed, the pH remains relatively constant.

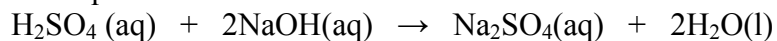
Homework Problems

- 6.1 Write the formula for each of the following:
- lithium hydroxide
 - iron(III) hydroxide
 - aluminum hydroxide
 - chlorous acid
 - strontium hydroxide
 - lead(II) hydroxide
 - phosphorous acid
 - ammonia
- 6.2 Calculate the pH of the following solutions:
- $[\text{H}_3\text{O}^+] = 5.6 \times 10^{-3}$
 - $[\text{H}_3\text{O}^+] = 3.8 \times 10^{-4}$
 - $[\text{H}_3\text{O}^+] = 2.7 \times 10^{-5}$
 - $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-9}$
- 6.3 For each of the following strong base solutions, determine $[\text{OH}^-]$, $[\text{H}_3\text{O}^+]$, and pH.
- 6.5×10^{-1} M NaOH
 - 4.7×10^{-2} M KOH
 - 1.2×10^{-2} M LiOH
 - 8.8×10^{-3} M NaOH
- 6.4 Write equations for the reaction of each of the following Brønsted-Lowry acids and bases. Identify the conjugated acids and bases.
- Acid: H_2O ; base: NH_3
 - Acid: NH_4^+ ; base: OH^-
 - Acid: HSO_4^- ; base: H_2O
 - Acid: HCl ; base: H_2PO_4^-

- 6.5 Four solutions of unknown H_2SO_4 concentration are titrated with solutions of NaOH . The following table lists the initial volume of each unknown H_2SO_4 solution titrated, the volume of NaOH solution required to reach end point, and the molarity of each NaOH solution used.

Calculate the molarity of the acid in each case.

The equation for the reaction is



	H_2SO_4 volume	NaOH volume	$[\text{NaOH}]$
a)	10.00 mL	16.50 mL	0.1820 M
b)	15.00 mL	10.85 mL	0.7550 M
c)	20.00 mL	24.10 mL	0.5923 M
d)	25.00 mL	12.77 mL	0.9505 M

- 6.6 Write an equation for each of the following buffer actions:
- The response of HClO/ClO^- buffer to the addition of H_3O^+ ions.
 - The response of HClO/ClO^- buffer to the addition of OH^- ions.
 - The response of $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer to the addition of H_3O^+ ions.
 - The response of $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ buffer to the addition of OH^- ions.