

Acids and Bases

Learning Objectives

- I. Acids and Bases Character and Concepts**
- II. Naming Acids and Bases**
- III. Conjugate Acid – Base Pairs**
- IV. Strengths of Acids and Bases**
- V. Dissociation of Weak Acids and Bases**
- VI. Dissociation of Water**
- VII. pH**
- VIII. Reactions of Acids and Bases**
- IX. Titration and Stoichiometry of Acid Base Reactions**
- X. Buffers**
- XI. Salts Hydrolysis**

I. Acids and Bases Character and Concepts

Acids are characterized by their sour taste and stingy character and they tend to turn blue litmus paper red. Bases on the other hand are characterized by their bitter taste and slippery and soapy character and they tend to turn red litmus paper blue.

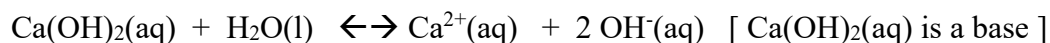
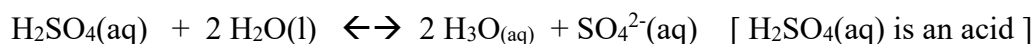
There are three acids and bases concepts:

1. Arrhenius Concept of Acids and Bases:

Arrhenius concept of acids and bases states that the acids any chemical compounds that form hydrogen ions in aqueous solution and that the bases any chemical compounds that form hydroxide ions in aqueous solution.

In other words, acids are chemicals that can release hydrogen ions into a solution (mainly water) and bases are chemical compounds that can release hydroxide ions into a solution (mainly water).

Examples:



Arrhenius concept of acids and bases was a great contribution and it did replace Lavoisier's old concept of the acids that acids are compounds that have oxygen.

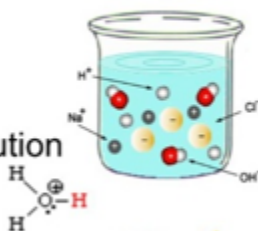
On the other hand, Arrhenius concept has several disadvantages:

- It did not take any consideration of solvent effect and role.
- It did not cover the concept of acids and bases in other solvent than water such as toluene which is not dissociating (no ionization).

The figures below summarized the Arrhenius Acids - Bases Concepts:

Arrhenius Acids Concept

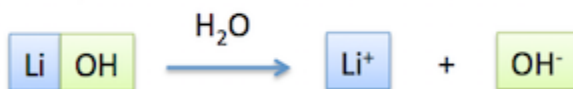
compound that **contains hydrogen** and ionizes in solution to form **hydrogen ions (H^+)**



Arrhenius Bases Concept



Potassium Hydroxide



Lithium Hydroxide

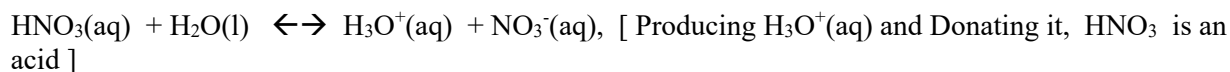
2. Bronsted – Lowry Concept of Acids and Bases:

Bronsted – Lowry concept of acid and bases replaces Arrhenius concept of acids and bases. It introduces the concept of Donor – Acceptor character of the acids and bases.

An acid according to Bronsted – Lowry is a chemical compound that produces hydronium ion when dissociating in water and furthermore they can donate this hydronium ion to another chemical compound which can accept this donated hydronium ion.

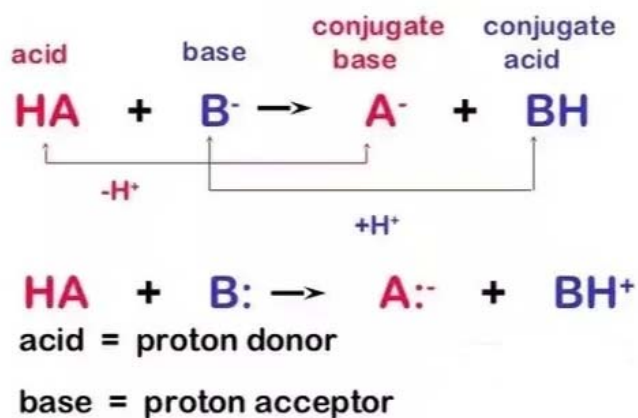
The acid is said to be hydronium ion donor and the base is said to be hydronium ion acceptor.

Example:



The figure below summarized the Bronsted – Lowry Acids - Bases Concepts:

BRONSTED - LOWRY THEORY



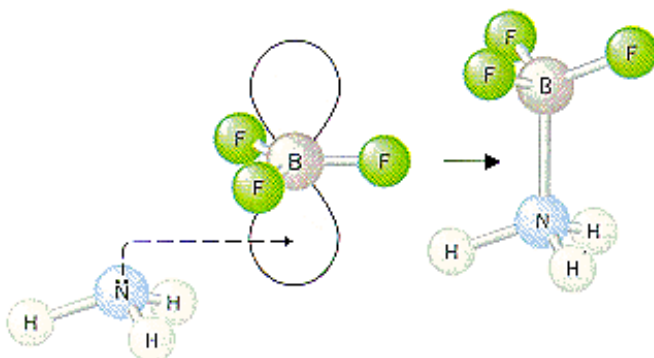
3. Lewis Concept of Acids and Bases:

Lewis concept of acids and bases is based on the **Electron Pair (Lone Pair Electron) Donor – Acceptor**. Bases have excess of lone pairs can donate their lone pair electrons and the bases are said to be **Electron Pair Donor**. The acids can accept these **Lone Pair Electrons** and said to be **Electron Pair Acceptor**.

The two examples taken from the reference below illustrate this **Electron Pair Donor – Acceptor** features of Acids and Bases.

<http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch11/lewis.php>

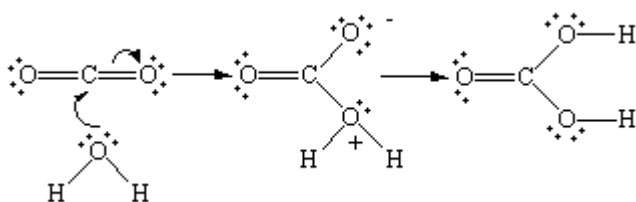
The first example:



The lone pair electrons on nitrogen atom of ammonia NH_3 are donated to vacant site of Boron in BF_3

NH_3 is said to be Lewis Base (Electron Pair Donor) and BF_3 is to be Lewis Acid (Electron Pair Acceptor)

The second example:



The lone pair electrons on oxygen atom are donated to vacant site of Carbon in CO_2

H_2O is said to be Lewis Base (Electron Pair Donor) and CO_2 is to be Lewis Acid (Electron Pair Acceptor)

II. Naming Acids and Bases

Naming Acids

Naming acids is divided in two categories:

a. Acids without Oxygen

Acids without oxygen such as HF , HCl , HBr , HI , HCN and H_2S :

Naming such acids without oxygen will follow the procedure below:

Start with the hydrogen and name it as Hydro

End with the second element and convert it with the ending “ic acid” ending.

Examples:

Acid Formula	Acid Name
HF	Hydrofluoric Acid
HCl	Hydrochloric Acid
HBr	Hydrobromic Acid
HI	Hydroiodic Acid
HCN	Hydrocyanic acid

Exception for this is H₂S acid which is named as **Hydrogen Sulfide**.

b. Acids with Oxygen (also known as Oxoacids):

Acids with oxygen such as HNO₃, H₂SO₄, H₃PO₄, HClO₃, and H₂CO₃

Naming such acids with oxygen will follow the procedure below:

Start with the middle nonmetal atom (not oxygen or hydrogen) and convert it and name it.

End with “ic acid”

Examples:

Acid Formula	Acid Name
HNO ₃	Nitric Acid
H ₂ SO ₄	Sulfuric Acid
H ₃ PO ₄	Phosphoric Acid
HClO ₃	Chloric Acid
H ₂ CO ₃	Carbonic Acid

Some oxoacids (acids contain oxygen) have different amount of oxygen atoms. In order to name such oxoacids, one will need to use a reference or standard point. Anions that end with “ate” will form the “the ic acid” and will be used as a reference acid point. Oxoacids with higher amount of oxygen than the standard acid point will have the prefix “per” is added. Less than this standard point acid will have the ending “ous acid” and more less oxygen (least amount of oxygen) will have the prefix of “hypo” and will end with “ous acid”. The table below shows several acids with this trend.

Acid Formula	Acid Name	Standard Point Acid
HClO₄	Perchloric Acid	
HClO₃	Chloric Acid	HClO₃
HClO₂	Chlorous Acid	
HClO	Hypochlorous Acid	

Acid Formula	Acid Name	Standard Point Acid
H₂SO₅	Persulfuric Acid	

H₂SO₄	Sulfuric Acid	H₂SO₄
H₂SO₃	Sulfurous Acid	
H₂SO₂	Hyposulfurous Acid	

Acid Formula	Acid Name	Standard Point Acid
HNO₄	Persulfuric Acid	
HNO₃	Sulfuric Acid	HNO₃
HNO₂	Sulfurous Acid	
HNO	Hyposulfurous Acid	

Acid Formula	Acid Name	Standard Point Acid
H₃PO₅	Perphosphoric Acid (also known as Peroxyphosphoric acid)	
H₃PO₄	Phosphoric Acid	H₃PO₄
H₃PO₃	Phosphorous Acid	
H₃PO₂	Hypophosphorous Acid	

The table below illustrates some acids with their corresponding anions:

<http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch2/names.html#acids>

Acids containing ions ending with <i>ide</i> often become				<i>hydro -ic acid</i>
Cl ⁻	chloride		HCl	hydrochloric acid
F ⁻	fluoride		HF	hydrofluoric acid
S ²⁻	sulfide		H ₂ S	hydrosulfuric acid
Acids containing ions ending with <i>ate</i> usually become				<i>-ic acid</i>
CH ₃ CO ₂ ⁻	acetate		CH ₃ CO ₂ H	acetic acid
CO ₃ ²⁻	carbonate		H ₂ CO ₃	carbonic acid
BO ₃ ³⁻	borate		H ₃ BO ₃	boric acid
NO ₃ ⁻	nitrate		HNO ₃	nitric acid
SO ₄ ²⁻	sulfate		H ₂ SO ₄	sulfuric acid
ClO ₄ ⁻	perchlorate		HClO ₄	perchloric acid

PO_4^{3-}	phosphate		H_3PO_4	phosphoric acid
MnO_4^-	permanganate		HMnO_4	permanganic acid
CrO_4^{2-}	chromate		H_2CrO_4	chromic acid
ClO_3^-	chlorate		HClO_3	chloric acid
Acids containing ions ending with <i>ite</i> usually become				<i>-ous acid</i>
ClO_2^-	chlorite		HClO_2	chlorous acid
NO_2^-	nitrite		HNO_2	nitrous acid
SO_3^{2-}	sulfite		H_2SO_3	sulfurous acid
ClO^-	hypochlorite		HClO	hypochlorous acid

Naming Bases

Naming bases is divided in three categories:

- Bases are made of metals that have no multiple charges. [No Roman Numbers are used]. Groups 1 A, 2A and part of group 3 A metals of the periodic table have all metals with multiple charges cations. For example group 1 A (oxidation number +1), group 2 A (oxidation number + 2) and group 3 A (oxidation number + 3).

Example:

Base Formula	Base Name	Group in Periodic Table
NaOH	Sodium Hydroxide	1 A
FrOH	Francium Hydroxide	1 A
Ca(OH)_2	Calcium Hydroxide	2 A
Ba(OH)_2	Barium Hydroxide	2 A
Al(OH)_3	Aluminum Hydroxide	3 A

- Bases are made of metals that have multiple charges. [Roman Numbers are used].

Some transition metals of group B in periodic table have multiple charges. The name of the bases of these transition metals will have the roman numbers to distinguish among the different charges of the same metal.

The table below shows these transition metals with their multiple charges.

https://saylordotorg.github.io/text_general-chemistry-principles-patterns-and-applications-v1.0/s06-03-naming-ionic-compounds.html

Cation	Systematic Name	Common Name	Cation	Systematic Name	Common Name
Cr ²⁺	chromium(II)	chromous	Cu ²⁺	copper(II)	cupric
Cr ³⁺	chromium(III)	chromic	Cu ⁺	copper(I)	cuprous
Mn ²⁺	manganese(II)	manganous*	Hg ²⁺	mercury(II)	mercuric
Mn ³⁺	manganese(III)	manganic*	Hg ₂ ²⁺	mercury(I)	mercurous [†]
Fe ²⁺	iron(II)	ferrous	Sn ⁴⁺	tin(IV)	stannic
Fe ³⁺	iron(III)	ferric	Sn ²⁺	tin(II)	stannous
Co ²⁺	cobalt(II)	cobaltous*	Pb ⁴⁺	lead(IV)	plumbic*
Co ³⁺	cobalt(III)	cobaltic*	Pb ²⁺	lead(II)	plumbous*
* Not widely used.					
[†]The isolated mercury(I) ion exists only as the gaseous ion.					

The periodic table below shows all charges for all elements of periodic table

[http://www.cabrillo.edu/~aromero/Common%20Files/Periodic%20Table%20\(Common%20Ionic%20Charges\).pdf](http://www.cabrillo.edu/~aromero/Common%20Files/Periodic%20Table%20(Common%20Ionic%20Charges).pdf)

Example:

Base Formula	Base Name	Group in Periodic Table
Pb(OH) ₄	Lead- IV -Hydroxide	4 A
Cr(OH) ₃	Chromium- III -Hydroxide	6 B
Fe(OH) ₂	Iron- II -Hydroxide	8 B
V(OH) ₅	Vanadium- V -Hydroxide	5 B
CuOH	Copper- I -Hydroxide	1 B

c. Bases are made of nonmetals only. No metals are part of their formulas.

Such bases have covalent bonds and not ionic bonds. Example of such bases are Lewis bases such as H₂O, NH₃, (CH₃-CH₂)₂O, PH₃ and CO. They have very unique names.

The names of such Lewis bases are shown in the table below:

Formula of Lewis Base	IUPAC Name of Lewis Base	Common Name of Lewis Base
H ₂ O	Dihydrogenmonoxide	Water
NH ₃	Nitrogen trihydride	Ammonia
(CH ₃ -CH ₂) ₂ O	Ethoxyethane	Diethyl Ether
PH ₃	Phosphorous trihydride	Phosphine
CO	Carbon monoxide	Carbonic Oxide

IUPAC stands for **International Union of Pure and Applied Chemistry**

A You Tube video explains the naming of Acids is given below:

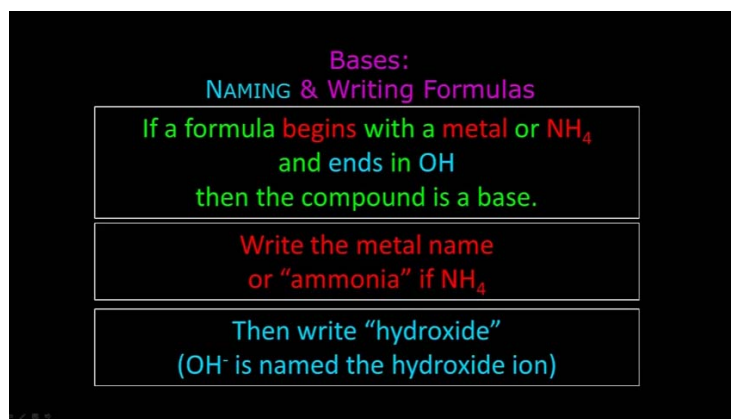
<https://www.youtube.com/watch?v=5Jb2u9ihfm4>



Naming Acids Introduction.mp4

The video below out of You Tube goes over naming Bases:

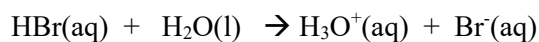
<https://www.youtube.com/watch?v=ypTV49Nu-BY>



Bases Naming and Writing.mp4

III. Conjugate Acid – Base Pairs

Bronsted – Lowry Acids and Bases when reacting, they form conjugate Acid – Base Pairs. Let us look at the Acid – Base Reaction below:



HBr is considered as Bronsted – Lowry Acid (H^+ Donor)

$\text{H}_2\text{O(l)}$ is considered as Bronsted – Lowry Base (H^+ Acceptor)

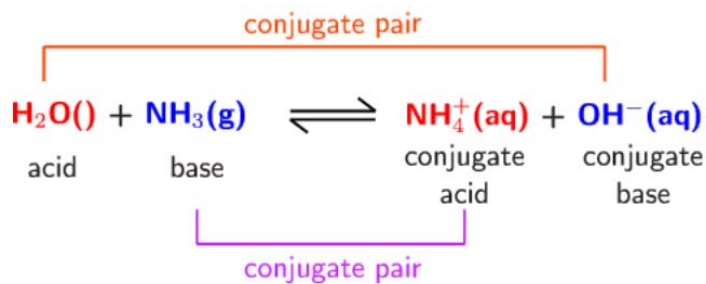
H_3O^+ is Conjugate Acid

Br^- is Conjugate Base

H_3O^+ is a conjugate Acid: It was a base (H_2O) that gained H^+ .

Br^- is a conjugate Base: It was an acid (HBr) that lost H^+ .

Another Example: <https://www.quora.com/What-is-conjugate-acid-base-pair>



A list of Acids and their Conjugate bases are listed below:

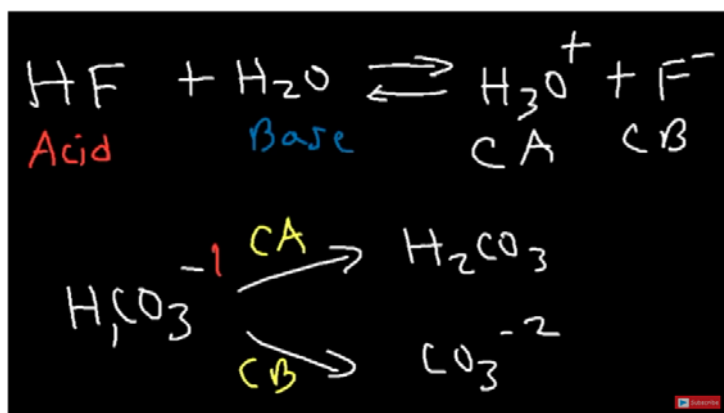
<http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch11/conjugat.php#pairs>

Typical Bronsted Acids and Their Conjugate Bases

Acid	Conjugate Base
H_3O^+	H_2O
H_2O	OH^-
OH^-	O^{2-}
HCl	Cl^-
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
NH_4^+	NH_3
NH_3	NH_2^-

A video of You Tube illustrate the concept of Conjugate Acid – Base Pairs:

<https://www.youtube.com/watch?v=OP6RKqSp1Xw>



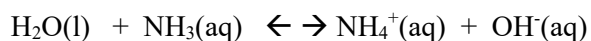
Amphoteric Compounds:

Amphoteric compounds are compounds that can act as an acid or as a base depending on the reaction they are participating in.

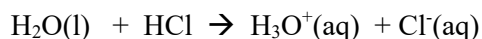
Examples of such amphoteric compounds are given amphoteric compounds:

H₂O is an amphoteric compound:

H₂O as an acid:

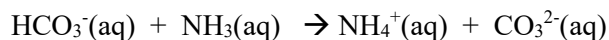


H₂O as a base:

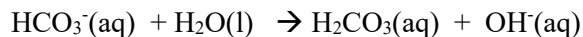


HCO₃⁻ is an amphoteric compound:

HCO₃⁻ as an acid:

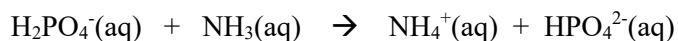


HCO₃⁻ as a base:

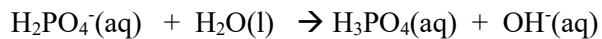


H₂PO₄⁻ is an amphoteric compound:

H₂PO₄⁻ as an acid:



H₂PO₄⁻ as a base:



IV. Strengths of Acids and Bases

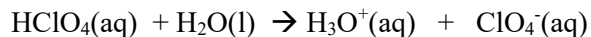
Acids and bases can dissociate in water to produce their corresponding conjugate acids – bases pairs. The degree of such dissociation will determine the strength of the acids and bases. Stronger acids and bases dissociate completely (about 100%) while weaker acids and bases dissociate partially or slightly. Weaker acids and bases stay undissociated for the most part.

Strong Acids:

There are seven strong acids that can dissociate completely (about 100%):

HCl, HBr, HI, HNO₃, H₂SO₄, HClO₃ and HClO₄

Strong Acids Dissociation:



Dissociation goes to completion (designated with one arrow)

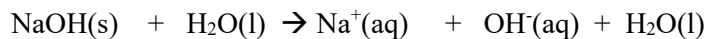
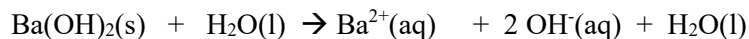
Strong Bases:

Bases are made of metals of group 1A and 2 A form strong bases that can dissociate completely (about 100%).

Group 1 A: LiOH, NaOH, KOH, RbOH, CsOH and FrOH

Group 2 A: Mg(OH)₂, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂ and Ra(OH)₂

Strong Bases Dissociation:

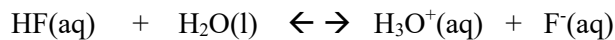


H₂O is a solvent.

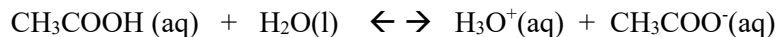
Dissociation goes to completion (designated with one arrow)

Weak Acids:

All other inorganic as well as organic acids are weak except those seven strong acids. Examples of inorganic acids are: H₂CO₃, HF, H₃PO₄ and H₂S. Examples of weaker organic acids are: CH₃COOH (acetic acid), H₂C₂O₄ (oxalic acid), HCOOH (formic acid), C₆H₈O₇ (citric acid) and C₆H₈O₆ (ascorbic acid).

Weak Acids Dissociation:

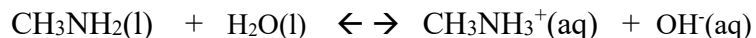
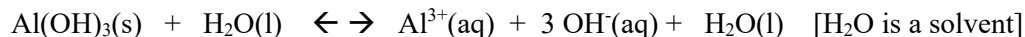
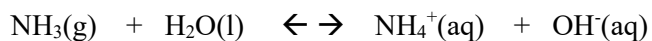
Dissociation does not go to completion (designated with double arrows)



Dissociation does not go to completion (designated with double arrows)

Weak Bases:



All other inorganic as well as organic bases are weak bases except those of group 1 A and 2 A metal hydroxides. Examples of inorganic weak bases are: H_2O , NH_3 , $\text{Fe}(\text{OH})_3$, $\text{Al}(\text{OH})_3$ and $\text{Cu}(\text{OH})_2$. Examples of organic bases are: CH_3NH_2 (methylamine), $\text{C}_5\text{H}_5\text{N}$ (pyridine), $\text{C}_3\text{H}_4\text{N}_2$ (imidazole), $\text{C}_2\text{H}_5\text{NO}_2$ (glycine) and other amines.

Weak Bases Dissociation:

Dissociation does not go to completion (designated with double arrows)

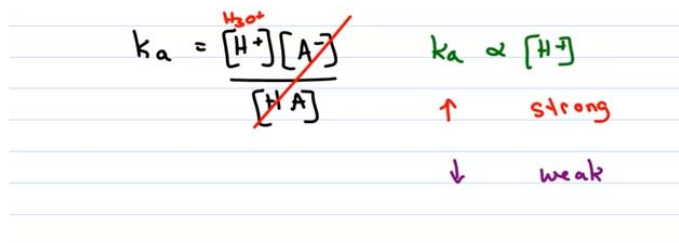
The table below illustrates the strength of acids in decreasing order with perchloric acid being the strongest acid and the strength of bases in increasing order with methide ion being the strongest base:

[https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_Chemistry_\(OpenSTAX\)/14%3A_Acid-Base_Equilibria/14.3%3A_Relative_Strengths_of_Acids_and_Bases](https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_Chemistry_(OpenSTAX)/14%3A_Acid-Base_Equilibria/14.3%3A_Relative_Strengths_of_Acids_and_Bases)

Acid			Base		
	perchloric acid	HClO_4	Do not undergo base ionization in water	ClO_4^-	perchlorate ion
	sulfuric acid	H_2SO_4		HSO_4^-	hydrogen sulfate ion
	hydrogen iodide	HI		I^-	iodide ion
	hydrogen bromide	HBr		Br^-	bromide ion
	hydrogen chloride	HCl		Cl^-	chloride ion
	nitric acid	HNO_3		NO_3^-	nitrate ion
	hydronium ion	H_3O^+	Undergo complete acid ionization in water	H_2O	water
	hydrogen sulfate ion	HSO_4^-		SO_4^{2-}	sulfate ion
	phosphoric acid	H_3PO_4		H_2PO_4^-	dihydrogen phosphate ion
	hydrogen fluoride	HF		F^-	fluoride ion
	nitrous acid	HNO_2		NO_2^-	nitrite ion
	acetic acid	$\text{CH}_3\text{CO}_2\text{H}$		CH_3CO_2^-	acetate ion
	carbonic acid	H_2CO_3		HCO_3^-	hydrogen carbonate ion
	hydrogen sulfide	H_2S		HS^-	hydrogen sulfide ion
	ammonium ion	NH_4^+		HN_3	ammonia
	hydrogen cyanide	HCN		CN^-	cyanide ion
	hydrogen carbonate ion	HCO_3^-		CO_3^{2-}	carbonate ion
	water	H_2O		OH^-	hydroxide ion
	hydrogen sulfide ion	HS^-		S^{2-}	sulfide ion
	ethanol	$\text{C}_2\text{H}_5\text{OH}$		$\text{C}_2\text{H}_5\text{O}^-$	ethoxide ion
	ammonia	NH_3		NH_2^-	amide ion
	hydrogen	H_2		H^-	hydride ion
	methane	CH_4		CH_3^-	methide ion
					

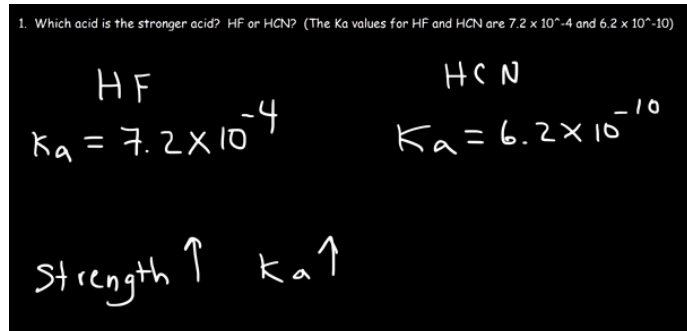
The videos below illustrate the ranking and the strength of acids and bases:

<https://www.youtube.com/watch?v=vfHnnASd3RM>



Ranking Acid Base Strength Using Ka pKa Values.mp4

https://www.youtube.com/watch?v=PdqFSw_SRug



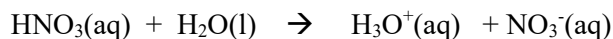
Acid Base Strength.mp4

Monoprotic, Diprotic and Triprotic Acids Dissociation:

Monoprotic acids have one proton, diprotic acids have two protons and triprotic acids have three protons. Examples are given below:

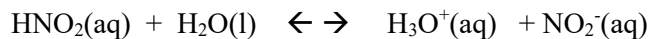
Strong Monoprotic Acids: HCl, HBr, HNO₃, HClO₄

Strong Monoprotic Acids Dissociation:



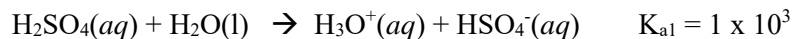
Weak Monoprotic Acids: HF, HCN, H₂O, HNO₂

Weak Monoprotic Acids Dissociation:



Strong Diprotic Acids: H₂SO₄

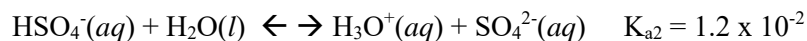
Strong Diprotic Acids Dissociation:



$K_{a1} = 1 \times 10^3$ is very large and this first of sulfuric acid dissociation is a strong acid dissociation and it is designated with one arrow indicating that the dissociation goes to completion.

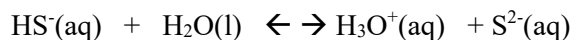
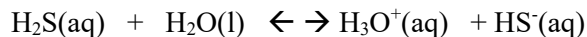
K_{a1} is called first acid dissociation constant and it will be discussed later.

The second step of sulfuric acid dissociation is a weak acid dissociation where K_{a2} is small, and it is designated with double arrows indicating that the dissociation is partial and incomplete.



Weak Diprotic Acids: H₂S, H₂CO₃

Weak Diprotic Acids Dissociation:

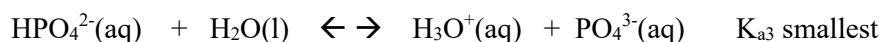
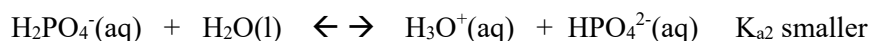
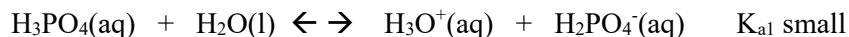


Strong Triprotic Acids: There are no strong triprotic acids.

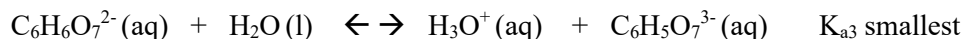
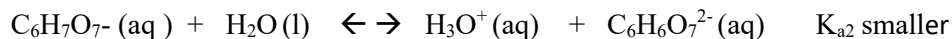
Weak Triprotic Acids: H_3PO_4 (phosphoric acid) and $\text{C}_6\text{H}_8\text{O}_7$ (citric acid)

Weak Triprotic Acids Dissociation:

Phosphoric Acid Dissociation:



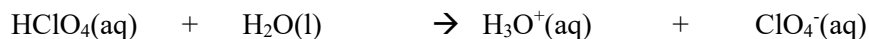
Citric Acid Dissociation:



Direction of Reaction:

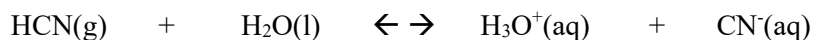
In acid - base reaction, the stronger acid and the stronger base favor the products formation where weaker acids and bases are produced:

<http://www2.ucdsb.on.ca/tiss/stretton/CHEM2/acid06.htm>



(Stronger Acid) (Stronger Base) (Weaker Acid) (Weaker Base) **Reaction favors products**

On the other hand, in weaker acids – bases reactions favor the reactants side.



Reaction favors Reactants

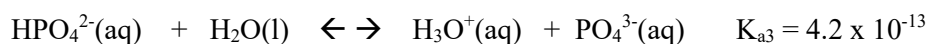
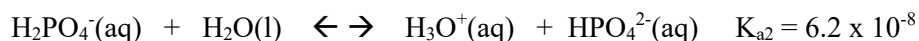
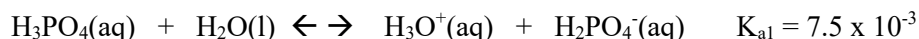
(Weaker Acid) (Weaker Base) (Stronger Acid) (Stronger Base)

V. Dissociation of Weak Acids and Bases

The dissociation of mono, di and triprotic weak acids has introduced the acid dissociation constant K_a . K_a is the ratio of the concentrations of the products over the concentrations of reactants.

Acid Dissociation Expression K_a

Phosphoric Acid Dissociation Constants at 25 °C:



$$K_{a1} = 7.5 \times 10^{-3} = \{[\text{H}_3\text{O}^+(\text{aq})] \times [\text{H}_2\text{PO}_4^-(\text{aq})]\} / [\text{H}_3\text{PO}_4(\text{aq})]$$

$$K_{a2} = 6.2 \times 10^{-8} = \{[\text{H}_3\text{O}^+(\text{aq})] \times [\text{HPO}_4^{2-}(\text{aq})]\} / [\text{H}_2\text{PO}_4^-(\text{aq})]$$

$$K_{a3} = 4.2 \times 10^{-13} = \{[\text{H}_3\text{O}^+(\text{aq})] \times [\text{PO}_4^{3-}(\text{aq})]\} / [\text{HPO}_4^{2-}(\text{aq})]$$

$[\text{H}_3\text{PO}_4(\text{aq})]$, $[\text{H}_2\text{PO}_4^-(\text{aq})]$, $[\text{HPO}_4^{2-}(\text{aq})]$ and $[\text{PO}_4^{3-}(\text{aq})]$ are molarities of the corresponding compounds and anions.

The general rule is: **The large K_a is, the Stronger the Acid is.**

A table of **Acids – Bases Strength and K_a Values** is given below:

<https://depts.washington.edu/eoopic/links/acidstrength.html>

In this table

- Strong acids are listed at the top left hand corner of the table and have K_a values > 1
- Acid with values less than one are considered weak.
- The strong bases are listed at the bottom right of the table and get weaker as we move to the top of the table.

Acids – Bases Strength and K_a Values

K_a	Acid		Base	
	Name	Formula	Formula	Name
Large	Perchloric acid	HClO_4	ClO_4^-	Perchlorate ion
3.2×10^9	Hydroiodic acid	HI	I^-	Iodide
1.0×10^9	Hydrobromic acid	HBr	Br^-	Bromide
1.3×10^6	Hydrochloric acid	HCl	Cl^-	Chloride
1.0×10^3	Sulfuric acid	H_2SO_4	HSO_4^-	Hydrogen sulfate ion
2.4×10^1	Nitric acid	HNO_3	NO_3^-	Nitrate ion
-----	Hydronium ion	H_3O^+	H_2O	Water
5.4×10^{-2}	Oxalic acid	$\text{HO}_2\text{C}_2\text{O}_2\text{H}$	$\text{HO}_2\text{C}_2\text{O}_2^-$	Hydrogen oxalate ion
1.3×10^{-2}	Sulfurous acid	H_2SO_3	HSO_3^-	Hydrogen sulfite ion
1.0×10^{-2}	Hydrogen sulfate ion	HSO_4^-	SO_4^{2-}	Sulfate ion
7.1×10^{-3}	Phosphoric acid	H_3PO_4	H_2PO_4^-	Dihydrogen phosphate ion
7.2×10^{-4}	Nitrous acid	HNO_2	NO_2^-	Nitrite ion
6.6×10^{-4}	Hydrofluoric acid	HF	F^-	Fluoride ion
1.8×10^{-4}	Methanoic acid	HCO_2H	HCO_2^-	Methanoate ion
6.3×10^{-5}	Benzoic acid	$\text{C}_6\text{H}_5\text{COOH}$	$\text{C}_6\text{H}_5\text{COO}^-$	Benzoate ion
5.4×10^{-5}	Hydrogen oxalate ion	$\text{HO}_2\text{C}_2\text{O}_2^-$	$\text{O}_2\text{C}_2\text{O}_2^{2-}$	Oxalate ion
1.8×10^{-5}	Ethanoic acid	CH_3COOH	CH_3COO^-	Ethanoate (acetate) ion
4.4×10^{-7}	Carbonic acid	CO_3^{2-}	HCO_3^-	Hydrogen carbonate ion
1.1×10^{-7}	Hydrosulfuric acid	H_2S	HS^-	Hydrogen sulfide ion
6.3×10^{-8}	Dihydrogen phosphate ion	H_2PO_4^-	HPO_4^{2-}	Hydrogen phosphate ion
6.2×10^{-8}	Hydrogen sulfite ion	HS^-	S^{2-}	Sulfite ion
2.9×10^{-8}	Hypochlorous acid	HClO	ClO^-	Hypochlorite ion
6.2×10^{-10}	Hydrocyanic acid	HCN	CN^-	Cyanide ion
5.8×10^{-10}	Ammonium ion	NH_4^+	NH_3	Ammonia
5.8×10^{-10}	Boric acid	H_3BO_3	H_2BO_3^-	Dihydrogen carbonate ion
4.7×10^{-11}	Hydrogen carbonate ion	HCO_3^-	CO_3^{2-}	Carbonate ion
4.2×10^{-13}	Hydrogen phosphate ion	HPO_4^{2-}	PO_4^{3-}	Phosphate ion
1.8×10^{-13}	Dihydrogen borate ion	H_2BO_3^-	HBO_3^{2-}	Hydrogen borate ion
1.3×10^{-13}	Hydrogen sulfide ion	HS^-	S^{2-}	Sulfide ion
1.6×10^{-14}	Hydrogen borate ion	HBO_3^{2-}	BO_3^{3-}	Borate ion
-----	water	H_2O	OH^-	Hydroxide

A table of **Base Dissociation Constant K_b values** is given below:

<https://www.ck12.org/book/CK-12-Chemistry-Concepts-Intermediate/section/21.13/>

In this table:

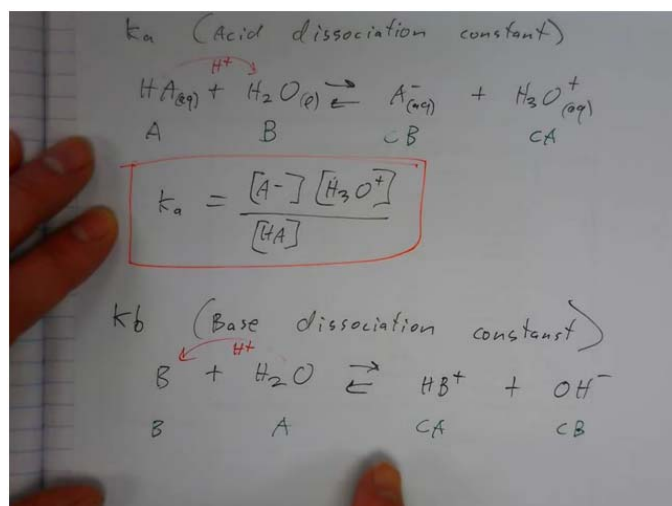
- Conjugate base of a weaker acid is a stronger base
- Conjugate base of a stronger acid is a weaker base

Base Dissociation Constant K_b Values

Name of Base	Ionization Equation	K_b
Methylamine	$\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$	5.6×10^{-4}
Ammonia	$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$	1.8×10^{-5}
Pyridine	$\text{C}_5\text{H}_5\text{N} + \text{H}_2\text{O} \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+ + \text{OH}^-$	1.7×10^{-9}
Acetate ion	$\text{CH}_3\text{COO}^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COOH} + \text{OH}^-$	5.6×10^{-10}
Fluoride ion	$\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-$	1.4×10^{-11}
Urea	$\text{H}_2\text{NCONH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{NCONH}_3^+ + \text{OH}^-$	1.5×10^{-14}

A video of You Tube illustrates the dissociation constant expression calculations:

<https://www.youtube.com/watch?v=kljVXEYMYCg>

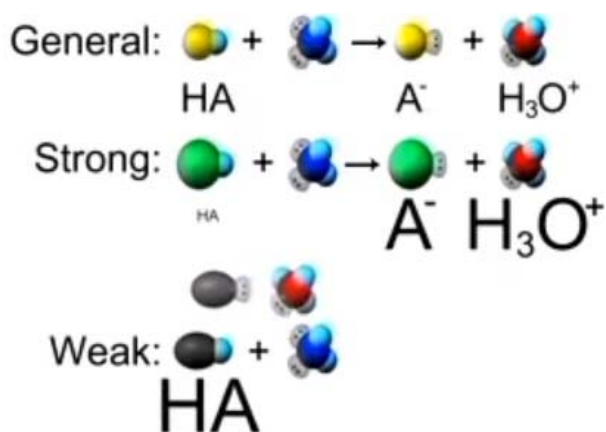




Acid Base Dissociation Constant Expressions.mp4

The dissociation of acids is illustrated in the video below:

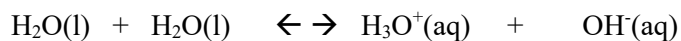
<https://www.youtube.com/watch?v=rKqYE5sZi1s>



Dissociation Strong and Weak Acids.mp4

VI. Dissociation of Water

Water is amphoteric compound which can behave as an acid or as a base depending on the condition it is in. The dissociation of water is shown below where one water molecule acts as an acid and the other molecule acts as a base:



Acid **Base** **Conjugate Acid** **Conjugate Base**

The dissociation of water is also called Auto-Ionization Water or Self-Ionization of Water.

The dissociation constant of water is called **Water Dissociation Constant**

Water Dissociation Constant Expression:

$$K = \{[H_3O^+(aq)] \times [OH^-(aq)]\} / \{[H_2O(l)] \times [H_2O(l)]\}$$

$H_2O(l)$ is pure liquid water and its concentration is not changed. Therefore, $H_2O(l)$ will be discarded from the expression and K will be replaced by K_w :

$$K_w = \{[H_3O^+(aq)] \times [OH^-(aq)]\}$$

It was determined by experiment at 25 °C that $[H_3O^+(aq)] = [OH^-(aq)] = 1.0 \times 10^{-7} \text{ mol/L}$

The value of $K_w = [1.0 \times 10^{-7}] \times [1.0 \times 10^{-7}] = 1.0 \times 10^{-14}$ at 25 °C

According to the concentrations of $[H_3O^+(aq)]$ and $[OH^-(aq)]$, one can determine the acidity and the basicity of a solution. For example:

If $[H_3O^+(aq)]$ is greater than $[OH^-(aq)]$, then the solution is said to be acidic.

If $[H_3O^+(aq)]$ is equal $[OH^-(aq)]$, then the solution is said to be neutral.

If $[OH^-(aq)]$ is greater than $[H_3O^+(aq)]$, then the solution is said to be basic.

Calculations Involving K_w , $[H_3O^+]$ and $[OH^-]$

The video link below is illustrating the concept of water dissociation:

<https://study.com/academy/lesson/dissociation-constant-and-autoionization-of-water.html>



Dissociation-Constant-and-Auto-ionization of water.mp4

Another Video You Tube illustrates the calculations of $[H_3O^+]$ and $[OH^-]$ using K_w

<https://www.youtube.com/watch?v=njgLEd9ox4A>



Calculating $[H_3O^+]$ and $[OH^-]$ Using K_w Guided Practice 1.mp4

$[H_3O^+]$ or $[OH^-]$ can be determined if one or the other is known using K_w .

Examples:

At room temperature 25°C , a solution has the $[H_3O^+] = 2.5 \times 10^{-3} \text{ M}$. Find the $[OH^-]$ of this solution.

Solution:

Using the formula below:

$$K_w = \{[H_3O^+(\text{aq})] \times [OH^-(\text{aq})]\} = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$[OH^-] = [1.0 \times 10^{-14}] / [H_3O^+] = [1.0 \times 10^{-14}] / [2.5 \times 10^{-3}] = 4.0 \times 10^{-12}$$

VII. pH

The pH is defined as a formula to describe the acidity and the basicity of a solution. The pH formula is given below:

$$\text{pH} = \text{Log}_{10}(1/[H_3O^+]) = -\text{Log}_{10}([H_3O^+])$$

$$\text{Abbreviated as } \text{pH} = -\text{Log}[H_3O^+]$$

$$\text{pOH} = \text{Log}_{10}(1/[OH^-]) = -\text{Log}_{10}([OH^-])$$

Abbreviated as $\text{pOH} = -\text{Log}[\text{OH}^-]$

Since: $K_w = \{[\text{H}_3\text{O}^+(\text{aq})] \times [\text{OH}^-(\text{aq})]\} = 1.0 \times 10^{-14}$

$$\text{p}K_w = -\text{Log} [1.0 \times 10^{-14}] = 14 = \text{p}[\text{H}_3\text{O}^+] + \text{p}[\text{OH}^-]$$

The pH scale is given below:

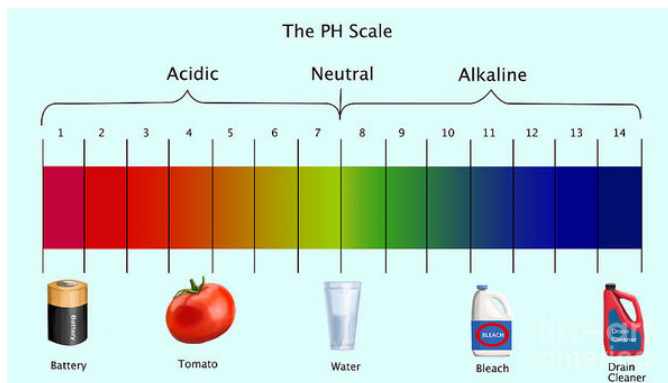
pH = 0 to 6 (acidic),

pH = 7 (neutral)

pH = 8 to 14 (basic)

The figure below depicts the scale of the pH with some examples:

<https://fineartamerica.com/featured/ph-scale-spencer-sutton.html?product=poster>



Other examples of the pH of common materials:

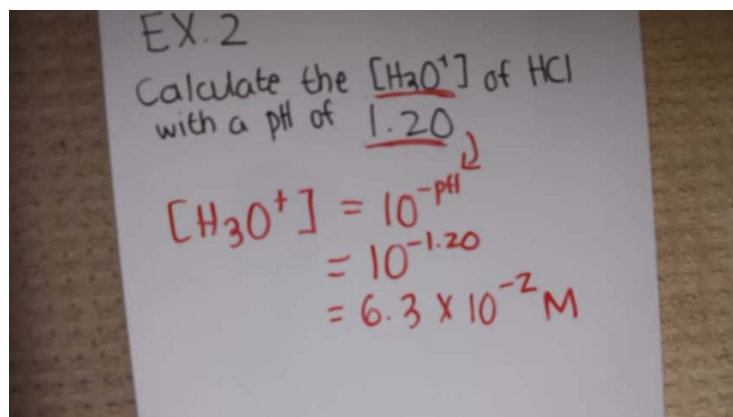
http://www1.lsbu.ac.uk/water/water_dissociation.html



The pH Calculations of Solutions

The pH calculations is illustrated in a You Tube video below:

<https://www.youtube.com/watch?v=bYo6dZ-aA-8>



pH and pOH Calculations.mp4

The pH and the pOH calculations with known molarities of H_3O^+ and OH^- examples are given below with some notes:

<http://www.everettcc.edu/files/programs/academic-resources/transitional-studies/support/tutoring-center/chemistry/w335-ph-worksheet-3.pdf>

Rule of pH significant figures:

The significant figures in the concentration of $[\text{H}^+]$ or $[\text{OH}^-]$ is equal to the number of decimal places in the pH or pOH and vice versa

- 1) What is the pH of a **0.0235** M HCl solution?

$$\text{pH} = -\log[\text{H}^+] = -\log(0.0235) = 1.629$$

Note: the molarity of HCl has 3 significant figures. Therefore, pH value will have 3 digits after the decimal. The number before the decimal is not considered.

- 2) What is the pOH of a **0.0235** M HCl solution?

$$\text{pH} = -\log[\text{H}^+] = -\log(0.0235) = 1.629$$

$$\text{pOH} = 14.000 - \text{pH} = 14.000 - 1.629 = 12.371$$

Note: the molarity of HCl has 3 significant figures. Therefore, pH value will have 3 digits after the decimal. The number before the decimal is not considered.

- 3) What is the pH of a **6.50×10^{-3}** M KOH solution?

$$\begin{aligned}\text{pOH} &= -\log[\text{OH}^-] = -\log(6.50 \times 10^{-3}) = 2.187 \\ \text{pH} &= 14.000 - \text{pOH} = 14.000 - 2.187 = 11.813\end{aligned}$$

Note: the molarity of KOH has 3 significant figures. Therefore, pH value will have 3 digits after the decimal. The number before the decimal is not considered.

- 4) A solution is created by measuring **3.60×10^{-3}** moles of NaOH and **5.95×10^{-4}** moles of HCl into a container and then water is added until the final volume is 1.00 L. What is the pH of this solution?

Since there is both acid and base we will assume a 1 mole acid:1 mole base ratio of neutralization. There is more base than acid so the leftover base is what will affect the pH of the solution.

$$3.60 \times 10^{-3} \text{ moles} - 5.95 \times 10^{-4} \text{ moles} = 3.01 \times 10^{-3} \text{ moles NaOH}$$

$$\underline{3.01 \times 10^{-3} \text{ moles NaOH} = 3.01 \times 10^{-3} \text{ M NaOH}}$$

1.00 L soln

$$\begin{aligned}\text{pOH} &= -\log[\text{OH}^-] = -\log(3.01 \times 10^{-3}) = 2.521 \text{ pH} = \\ 14.000 - \text{pOH} &= 14.000 - 2.521 = 11.479\end{aligned}$$

Note: the molarities have 3 significant figures. Therefore, pH value will have 3 digits after the decimal. The number before the decimal is not considered.

- 5) What is the pH of a 6.2×10^{-5} M NaOH solution?

$$\begin{aligned}\text{pOH} &= -\log[\text{OH}^-] = -\log(6.2 \times 10^{-5}) = 4.21 \text{ pH} = \\ 14.00 - \text{pOH} &= 14.00 - 4.21 = 9.79\end{aligned}$$

Note: the molarity of NaOH has 2 significant figures. Therefore, pH value will have 2 digits after the decimal. The number before the decimal is not considered.

- 6) A solution with a H^+ concentration of 1.00×10^{-7} M is said to be neutral. Why?

$$\text{pH} = -\log[\text{H}^+] = -\log(1.00 \times 10^{-7}) = 7.000$$

$$\text{pOH} = 14.000 - \text{pH} = 14.000 - 7.000 = 7.000$$

Note: the molarity of solution has 3 significant figures. Therefore, pH value will have 3 digits after the decimal. The number before the decimal is not considered.

$$\text{pOH} = -\log[\text{OH}^-] = 7.000 \text{ we can use this to find the } \text{OH}^- \text{ concentration}$$

$$-\log[\text{OH}^-] = 7.000$$

$$[\text{OH}^-] = 10^{-\text{pOH}} = 10^{-7.000}$$

$$[\text{OH}^-] = 1.00 \times 10^{-7} \text{ M}$$

The concentrations of H^+ and OH^- are equal, as are the pH and pOH, so the solution must be neutral.

Calculations of H_3O^+ and OH^- molarities from known pH and the pOH examples are given below with some notes:

http://sciencelas.weebly.com/uploads/5/7/8/7/57876455/11-11lab_ph_calculations_wkst-key.pdf

1. a) What is the hydrogen ion concentration of an aqueous HCl solution that has a pH of 3.0?

$$\text{pH} = 3.0$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.0} = 1 \times 10^{-3}$$

Note: the pH has 1 digit after decimal. Therefore, $[\text{H}_3\text{O}^+]$ molarity value will have 1 significant figure.

- b) What is the hydroxide ion concentration of this same solution?

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-3}$$

$$[\text{OH}^-] = K_w / [\text{H}_3\text{O}^+] = [1.0 \times 10^{-14}] / [1 \times 10^{-3}] = 1 \times 10^{-11}$$

Note: the pH has 1 digit after decimal. Therefore, $[\text{OH}^-]$ molarity value will have 1 significant figure.

- c) Which ion, H^+ or OH^- , is in greater concentration?

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-3}$$

$$[\text{OH}^-] = 1 \times 10^{-11}$$

$[\text{H}_3\text{O}^+]$ is greater than $[\text{OH}^-]$

- d) Is this solution acidic or basic?

$[\text{H}_3\text{O}^+]$ is greater than $[\text{OH}^-]$ and the solution is acidic.

2. Find the $[\text{H}^+]$ and the $[\text{OH}^-]$ of a solution with a pH of 3.494

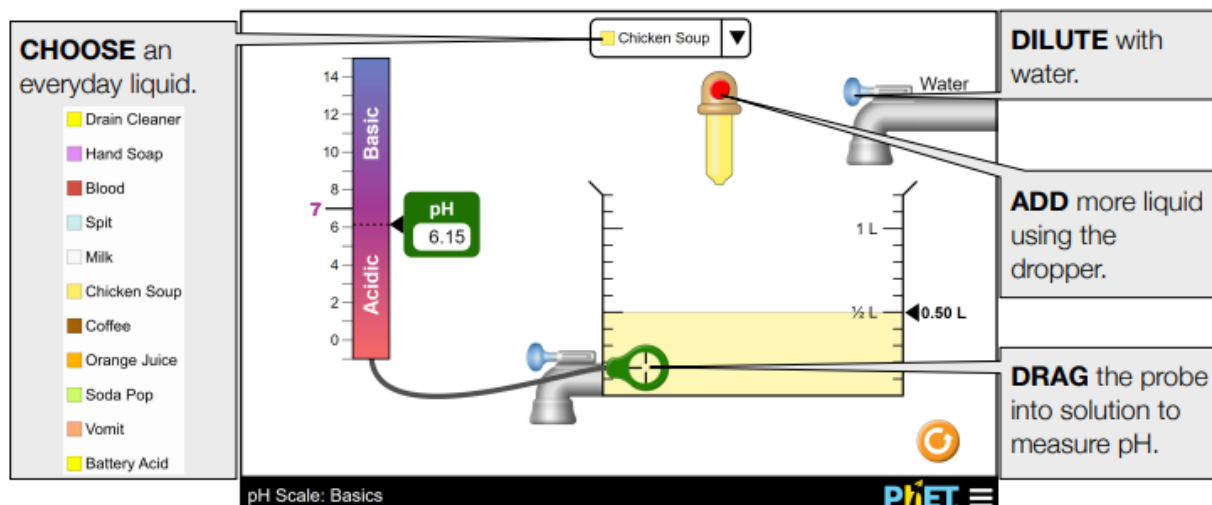
pH of 3.494

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-3.494} = 3.21 \times 10^{-4}$$

$$[\text{OH}^-] = K_w / [\text{H}_3\text{O}^+] = [1.0 \times 10^{-14}] / [3.206 \times 10^{-4}] = 3.12 \times 10^{-11}$$

Note: the pH has 3 digits after decimal. Therefore, $[H_3O^+]$ and $[OH^-]$ molarities values will have 3 significant figures.

A Phet simulation for the pH determination is given below:



Simulation Activity

Directions: Use specific examples to demonstrate each of the following learning goals.

1. Determine if a solution is acidic or basic using
 - a. pH
 - b. H_3O^+/OH^- ratio (molecular size representation of just the ions in the water equilibrium)
 - c. Hydronium/Hydroxide concentration
2. Relate liquid color to pH.
3. Predict if dilution and volume will increase, decrease or not change the pH
4. Organize a list of liquids in terms of acid or base strength in relative order with supporting evidence.
5. Write the water equilibrium expression. Describe how the water equilibrium varies with pH.

Directions: Use specific examples to demonstrate each of the following learning goals.

1. Determine if a solution is acidic or basic using
 - a. pH

- b. molecular representation
 - c. Hydronium/Hydroxide concentration
- 2. Relate liquid color to pH.
- 3. Predict if dilution and volume will increase, decrease or not change the pH
- 4. Organize a list of liquids in terms of acid or base strength in relative order with supporting evidence.
- 5. Write the water equilibrium expression. Describe how the water equilibrium varies with pH.

Is this solution acidic or basic?

$$[\text{H}_3\text{O}^+] = 3.21 \times 10^{-4}$$

$$[\text{OH}^-] = 3.12 \times 10^{-11}$$

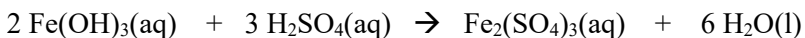
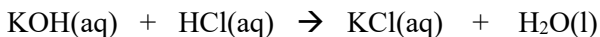
$[\text{H}_3\text{O}^+]$ is greater than $[\text{OH}^-]$, the solution is **acidic**.

VIII. Reactions of Acids and Bases

Bases Reactions:

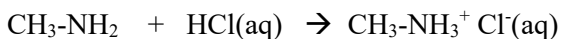
- a. Reactions of the inorganic bases with the acids are producing salts and water:

Examples:



- b. Reactions of the organic bases with the acids are producing salts only and no water:

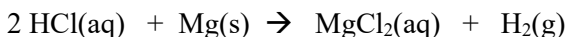
Examples:



Acids Reactions:

- Reactions of the acids with inorganic bases yield salt and water as shown in the above example.
- Reactions of the acids with organic bases yield salt only and no water as shown in the above example.
- Reactions of the acids with metals yield hydrogen and salts:

Example:



- Reactions of the acids with the bicarbonates (hydrogen carbonate) yield salt, water and carbondioxide:

Example:



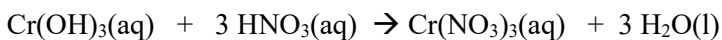
- Reactions of the acids with the carbonates yield salt, water and carbondioxide:



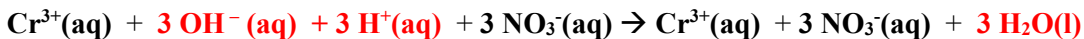
The net chemical equations of the acids with hydroxide bases:

Example:

The complete molecular equation:



The complete ionic chemical equation:



By cancelling out the spectator ions $\text{Cr}^{3+}(\text{aq})$ and $\text{NO}_3^-(\text{aq})$ from both sides of the chemical equation, one obtains the **net ionic chemical equation**.

The net ionic chemical equation:



The net ionic chemical equation above represents the **neutralization reaction** (double replacement reaction) of acids with hydroxide bases.

IX. Titration and Stoichiometry of Acid Base Reactions

Titration is a procedure by which an acid with a base are neutralized. The purpose of such procedure is to determine the concentration of an acid or a base to a precise value (several trials are needed to obtain the averaged precise value).

The process of titration is using an indicator which is an organic compound (a dye) that can change its color by changing the medium from acidic to basic or vice versa. Phenolphthalein is colorless in acidic medium and it has a pink color in basic medium.

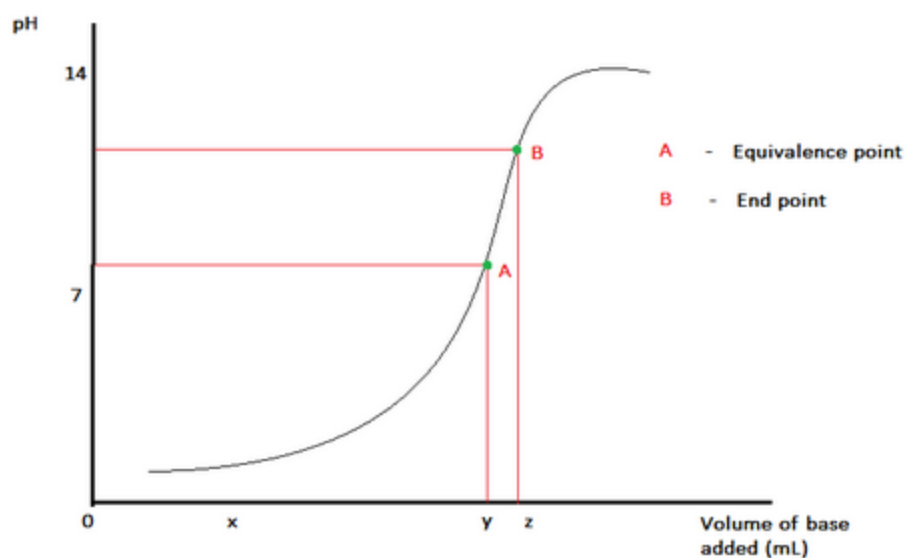
In the titration process, one can differentiate between two important points:

- a. End point: The point by which the indicator is changing its color. The signaling of color change is observed by an addition of one drop of the titrant from the buret after the neutralization of the acid with the base is reached.
- b. Equivalent point: the point by which neutralization is reached and the number of moles of the acid equal number of moles of the base.

Note that: the end point comes always after the equivalent point. One drop of the titrant signals the color change.

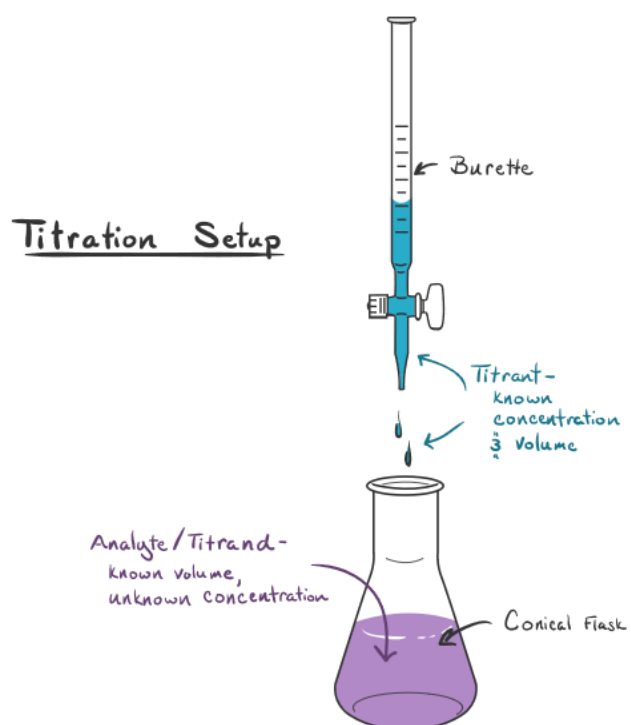
The figure below shows the difference between the end point and equivalent point of the titration of HCl with NaOH.

<http://pediaa.com/difference-between-equivalence-point-and-endpoint/>



A typical setup of the titration is given below:

<https://www.khanacademy.org/test-prep/mcat/chemical-processes/titrations-and-solubility-equilibria/a/acid-base-titration-curves>



A You Tube video goes over the titration procedure in details:

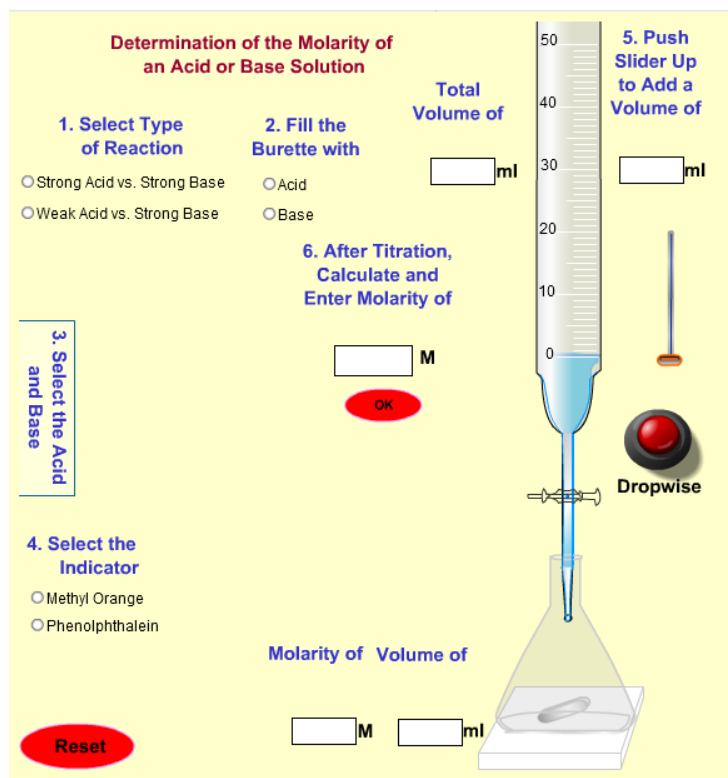
<https://www.youtube.com/watch?v=i7jnSafIMuc>



Acid-Base Titration Lab.mp4

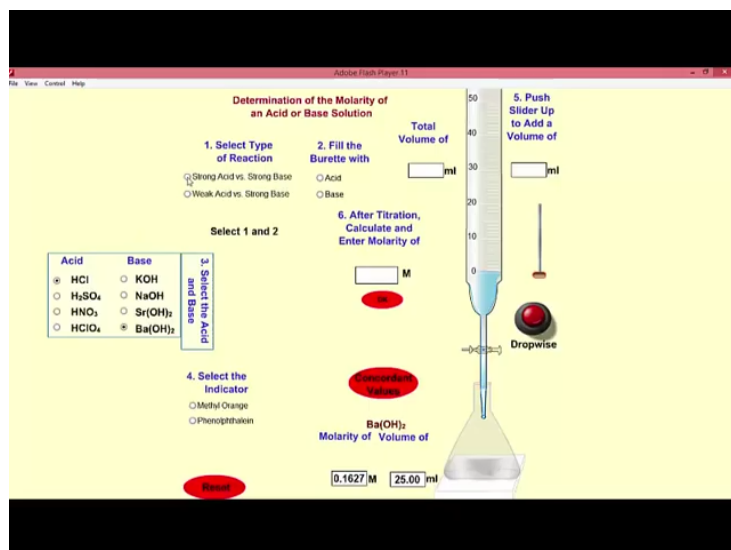
A simulation of acid base titration is illustrated below:

https://pages.uoregon.edu/tgreenbo/acid_base.html



A video of You Tube explains the simulation above:

<https://www.youtube.com/watch?v=Iruv4YoB3gA>



The directions how to use the simulation is given in the video as well in the simulation activity below:



Titration Simulation.mp4

Simulation Activity

- 1.) Select reaction type: Strong acid vs. Strong base
- 2.) Fill the burette with: Base
- 3.) Select the acid and base: Acid = HCl ; Base = NaOH
- 4.) Select the indicator: Phenolphthalein
- 5.) Push slider up to add a volume of base: Add the NaOH. You can begin by just sliding the slider up, adding about 10-12 mL of NaOH. After that, continue adding the NaOH dropwise, by clicking on the red circle. Continue adding NaOH to the flask until the solution in the flask turns pink. Once the solution has turned pink, you have reached the neutralization point, and your titration is complete.
- 6.) Print screen and copy and paste the graph into page 2.
- 7.) Calculate and enter molarity of the base: Fill in the appropriate data in the data list below using the values shown in the simulation. (Be sure to include units!) Then, calculate the molarity of NaOH used. (Hint: Use formula $M_a V_a = M_b V_b$) Enter the value you get in the box provided in the simulation. Make sure your value is correct!
- 8.) Repeat steps 1-7 for titration 2 of your choosing.

Submission Instructions

****You will be submitting only Page 2 of this document!!****

Fill in the document in word format (use equation editor where applicable)

Save too I:\MMSTC\Homework\Acid Base Titration Lab in your class folder

File Name: Last name_class

Acid-Base Titration Virtual Lab

Titration 1

1.) Write the equation for the neutralization reaction in this titration:

2.) What indicator is used? What color does the indicator change to?

3.) Data:

Molarity of Acid: _____

Volume of Acid: _____

Volume of Base: _____

4.) Calculate the molarity of the base. (Show all work, including the formula & units!)

Molarity of Base: _____

5. Graph

Titration 2

1.) Write the equation for the neutralization reaction in this titration:

2.) What indicator is used? What color does the indicator change to?

3.) Data:

Molarity of _____: _____

Volume of _____: _____

Volume of unknown: _____

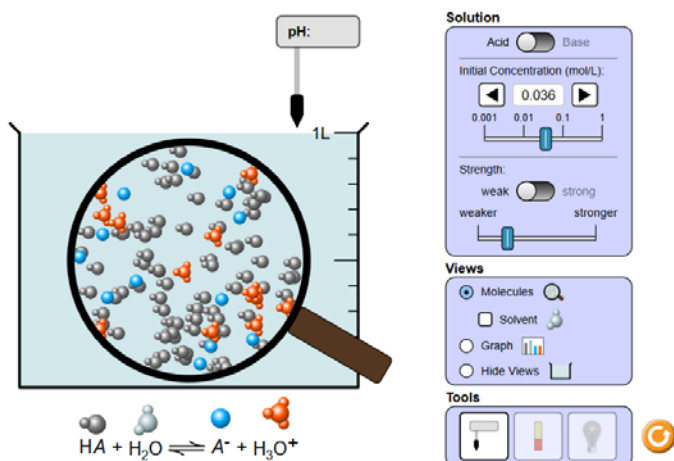
4.) Calculate the molarity of the unknown. (Show all work, including the formula & units!)

Molarity of unknown: _____

5. Graph

Another simulation is given by Phet Simulation which illustrates the Acid – Base Reaction (Neutralization):

https://phet.colorado.edu/sims/html/acid-base-solutions/latest/acid-base-solutions_en.html






Simulation Activity:




Acid Bases WebLab PhET Chromebook Version



<https://phet.colorado.edu/en/simulation/acid-base-solutions> Click <Introduction> to begin.

Part 1: Procedure

1. The lab has 2 tools that allow you to test for pH values: A probe , and pH paper . Use each one by dipping it into the solution to be tested. Try all the given types of solutions and fill in the Data Chart with the pH value 0-14.

2. The circuit with a battery and bulb as shown:  is the tool used to test for conduction of a solution. By dipping the wire leads into the solution, the bulb with either **remain unlit**, be **dimly lit**, be **somewhat bright** or **very bright**. Test each solution and record your observation for the bulbs brightness in the chart below.

Part 1: Data	pH Value from Probe	Color & pH Value from pH Paper	Observations from Circuit Tool Describe the brightness
Water (H ₂ O) 			
Strong Acid (HA) 			
Weak Acid (A) 			

Strong Base (MOH) 			
Weak Base (B) 			

Part 1: Analysis

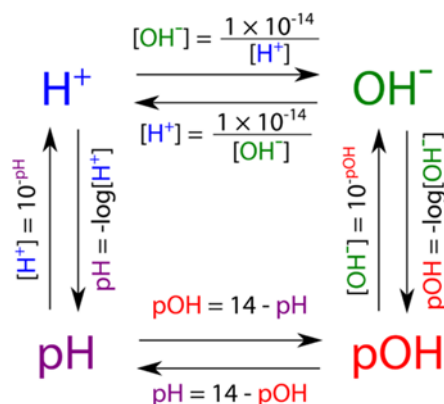
1. What pH value range is observed: a. for acids? _____ b. for bases? _____

2. *Why* are some solutions better conductors of electricity?

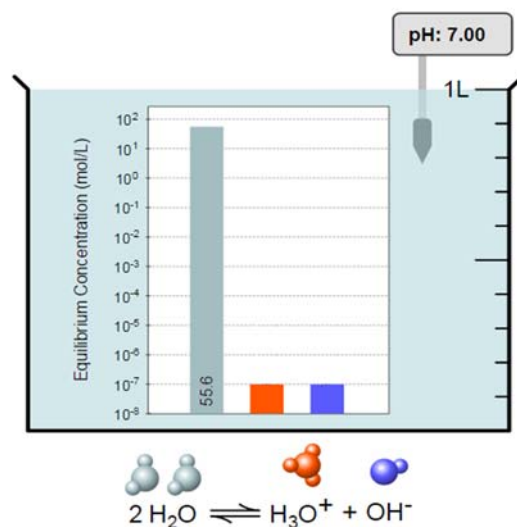
Part 2 Procedure, Data & Analysis:

Recall: The amount of ionization or dissociation of ions determines the strength of an acid or base. The concentration of $[H_3O^+]$, hydronium and $[OH^-]$, hydroxide ions can be used to calculate pH and pOH as shown on the diagram here:

Note: we use $[H_3O^+]$ and $[H^+]$ interchangeably.



- Click on Water Solution, Graph View, Probe Tool. Insert the probe in the water. Notice that the initial concentration of the solution is given before any ionization or dissociation takes place.
- Fill in the missing concentration values for the hydronium and hydroxide ions on the diagram here: Use the concentration *value* for $[H_3O^+]$ to calculate the pH. Show work:



3. Use the concentration *value for* $[OH^-]$ to calculate the pOH. Show work:

4. Did your answer to #2 match the pH given in the simulation? _____

5. Is the answer to #3 equal to: $(14 - \text{pH})$? _____ Show work: _____

6. Is the solution an acid, a base or neutral based upon the calculated pH? _____

HONORS ONLY: Attach notebook paper to show calculations for the pH and pOH for the other solutions.

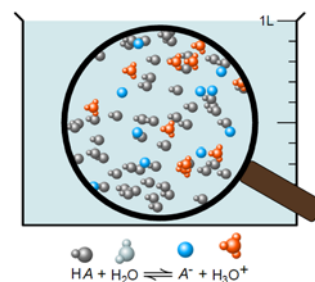
Part 3 Procedure, Analysis, Conclusion: My Solution



Across the bottom of the screen, click the **My Solution** button. The default setting shows a weak acid with a concentration of 0.010 M. Insert the pH probe to show an initial pH of 4.50. The beaker is shown below:

1. Slide the initial concentration bar to the right to increase the number of solute molecules and then slide it to the left.

What effect does changing the concentration have on the pH value? (Be specific)




2. Return to your default setting and insert the probe. Now slide the strength to the right to make the acid stronger.

a. As you increase the strength, describe the change in the number of blue A^- ions, orange H_3O^+ ions and the original HA acid:


b. As you increase the strength, describe the change in the concentrations of the ions in the solution?
Hint: Click <Graph> to see how the concentrations rise and fall.

3. Yes or No? Does the pH seem to depend upon the concentration of $[H_3O^+]$ ions?

4. We always assume that strong acids will 100% ionize in water. Click reset  and move the slider to strength: strong. Insert the probe. Record pH. Observe the number of ions in the beaker and click <Graph> to observe the concentrations.

a. pH Value = _____

b. YES or NO? Does the beaker contain particles that have not ionized and have 0% concentration? If so, what particle seems missing? _____. Why is it likely missing?

5. Click reset  and change to a base. Repeat #1-4 above and describe any different results or simply write, "Same results for bases."

#1:

#2:

#3:

#4:

Conclusions: If the answer is no, explain why not.

6. YES or NO? Can a weak acid be concentrated?

7. YES or NO? Can a strong acid be dilute?

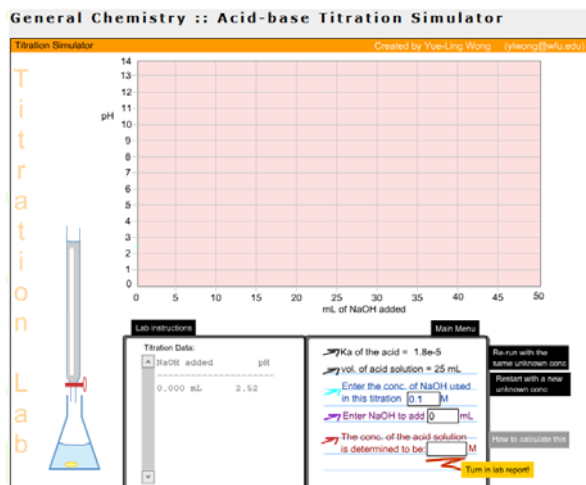
8. YES or NO? For acids, can increasing the initial concentration increase the pH?

9. YES or NO? For Bases, can increasing the initial concentration increase the pH?

Try different combinations of initial concentration and strength. Dip probe, click <Graph> to record ion concentration. Use concentration to calculate pH and verify probe data.

Another Acid Base Simulation is shown below:

<https://users.wfu.edu/~yhwong/chem/titrationsimulator/index.html>



Simulation Activity

Determination of acetic acid concentration is the desired simulation to do and will not do the rest of the simulations..

The directions are given within the simulation itself. A snapshot of such directions is given below:

Do all steps required and discussed in the simulation for the acetic acid concentration determination.

General Chemistry :: Acid-base Titration Simulator

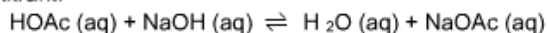
Titration Simulator

Created by Yue-Ling Wong (ylwong@wfu.edu)

Click anywhere to start the titration simulator when you are ready

About This Weak Acid Strong Base Titration Simulator

Objective of this exercise #1: To determine the concentration of a 25.0 mL solution of acetic acid (HOAc) by titration using NaOH (aq) as the titrant.



Instructions:

1. Enter the concentration of the NaOH solution you want to use in the titration.
Note: You cannot change the NaOH concentration in the middle of the titration.
2. Enter the aliquot volume of the NaOH (aq) you want to add. You can change the aliquot volume any time during the titration.
Tip: To determine the end point accurately, always add smaller aliquots when close to the end point.
3. To add the NaOH(aq) to the acid solution, click and hold down the mouse on the buret.
4. Observe the change in the pH in the graph, and determine the end point.
The solution pH versus the volume of NaOH(aq) added will be recorded.
5. Calculate the concentration of the acid solution. Enter the answer and click on the yellow tab labeled "Turn in lab report" to check your answer.

Additional activity:

Choose one or two values of mL NaOH (aq) added from the list of the titration data, and calculate the pH to see if your calculated pH matches that recorded in the titration data.

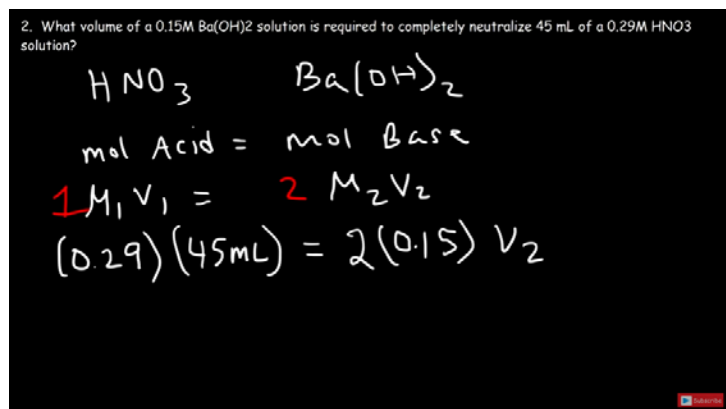
Tips:

1. You have only 50 mL of NaOH solution for each trial. If you use up 50 mL before the end point, you can always titrate again.
2. In the first trial, run a fast titration using a little higher concentration of NaOH (aq), e.g. 0.5 to 1.0 M, to get a rough estimate of the end point and the acid concentration. Then, run the titration again using a lower concentration of NaOH -- in the ranges of 0.5 to 2 times of the acid concentration, depending on the acid concentration.
3. You can re-start the titration of the same concentration of acetic acid, or a different concentration any time by clicking on the tabs on the "lab notebook".

Stoichiometry of Acid Base Reaction:

A video of You Tube explains the stoichiometry of Acid Base Reaction:

<https://www.youtube.com/watch?v=HJvALCcKYAc>



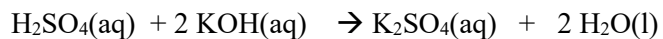
Acid Base Titration Problems Basic Introduction Calculations Examples Solution Stoichiometry.mp4

Examples Stoichiometry of Acid Base Reaction:

Mole Ratio of Acid to Base is NOT 1:1

Example:

What volume of 0.800 M KOH is needed to titrate 15.0 ml of 0.500 M H₂SO₄?



Mole Ratio of Acid to Base is 1:2

The first step is to calculate the moles of the analyte H₂SO₄:

Converting the milliliters to Liters:

$$15.0 \text{ mL} \times [1\text{L} / 1000 \text{ mL}] = 0.0150 \text{ L H}_2\text{SO}_4$$

$$0.0150 \text{ L H}_2\text{SO}_4 \times [0.500 \text{ mol H}_2\text{SO}_4 / \text{L H}_2\text{SO}_4] = 0.00750 \text{ moles H}_2\text{SO}_4$$

The second step is to calculate moles titrant KOH:

$$\text{The mole ratio of KOH : H}_2\text{SO}_4 = 2: 1$$

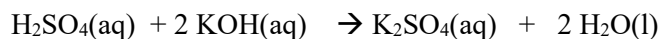
$$[0.00750 \text{ mol H}_2\text{SO}_4] \times [2 \text{ mol KOH} / 1 \text{ mol H}_2\text{SO}_4] = 0.0150 \text{ mol KOH}$$

The third step is to calculate the volume of KOH needed by dividing its number of moles by its molarity:

$$[0.0150 \text{ mol KOH}] \times [\text{L KOH} / 0.800 \text{ mol KOH}] = 0.01875 \text{ L} = 0.0188 \text{ L}$$

Example:

What mass of water can be produced if 156 grams of sulfuric acid solution H_2SO_4 react with 250. ml of 1.2 M potassium hydroxide solution KOH?



Mole Ratio of Acid to Base is 1:2

Calculations of the water mass based on H_2SO_4 :

First step to calculate number of moles H_2SO_4 using its number of molar mass:

$$\text{Molar mass of H}_2\text{SO}_4 = 2 \text{ H} + 1 \text{ S} + 4 \text{ O} = [2 \times (1.00 \text{ g/mol}) + 1 \times (32.0 \text{ g/mol}) + 4 \times (16.0 \text{ g/mol})] = 98.0 \text{ g/mol}$$

$$[156 \text{ g H}_2\text{SO}_4] \times [\text{mol H}_2\text{SO}_4 / 98.0 \text{ g H}_2\text{SO}_4] = 1.59 \text{ mol H}_2\text{SO}_4$$

The second step to calculate the mass of water based on H_2SO_4 :

$$[1.59 \text{ mol H}_2\text{SO}_4] \times [2 \text{ mol H}_2\text{O} / 1 \text{ mol H}_2\text{SO}_4] = 3.18 \text{ mol H}_2\text{O}$$

$$\text{Using the molar mass of H}_2\text{O} = 2 \text{ H} + 1 \text{ O} = 2 \times (1.00 \text{ g/mol}) + 1 \times (16.0) = 18.0 \text{ g H}_2\text{O} / \text{mol H}_2\text{O}$$

$$\text{Mass of H}_2\text{O} = [3.18 \text{ mol H}_2\text{O}] \times [18.0 \text{ g H}_2\text{O} / \text{mol H}_2\text{O}] = 57.2 \text{ g H}_2\text{O}$$

The 57.2 g H_2O is the theoretical amount of H_2O if all the H_2SO_4 is reacting to completion

Calculations of the water mass based on KOH:

First step to calculate number of moles KOH using its molarity and volume in Liters:

Converting the volume from milliliters to Liters:

$$250. \text{ mL KOH} \times [1 \text{ L} / 1000 \text{ mL}] = 0.250 \text{ L KOH}$$

$$[0.250 \text{ L KOH}] \times [1.2 \text{ mol KOH} / \text{L KOH}] = 0.300 \text{ mol KOH}$$

The second step to calculate the mass of water based on KOH:

$$[0.300 \text{ mol KOH}] \times [2 \text{ mol H}_2\text{O} / 2 \text{ mol KOH}] = 0.300 \text{ mol H}_2\text{O}$$

$$\text{Mass of H}_2\text{O} = [0.300 \text{ mol H}_2\text{O}] \times [18.0 \text{ g H}_2\text{O} / \text{mol H}_2\text{O}] = 5.40 \text{ g H}_2\text{O}$$

The 5.40 g H₂O is the theoretical amount of H₂O if all the KOH is reacting to completion.

The results above indicate that KOH is the limiting reactant because KOH produces less water than the H₂SO₄ and hence only 5.40 g of water can be made.

Example:

Determine the concentration of sulfuric acid required if it takes 28.5 ml of sulfuric acid to titrate a 36.0 ml sample of 0.450 M KOH.

What are we trying to find? That's right, the concentration or molarity of the sulfuric acid. Recall that molarity = mol solute/L solution. Since we have the volume and the molarity (concentration) of the KOH, we'll begin with it.

$$0.0360 \text{ L KOH} \left(\frac{0.450 \text{ mol KOH}}{1 \text{ L KOH}} \right) \left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol KOH}} \right) = 0.00810 \text{ mol H}_2\text{SO}_4 \text{ needed}$$

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liter of solution}} = \frac{0.00810 \text{ mol H}_2\text{SO}_4}{0.0285 \text{ L}} = 0.284 \text{ M or } 0.284 \text{ mol/L}$$

Another Example of the acid – base stoichiometry calculations:

Mole Ratio of Acid to Base is 1:1

Example:

Calculate the molarity of an acetic acid solution if 34.57 mL of this solution are needed to neutralize 25.19 mL of 0.1025 M sodium hydroxide



The first step is to calculate the moles of titrant (in this case, the NaOH):

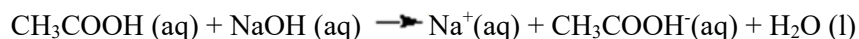
Converting the volume from milliliters to Liters:

$$25.19 \text{ mL} \times [1 \text{ L} / 1000 \text{ mL}] = 0.02519 \text{ L}$$

Multiplying the volume in Liter by molarity of NaOH:

$$0.02519 \text{ L} \times [0.1025 \text{ mol} / \text{L}] = 0.002582 \text{ mol NaOH}$$

The second step is to use the balanced chemical equation to calculate the moles of analyte (in this case, the CH_3COOH) present:



Mole Ratio of Acetic Acid to Sodium hydroxide: 1: 1

$$0.02582 \text{ mol NaOH} \times [1 \text{ mol CH}_3\text{COOH} / 1 \text{ mol NaOH}] = 0.002582 \text{ mol CH}_3\text{COOH}$$

The third step is to use the volume of analyte to find the concentration of the analyte.

Converting the volume of the analyte from mL into L:

$$0.02582 \text{ L} \times [1 \text{ L} / 1000 \text{ mL}] = 0.03457 \text{ L CH}_3\text{COOH}$$

$$\text{The concentration of acetic acid:} = 0.002582 \text{ mol CH}_3\text{COOH} / 0.03457 \text{ L CH}_3\text{COOH} = 0.07469 \text{ M}$$

X. Buffers

Buffers are solutions that resist a change of the pH and maintaining the pH of such solutions unchanged by neutralizing small amounts of added acids or bases.

Examples of such buffers solutions are:

- a. Human blood
- b. $\text{NH}_3 / \text{NH}_4^+$
- c. $\text{HCO}_3^- / \text{CO}_3^{2-}$
- d. $\text{CH}_3\text{COOH} / \text{CH}_3\text{COO}^-$
- e. $\text{H}_3\text{PO}_4 / \text{H}_2\text{PO}_4^-$

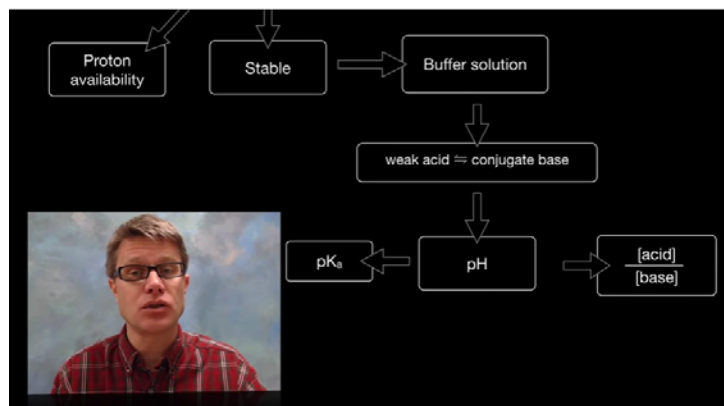
From the above examples, one can deduce that the buffer solutions are made of:

- 1. Weak Acid Mixed with Its Conjugate Base
- 2. Weak Base Mixed with its Conjugate Acid

Note: Strong Acids and Strong Bases can NOT form buffer solutions because they cannot be hydrolyzed in aqueous solutions.

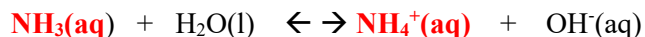
A video of you tube illustrates the concept of the buffer solutions:

<https://www.youtube.com/watch?v=rIvEvwViJGk>



pH and Buffers.mp4

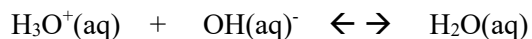
Let us consider the buffer solution system of $\text{NH}_3 / \text{NH}_4\text{Cl}$



The Buffer System is made of $\text{NH}_3 / \text{NH}_4^+$

Now let us examine the buffer function:

If an external acid $[\text{H}_3\text{O}^+]$ is added to this buffer solution, then the external acid $[\text{H}_3\text{O}^+]$ will react with $[\text{OH}^-]$ of the buffer solution to produce water:



When OH^- is removed by H_3O^+ , **the equilibrium will shift to the right to the products' side of the equilibrium.** This means the amount of NH_4^+ will slightly be increased on the expense of NH_3 which will be slightly decreased.

On the other hand, when an external base $[\text{OH}^-]$ is added to the buffer solution, then the external base $[\text{OH}^-]$ will cause **the equilibrium to shift to the side of the reactants.** This means that the amount of NH_3 will slightly be increased on the expense of NH_4^+ which will be slightly decreased.

The increase and the decrease of the concentrations of NH_3 and NH_4^+ is very small, thus the pH of the buffer solution is maintained.

pH Calculations of the Buffer Solutions

<https://www.youtube.com/watch?v=7Us44X98r-E>

2. What is the pH of a solution consisting of 0.15M NH₄Cl and 1.5M NH₃? The K_b of NH₃ is 1.8×10^{-5} .

$pH = pK_a + \log \left(\frac{B}{A} \right)$ ↖ NB

$pH > pK_a$ $pH = 10.255$

$pH = 9.255 + \log \left(\frac{1.5}{0.15} \right)$

$= 10.255$

$pK_b = -\log(K_b) = -\log(1.8 \times 10^{-5}) = 4.745$

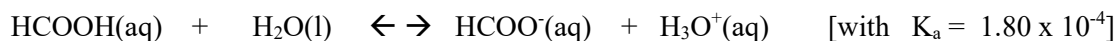
$pK_a = 14 - pK_b = 14 - 4.745 = 9.255$



Buffer Solution pH Calculations Henderson Hasselbalch Equation Explained Chemistry Problems.mp4

Example of K_a of Formic Acid

Let us consider the buffer solution: HCOOH / HCOONa [Formic Acid / Sodium Formate]:



Setting up the K_a expression:

$$K_a = \frac{[\text{HCOO}^-\text{(aq)}] \times [\text{H}_3\text{O}^+\text{(aq)}]}{[\text{HCOOH(aq)}]}$$

Isolating [H₃O⁺(aq)]:

$$[\text{H}_3\text{O}^+\text{(aq)}] = \frac{K_a \times [\text{HCOOH(aq)}]}{[\text{HCOO}^-\text{(aq)}]}$$

$$[\text{H}_3\text{O}^+\text{(aq)}] = K_a \times \frac{[\text{HCOOH(aq)}]}{[\text{HCOO}^-\text{(aq)}]}$$

Since the pH is defined as $-\text{Log} [\text{H}_3\text{O}^+]$, then taking the Log of both sides of the equations, one obtains the final equation of the pH of the buffer solution:

$$-\text{Log} [\text{H}_3\text{O}^+\text{(aq)}] = -\text{Log} \left[K_a \times \frac{[\text{HCOOH(aq)}]}{[\text{HCOO}^-\text{(aq)}]} \right]$$

$$-\text{Log} [\text{H}_3\text{O}^+(\text{aq})] = -\text{Log } K_a - \text{Log} \left\{ \frac{[\text{HCOOH}(\text{aq})]}{[\text{HCOO}^-(\text{aq})]} \right\}$$

$$\text{pH} = \text{p}K_a + \text{Log} \frac{[\text{HCOO}^-(\text{aq})]}{[\text{HCOOH}(\text{aq})]}$$

The equation in red above is known as Henderson - Hasselbalch Equation and it is used to calculate the pH of the buffer solutions.

Let us look at simple calculations of the pH of some buffer solutions:

Example:

What is the pH of a buffer solution that is 0.25 M in HF and 0.10 M in NaF? (K_a for HF is 6.8×10^{-4})

Solution:

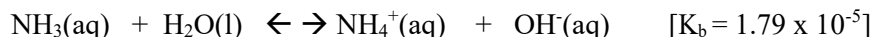
$$\text{pH} = \text{p}K_a + \text{Log} \frac{[\text{F}^-(\text{aq})]}{[\text{HF}(\text{aq})]}$$

$$\text{pH} = -\text{Log}(6.8 \times 10^{-4}) + \text{Log} (0.10 \text{ M} / 0.25)$$

$$\text{pH} = 3.16749 - 0.39794 = 2.76955 = 2.77 \quad [2 \text{ digits after decimal}]$$

Example:

Determine the pH of a buffer that is 0.250 M in NH_3 and 0.500 M in NH_4Cl . (K_b for NH_3 is 1.79×10^{-5})



$$K_b = 1.79 \times 10^{-5} = \frac{[\text{NH}_4^+(\text{aq})] \times [\text{OH}^-(\text{aq})]}{[\text{NH}_3(\text{aq})]}$$

Isolating $[\text{OH}^-(\text{aq})]$:

$$[\text{OH}^-(\text{aq})] = \frac{(1.79 \times 10^{-5}) \times [\text{NH}_3(\text{aq})]}{[\text{NH}_4^+(\text{aq})]}$$

$$[\text{OH}^-(\text{aq})] = \frac{(1.79 \times 10^{-5}) \times (0.250 \text{ M})}{(0.500 \text{ M})} = 8.95 \times 10^{-6}$$

$$K_w = 1.00 \times 10^{-14} = [\text{OH}^-] \times [\text{H}_3\text{O}^+] = [8.95 \times 10^{-6}] \times [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = (1.00 \times 10^{-14}) / (8.95 \times 10^{-6}) = 1.12 \times 10^{-9}$$

$$\text{pH} = -\text{Log} (1.12 \times 10^{-9}) = 8.950 \quad (3 \text{ digits after decimal})$$

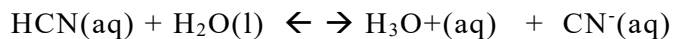
Example:

A buffer solution of pH 8.50 is desired.

Starting with 0.0100 mol of KCN and the usual inorganic reagents of the laboratory. How would you

prepare 1.00 L of the buffer solution? ($K_a = 4.93 \times 10^{-10}$)

The buffer solution will be made of the weak acid HCN and its conjugate base KCN (CN^-):



$$\text{pH} = \text{pK}_a + \text{Log} \{ [\text{CN}^-(\text{aq})] / [\text{HCN(aq)}] \}$$

$$8.50 = (-\text{Log}(4.93 \times 10^{-10}) + \text{Log} [(0.0100 \text{ mol/1.00 L}) / [\text{HCN}]]$$

$$8.50 = 9.307 + \text{Log} (0.0100 \text{ M}) / [\text{HCN}]$$

$$8.50 - 9.307 = -0.807 = \text{Log} (0.0100 \text{ M}) / [\text{HCN}]$$

$$-0.807 = -\text{Log} \{ [\text{HCN}] / (0.0100 \text{ M}) \}$$

$$0.807 = \text{Log} \{ [\text{HCN}] / (0.0100 \text{ M}) \}$$

Taking the anti Log of both sides:

$$10^{0.807} = \{ [\text{HCN}] / (0.0100 \text{ M}) \}$$

$$6.412 = \{ [\text{HCN}] / (0.0100 \text{ M}) \}$$

$$[\text{HCN}] = 6.412 \times 0.0100 = 0.0641 \text{ M}$$

Both concentrations of the buffer solutions are known:

$$[\text{HCN}] = 0.0641 \text{ M and } [\text{CN}^-] = 0.0100 \text{ M}$$

Now we have to convert the molarities into grams which can be weighed experimentally:

$$[\text{HCN}] = 0.0641 \text{ mol/L}$$

$$\text{The molar mass of HCN} = 1\text{H} + 1\text{C} + 1\text{N} = (1 \times 1.00 \text{ g/mol}) + (1 \times 12.0 \text{ g/mol}) + (1 \times 14.0 \text{ g/mol}) = 27.0 \text{ g/mol}$$

The amount of the buffer solution to be prepared is 1.00 L

Amount of grams of HCN:

$$[0.0641 \text{ mol HCN} / \text{L HCN}] \times [27.0 \text{ g HCN} / \text{mol HCN}] \times [1.00 \text{ L HCN}] = 1.73 \text{ g HCN}$$

Amount of grams of KCN:

$$[\text{CN}^-] = [\text{KCN}] = 0.0100 \text{ mol/L}$$

Molar mass of KCN = 1K + 1C + 1N = (1x39.0g/mol) + (1x12.0g/mol) + (1x14.0g/mol) = 65.0 g/mol

$[0.0100 \text{ mol KCN} / \text{L KCN}] \times [65.0 \text{ g KCN} / \text{mol KCN}] \times [1.00 \text{ L KCN}] = 0.650 \text{ g KCN}$

The buffer solution is made by weighing 1.73 grams of HCN and 0.650 grams KCN in 1.00 Liter H₂O

XI. Salts Hydrolysis

Salt hydrolysis involves the dissociation of the salt in water forming an aqueous solution.

Hydrolysis of the salt in water should yield weak acid or weak base. Hydrolysis of the salt will not yield strong acid or strong base.

The hydrolysis of the salt determines if the salt is acidic, neutral or basic when dissolved in water.

Examples:

Determine if the salts below are acidic, neutral or basic. Explain your choice.

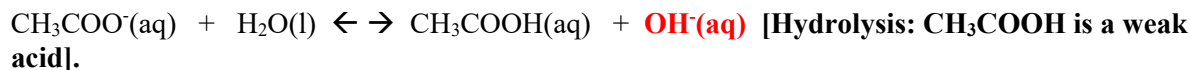
- a. CH₃COOK
- b. NH₄Cl
- c. NaCl
- d. NaNO₂
- e. NH₄CH₃COO

CH₃COOK (Potassium Acetate)

The first step is to dissolve the salt in water to produce the cation and the anion of the salt:



The second step is to hydrolyze the cation and the anion of the salt with H₂O:



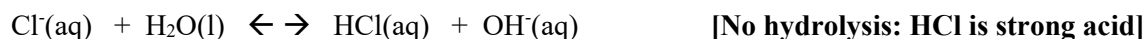
CH₃COOH is a weak acid and hence hydrolysis does occur. The hydrolysis yields the weak acid CH₃COOH and OH⁻. Thus CH₃COOK is said to be basic and it yields the basic OH⁻.

NH₄Cl (Ammonium Chloride)

The first step is to dissolve the salt in water to produce the cation and the anion of the salt:



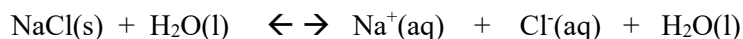
The second step is to hydrolyze the cation and the anion of the salt with H₂O:



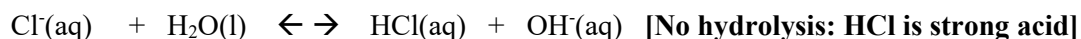
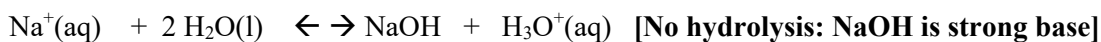
NH₄OH is a weak base and hence hydrolysis does occur. The hydrolysis yields weak base NH₄OH and **H₃O⁺**. Thus **NH₄Cl** is said to be **acidic** and it yields the acidic **H₃O⁺**.

NaCl (Sodium Chloride)

The first step is to dissolve the salt in water to produce the cation and the anion of the salt:



The second step is to hydrolyze the cation and the anion of the salt with H₂O:

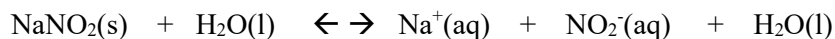


Both Na⁺ and Cl⁻ cannot be hydrolyzed. Then what is the pH of NaCl?

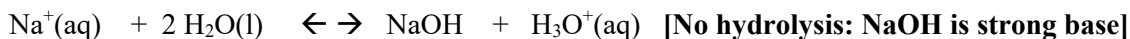
The pH of the NaCl in water is actually the pH of water itself because water is the only compound in the solution of NaCl can dissociate and can hydrolyze (auto ionization of water). Thus the pH of NaCl solution is the pH of water dissociation which equals to 7.0. Thus **NaCl** is said to be **neutral**.

NaNO₂ (Sodium Nitrite)

The first step is to dissolve the salt in water to produce the cation and the anion of the salt:



The second step is to hydrolyze the cation and the anion of the salt with H₂O:



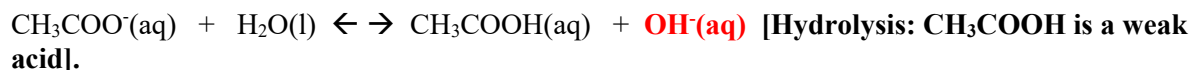
HNO₂ is a weak acid and hence hydrolysis does occur. The hydrolysis yields weak acid HNO₂ and **OH⁻**. Thus **NaNO₂** is said to be **basic** and it yields the basic **OH⁻**.

NH₄CH₃COO

The first step is to dissolve the salt in water to produce the cation and the anion of the salt:



The second step is to hydrolyze the cation and the anion of the salt with H₂O:



Both NH₄⁺ and CH₃COO⁻ are hydrolyzed. NH₄⁺ hydrolysis yields the acidic **H₃O⁺**, and CH₃COO⁻ yields the basic **OH⁻**. Then what is the pH of NH₄CH₃COO?

If both the cation and anion are hydrolyzed, then one has to inspect the K_a and K_b values of the produced weak acid and weak base respectively.



Comparing the values of K_a with K_b:

If K_a > K_b, then the salt hydrolyzed solution is acidic

If K_a < K_b, then the salt hydrolyzed solution is basic

If K_a = K_b, then the salt hydrolyzed solution neutral

NH₄CH₃COO is hydrolyzed and it does produce H₃O⁺ and OH⁻ with **K_a = K_b**. Thus **NH₄CH₃COO** is said to be **neutral**.

End of Chapter Questions:

I. Acids and Bases Character and Concepts

1. An Arrhenius acid is:
 - a) a substance that has hydroxide group and dissociates in aqueous solution to give hydroxide ion
 - b) a substance that can donate electrons
 - c) a substance that can donate hydroxide ion
 - d) a substance that has hydrogen and dissociates in aqueous solution to give hydrogen ions
2. An Arrhenius base is:
 - a) a substance that has hydroxide group and dissociates in aqueous solution to give hydroxide ion
 - b) a substance that can donate electrons
 - c) a substance that can donate hydroxide ion
 - d) a substance that has hydrogen and dissociated in aqueous solution to give hydrogen ions
3. An acid is:
 - a) a substance that tastes bitter
 - b) a substance that has no taste
 - c) a substance that tastes sour
 - d) none of the above
4. A base is:
 - a) a substance that tastes bitter
 - b) a substance that has no taste
 - c) a substance that tastes sour
 - d) none of the above
5. A Bronsted Lowry acid is:
 - a) a substance that can accept hydronium ion in an aqueous solution
 - b) a substance that can donate hydronium ion in an aqueous solution
 - c) a substance that can donate valence electrons
 - d) a substance that can accept valence electrons
6. A Bronsted Lowry base is:
 - a) a substance that can accept hydronium ion in an aqueous solution
 - b) a substance that can donate hydronium ion in an aqueous solution
 - c) a substance that can donate valence electrons
 - d) a substance that can accept valence electrons
7. A Lewis acid:

- a) a substance that can accept hydronium ion in an aqueous solution
- b) a substance that can donate hydronium ion in an aqueous solution
- c) a substance that can donate valence electrons
- d) a substance that can accept valence electrons

8. A Lewis base:

- a) a substance that can accept hydronium ion in an aqueous solution
- b) a substance that can donate hydronium ion in an aqueous solution
- c) a substance that can donate valence electrons
- d) a substance that can accept valence electrons

II. Naming Acids and Bases

9. Name the base CsOH:

- a) cesium dihydroxide
- b) cesium oxyhydrogen
- c) cesium hydroxide
- d) cesium – I – hydroxide

10. Name the acid HClO

- a) hydrochloric acid
- b) chloric acid
- c) perchloric acid
- d) hypochlorous acid
- e) chlorous acid

11. Name the base Fe(OH)₃

- a) Iron hydroxide
- b) Iron -II- hydroxide
- c) Iron -III- hydroxide
- d) Iron -I- hydroxide

12. A Bronsted Lowry acid is:

- a) a substance that can accept hydronium ion in an aqueous solution
- b) a substance that can donate hydronium ion in an aqueous solution
- c) a substance that can donate valence electrons
- d) a substance that can accept valence electrons

13. Name the base Be(OH)₂

- a) beryllium -II- hydroxide

- b) beryllium -I- hydroxide
- c) beryllium hydroxide
- d) none of the above

14. Name the acid H_2Se

- a) hydroselenic -I- acid
- b) hydroselenic -II- acid
- c) hydroselenic acid
- d) none of the above

15. Name the base $\text{Mn}(\text{OH})_4$

- a) manganese -I- hydroxide
- b) manganese -II- hydroxide
- c) manganese -III- hydroxide
- d) manganese -IV- hydroxide
- e) manganese hydroxide

16. Name the acid H_2S

- a) sulfuric acid
- b) sulfurous acid
- c) hyposulfurous acid
- d) hydrosulfuric acid
- e) none of the above

17. Name the base $\text{Al}(\text{OH})_3$

- a) aluminum -III- hydroxide
- b) aluminum hydroxide
- c) aluminum -I- hydroxide
- d) none of the above

18. Name the acid HClO_4

- a) chloric acid
- b) chlorous acid
- c) perchloric acid
- d) hypochlorous acid

19. The formula of hydrochromic acid is:

- a) H_2CrO_4
- b) $\text{H}_2\text{Cr}_2\text{O}_3$
- c) H_2Cr

d) H_2CrO

20. The formula of hyponitrous acid:

- a) HNO_3
- b) HNO
- c) HNO_2
- d) none of the above

21. The formula of the base lead -IV- oxide:

- a) $\text{Pb}(\text{OH})_2$
- b) $\text{Pb}(\text{OH})$
- c) $\text{Pb}(\text{OH})_4$
- d) $\text{Pb}_2(\text{OH})_3$

22. The formula of the base ammonium hydroxide:

- a) $\text{NH}_4(\text{OH})_2$
- b) $\text{NH}_4(\text{OH})_3$
- c) $(\text{NH}_4)_3(\text{OH})_2$
- d) NH_4OH
- e) none of the above

23. The formula of iodous acid:

- a) HI
- b) HIO
- c) HIO_3
- d) HIO_2
- e) none of the above

24. The formula of base calcium hydroxide:

- a) $\text{Ca}(\text{OH})$
- b) $\text{Ca}(\text{OH})_2$
- c) $\text{Ca}(\text{OH})_3$
- d) $\text{Ca}_2(\text{OH})$
- e) none of the above

25. The formula of hydroxic acid:

- a) HO
- b) H_2O_2
- c) H_2O
- d) H_2O_3

e) none of the above

26. The formula of the base vanadium pentahydride:

- a) V_5H
- b) V_5H_3
- c) V_3H_5
- d) VH_5
- e) none of the above

27. The formula of manganic acid:

- a) $MnOH$
- b) $HMnO_2$
- c) $HMnO_4$
- d) $HMnO_3$
- e) none of the above

28. The formula of the base nitrogen trihydride (ammonia):

- a) NH
- b) NH_3
- c) NH_2
- d) NHO
- e) none of the above

29. The formula of the pertitanic acid:

- a) $HTiO_4$
- b) $HTiO_3$
- c) $HTiO_2$
- d) $HTiO$
- e) none of the above

30. The formula of the base tin-IV-hydroxide:

- a) $Sn(OH)_4$
- b) $Sn(OH)_3$
- c) $Sn(OH)_2$
- d) $Sn(OH)$
- e) one of the above

III. Conjugate Acid – Base Pairs

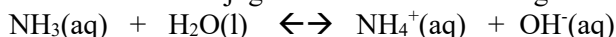
31. What is the conjugate acid?

- a) an ion that is produced when an acid accepts hydrogen ion
- b) an ion that is produced when a base donates hydrogen ion
- c) an ion that is produced when an acid donates hydrogen ion
- d) an ion that is produced when a base accepts hydrogen ion

32. What is the conjugate base?

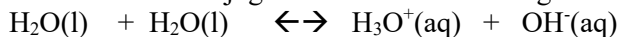
- a) an ion that is produced when an acid accepts hydrogen ion
- b) an ion that is produced when a base donates hydrogen ion
- c) an ion that is produced when an acid donates hydrogen ion
- d) an ion that is produced when a base accepts hydrogen ion

33. Determine the conjugate acid in the following reaction:



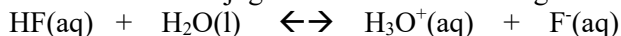
- a) NH_3
- b) H_2O
- c) NH_4^+
- d) OH^-

34. Determine the conjugate base in the following reaction:



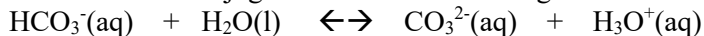
- a) H_2O
- b) H_3O^+
- c) OH^-
- d) none of the above

35. Determine the conjugate acid in the following reaction:



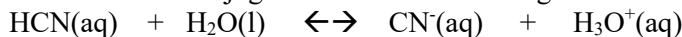
- a) HF
- b) H_2O
- c) H_3O^+
- d) F^-

36. Determine the conjugate base in the following reaction:



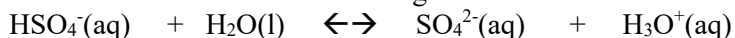
- a) HCO_3^-
- b) H_2O
- c) CO_3^{2-}
- d) H_3O^+

37. Determine the conjugate base in the following reaction:



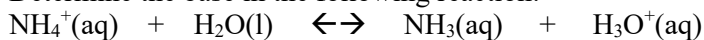
- a) HCN
- b) H_2O
- c) CN^-
- d) H_3O^+

38. Determine the acid in the following reaction:



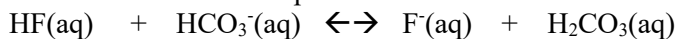
- a) HSO_4^-
- b) H_2O
- c) SO_4^{2-}
- d) H_3O^+

39. Determine the base in the following reaction:



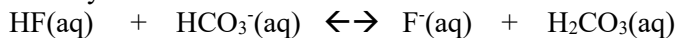
- a) NH_4^+
- b) H_2O
- c) NH_3
- d) H_3O^+

40. Which is an acid – base pair in the reaction?



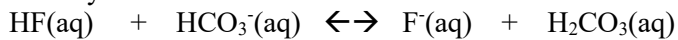
- a) HF and HCO_3^-
- b) F^- and H_2CO_3
- c) HF and H_2CO_3
- d) F^- and HCO_3^-
- e) F^- and HF

41. Identify the acids in the reaction?



- a) HF and HCO_3^-
- b) F^- and H_2CO_3
- c) HF and H_2CO_3
- d) F^- and HCO_3^-
- e) F^- and HF

42. Identify the bases in the reaction?



- a) HF and HCO_3^-
- b) F^- and H_2CO_3
- c) HF and H_2CO_3
- d) F^- and HCO_3^-
- e) F^- and HF

43. The conjugate base of H_2PO_4^-

- a) HPO_4^{2-}
- b) H_3PO_4
- c) PO_4^{3-}
- d) none of the above

44. The conjugate acid of HSO_4^-

- a) H_2SO_4
- b) SO_4^{2-}
- c) H_3SO_4
- d) none of the above

IV. Strengths of Acids and Bases

45. Determine the weakest acid:

- a) HSO_4^- (K_a at 25°C is 1.2×10^{-2})
- b) CH_3COOH (K_a at 25°C is 1.8×10^{-5})
- c) none of the above

46. Determine the strongest acid:

- a) HNO_2 (K_a at 25°C is 4.6×10^{-4})
- b) HF (K_a at 25°C is 3.5×10^{-4})
- c) HCNO (K_a at 25°C is 2.0×10^{-4})
- d) HCOOH (K_a at 25°C is 1.8×10^{-4})
- e) none of the above

47. Determine the strongest acid:

- a) HClO (K_a at 25°C is 2.9×10^{-8})
- b) HBrO (K_a at 25°C is 2.8×10^{-9})
- c) HCN (K_a at 25°C is 4.9×10^{-10})
- d) none of the above

48. Determine the weakest acid:

- a) H_2CO_3 (K_a at 25°C is 4.8×10^{-11})
- b) H_3PO_4 (K_a at 25°C is 7.5×10^{-3})
- c) $\text{C}_6\text{H}_5\text{OH}$ (K_a at 25°C is 1.6×10^{-10})
- d) H_2O (K_a at 25°C is 1.0×10^{-14})
- e) none of the above

49. Determine the weakest base:

- a) ethylamine ($\text{C}_2\text{H}_5\text{NH}_2$) (K_b at 25°C is 4.5×10^{-4})
- b) butylamine ($\text{C}_4\text{H}_9\text{NH}_2$) (K_b at 25°C is 4.0×10^{-4})
- c) propylamine ($\text{C}_3\text{H}_7\text{NH}_2$) (K_b at 25°C is 3.5×10^{-4})
- d) none of the above

50. Determine the strongest base:

- a) trimethylamine ($\text{C}_3\text{H}_9\text{N}$) (K_b at 25°C is 6.4×10^{-5})
- b) allylamine ($\text{C}_3\text{H}_5\text{NH}_2$) (K_b at 25°C is 3.1×10^{-5})
- c) benzylamine ($\text{C}_6\text{H}_5\text{CH}_2\text{NH}_2$) (K_b at 25°C is 2.2×10^{-5})
- d) ammonia (NH_3) (K_b at 25°C is 1.8×10^{-5})

e) none of the above

51. Determine the strongest base:

- a) morpholine ($\text{C}_4\text{H}_9\text{NO}$) (K_b at 25°C is 3.2×10^{-6})
- b) strychnine ($\text{C}_{21}\text{H}_{22}\text{N}_2\text{O}_2$) (K_b at 25°C is 1.8×10^{-6})
- c) ethyleneimine ($\text{C}_2\text{H}_5\text{N}$) (K_b at 25°C is 1.1×10^{-6})
- d) hydrazine (N_2H_4) (K_b at 25°C is 1.0×10^{-6})
- e) none of the above

52. Determine the weakest base:

- a) hydroxyl amine (NH_2OH) (K_b at 25°C is 8.8×10^{-9})
- b) pyridine ($\text{C}_5\text{H}_5\text{N}$) (K_b at 25°C is 1.7×10^{-9})
- c) aniline ($\text{C}_6\text{H}_5\text{NH}_2$) (K_b at 25°C is 7.5×10^{-10})
- d) none of the above

53. Determine the strongest base:

- a) urea [$(\text{NH}_2)_2\text{CO}$] (K_b at 25°C is 1.0×10^{-14})
- b) codeine ($\text{C}_{18}\text{H}_{21}\text{O}_3\text{N}$) (K_b at 25°C is 8.9×10^{-7})
- c) hydrazinium ion (N_2H_5^+) (K_b at 25°C is 8.9×10^{-16})
- d) none of the above

V. Dissociation of Weak Acids and Bases

54. Which acid among the following weak acids dissociates to give the **strongest** conjugate base?

[Hint: use K_a values from the table]

- a) HClO
- b) CH_3COOH
- c) HF
- d) HNO_2
- e) HCN

55. Which acid among the following weak acids dissociates to give the **weakest** conjugate base?

[Hint: use K_a values from the table]

- a) HClO
- b) CH_3COOH
- c) HF
- d) HNO_2
- e) HCN

56. Which acid among the following weak acids dissociates to give the **weakest** conjugate base?

[Hint: use K_a values from the table]

- a) H_2CO_3
- b) NH_4^+
- c) HCO_3^-

d) H_2O

57. Which acid among the following weak acids dissociates to give the **strongest** conjugate base?
[Hint: use K_a values from the table]

- a) H_2CO_3
- b) NH_4^+
- c) HCO_3^-
- d) H_2O

VI. Dissociation of Water

58. What is the hydronium ion concentrations in a $3.0 \times 10^{-2} \text{ M}$ NaOH solution?

- a) $4.0 \times 10^{-12} \text{ M}$
- b) $3.3 \times 10^{-13} \text{ M}$
- c) $2.5 \times 10^{-11} \text{ M}$
- d) $1.7 \times 10^{-10} \text{ M}$
- e) none of the above

59. What is the hydroxide ion concentrations in a $3.0 \times 10^{-2} \text{ M}$ NaOH solution?

- a) $4.0 \times 10^{-4} \text{ M}$
- b) $3.3 \times 10^{-3} \text{ M}$
- c) $2.5 \times 10^{-1} \text{ M}$
- d) $3.0 \times 10^{-2} \text{ M}$
- e) none of the above

60. What is formula used to determine the concentration of the hydroxide ion $[\text{OH}^-]$ in an aqueous solution using the water dissociation constant K_w ?

- a) $[\text{OH}^-] = [\text{H}_3\text{O}^+] \times K_w$
- b) $[\text{OH}^-] = [\text{H}_3\text{O}^+] / K_w$
- c) $[\text{OH}^-] = K_w / [\text{H}_3\text{O}^+]$
- d) none of the above

61. What is formula used to determine the concentration of the hydronium ion $[\text{H}_3\text{O}^+]$ in an aqueous solution using the water dissociation constant K_w ?

- a) $[\text{H}_3\text{O}^+] = [\text{OH}^-] \times K_w$
- b) $[\text{H}_3\text{O}^+] = [\text{OH}^-] / K_w$
- c) $[\text{H}_3\text{O}^+] = K_w / [\text{OH}^-]$
- d) none of the above

62. What is the hydroxide ion concentrations in a $1.0 \times 10^{-2} \text{ M}$ HCl solution?

- a) $2.0 \times 10^{-10} \text{ M}$
- b) $3.3 \times 10^{-11} \text{ M}$
- c) $1.0 \times 10^{-12} \text{ M}$
- d) $3.0 \times 10^{-13} \text{ M}$

e) none of the above

63. What is the hydronium ion concentrations in a 1.0×10^{-2} M HCl solution?

- a) 2.0×10^{-1} M
- b) 1.0×10^{-2} M
- c) 3.0×10^{-3} M
- d) 4.0×10^{-4} M
- e) none of the above

64. What is the hydroxide ion concentration in 2.50 M $\text{Ca}(\text{OH})_2$?

- a) 2.50×10^{-1} M
- b) 2.0×10^{-2} M
- c) 3.0×10^{-3} M
- d) 2.5 M
- e) none of the above

65. What is the hydronium ion concentration in 2.50 M $\text{Ca}(\text{OH})_2$?

- a) 2.50×10^{-12} M
- b) 2.0×10^{-13} M
- c) 4.0×10^{-15} M
- d) 2.5×10^{-11} M
- e) none of the above

66. Which ion [hydronium H_3O^+ or hydroxide OH^-] has higher concentration in acidic solution?

- a) both ions have the same concentrations
- b) H_3O^+
- c) OH^-
- d) none of the above

67. Which ion [hydronium H_3O^+ or hydroxide OH^-] has higher concentration in basic solution?

- a) both ions have the same concentrations
- b) H_3O^+
- c) OH^-
- d) none of the above

68. Which ion [hydronium H_3O^+ or hydroxide OH^-] has higher concentration in neutral solution?

- a) both ions have the same concentrations
- b) H_3O^+
- c) OH^-
- d) none of the above

69. Which statement is true for a basic solution?

- a) H_3O^+ must be more than 1.0×10^{-7} M
- b) OH^- must be more than 1.0×10^{-7} M
- c) OH^- must be equal to 1.0×10^{-7} M

d) none of the above

70. Which statement is true for an acidic solution?

- a) H_3O^+ must be more than $1.0 \times 10^{-7} \text{ M}$
- b) OH^- must be more than $1.0 \times 10^{-7} \text{ M}$
- c) H_3O^+ must be equal to $1.0 \times 10^{-7} \text{ M}$
- d) none of the above

71. An acidic solution contains H_3O^+ while a basic solution contains OH^- . Is this statement true or false?

- a) true
- b) false
- c) cannot be determined
- d) none of the above

72. Determine if the solution with hydroxide ion $[\text{OH}^-]$ concentration of 1.65×10^{-8} is acidic, basic or neutral.

- a) basic
- b) acidic
- c) neutral
- d) none of the above

73. Determine if the solution with hydronium ion $[\text{H}_3\text{O}^+]$ concentration of 3.89×10^{-7} is acidic, basic or neutral.

- a) basic
- b) acidic
- c) neutral
- d) none of the above

74. Determine if the solution with hydroxide ion $[\text{OH}^-]$ concentration of 1.00×10^{-7} is acidic, basic or neutral.

- a) basic
- b) acidic
- c) neutral
- d) none of the above

VII. pH

75. What is the pH of 0.00455 M HCl?

- a) 2.342
- b) 3.415
- c) 1.984
- d) 2.980
- e) none of the above

76. What is the pOH of 0.00455 M HCl?

- a) 10.976
- b) 11.658
- c) 12.363
- d) 13.897
- e) none of the above

77. What is the pH of the solution 6.50×10^{-3} M LiOH?

- a) 9.097
- b) 12.432
- c) 10.678
- d) 11.813
- e) none of the above

78. What is the pH of the solution 3.233×10^{-3} M HNO₃?

- a) 3.3423
- b) 3.4653
- c) 2.4904
- d) 2.9876
- e) none of the above

79. What is the pOH of the solution 3.233×10^{-3} M HNO₃?

- a) 9.1123
- b) 10.2901
- c) 12.3871
- d) 11.5096
- e) none of the above

80. A solution is created by measuring 3.60×10^{-3} moles of NaOH and 5.95×10^{-4} moles of HCl into a container and then water is added until the final volume is 1.00 L. What is the pH of this solution?

- a) 11.479
- b) 10.429
- c) 9.984
- d) 12.547
- e) none of the above