Acids and Bases

What Are Acids and Bases?

- Both are corrosive to most things
  - Acids to metals, and most other stuff
  - Bases not usually to metals, but to most other stuff
- When you combine an acid with a base you get a solution that is less acidic and less basic. (This means when you combine 2 things that are corrosive, you end out with a less corrosive solution.)

What Are Acids and Bases?

- Acids
  - Taste Sour
  - Corrode Metals
  - Change Litmus Red
  - Become Less Sour and Corrosive when mixed with Bases
- Bases
  - Feel Slippery
  - Change Litmus Blue
  - Become Less Corrosive when mixed with Acids.

Models of Acids and Bases

There are many ways of classifying acids and bases. These have developed over time as our understanding of underlying chemistry has increased and as our needs have changed.

- **Arrhenius Concept of Acids and Bases**
- **Brϕnsted - Lowry concept of Acids and Bases**
- **Lewis Definition of Acids and Bases**

Acids and Bases

Arrhenius Concept of Acids and Bases (1887)

According to Savante Arrhenius acids form hydrogen ions and bases form hydroxide ions when dissolved in water.

In this definition an acidic solution is one with plenty of H^+ ions and few OH^- ions.

A basic solution has few H^+ ions and many OH^- ions.

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What

- In basic solutions it is the OH^- that makes the solution behave like a base. In KOH, NaOH, $Ba(OH)_2$, solutions it is the hydroxyde ion that makes the solution, corrosive, feel slipery, etc.
- In acidic solutions it is H+ that cause the sour taste, corrosiveness, and the reaction with metals.

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Where do the hydrogen ions and hydroxide ions come from.

- When HCl is dissolved in water it gives H^+ ions:

$$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

- When NaOH dissolves in water it gives OH^- ion:

$$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

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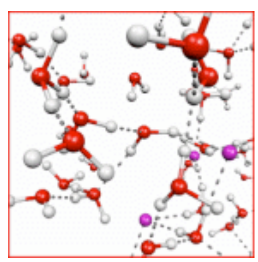
What happens to HCl in water?



•The H^+ combine with H_2O to form H_3O^+.

•The Cl^- are solvated by the water just as they would if the Cl^- were from NaCl or $CaCl_2$.

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H^+ concentration

- What would be the concentration of H^+ in a 1.500 L solution that contained 0.0598 moles of HCl ?

$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

- How many moles of H^+ would there be in the solution? There is 1 H^+ for every HCl so there will be 0.0598 moles of H^+.
- To get the molarity (M)…

$$[H^+] = \left(\frac{0.598 moles H^+}{1.500 L solution}\right) = 0.399 M H^+$$

$$= 0.399 M HCl$$

- How much of 12 M HCl do you need to make 0.350 L of a 0.010M solution?

1. How many moles HCl are needed to make 0.350 L of a 0.010 M HCl solution?

$$0.350 L\ Solution_{(dilute)} \left(\frac{0.010\ moles\ HCl}{1 L\ Solution_{(dilute)}}\right) = 0.0035\ moles\ HCl$$

2. How many L of 12M HCl are needed to get 0.0035 moles HCl?

$$0.0035\ moles\ HCl \left(\frac{1 L\ Solution_{(conc.)}}{12\ moles\ HCl}\right) = 0.00029 L solution_{(conc.)}$$

$$= 0.29 ml$$

Acidity-How Acidic is a solution?

- The 12 M HCl solution or the 0.01M HCl solution? (12M)
- The 12 M HCl is more acidic.
- What if you had an acid that gave up 2 H^+ not just 1. AH_2

$AH_2 \rightarrow A^{2-} + 2H^+$

- Would 1 M solution of HCl and a 1 M solution of AH_2 be equally acidic/corrosive? The solution made with AH2 has 2x the number of H+ so it is expected to be more corrosive

Basic Solution

- When NaOH is dissolved in water it dissociates completely, to give Na^+ and OH- ions.
  - $NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
- If 0.349 moles of NaOH are in a 1.4 L solution. What is the molarity OH^- in the resulting solution?

$$?\ M OH^- = \left(\frac{moles OH^-}{L solution}\right) = \left(\frac{0.349 moles OH^-}{1.4 L solution}\right)$$

$$= 0.249 M OH^- = 0.25 M OH^-$$

Basic Solution

- When $Ba(OH)_2$ dissolved in water it dissociates completely, to give Ba^+ and OH- ions.
  - $Ba(OH)_{2(s)} \rightarrow Ba^+_{(aq)} + 2OH^-_{(aq)}$
- If 0.349 moles of $Ba(OH)_2$ are in a 1.4 L solution. What is the molarity OH^- in the resulting solution?

$? \quad M\,OH^- =$

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Mixing Acidic And Basic Solutions

- As was mentioned earlier when acidic and basic solutions are mixed the resulting solution is less basic than the initial solution. Why does this happen?

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Neutralization-Why mixing 2 corrosive solutions can make a less corrosive one.

$HCl_{(g)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$ (stong acid)

$NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$ (strong base)

Put acidic and basic solutions together

$Na^+_{(aq)} + OH^-_{(aq)} + Cl^-_{(aq)}) + H^+_{(aq)}$

$\rightarrow H_2O + Cl^-_{(aq)} + Na^+_{(aq)}$

(Take out the ions that are on both sides of the equation, they just confuses things.)

$OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O$

The reaction gets rid of some of the hydrogen ions and some of the hydroxides.

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Neutralization

If the number of moles of OH^- = number of moles of H^+ then the solution is completely neutral.

$OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O$

If you mix together a solution that has 0.01 moles HCl and 0.02 moles NaOH, would you expect the solution to have a neutral pH? No

Would you expect it to be basic or acidic? Basic-All of the H^+ will have been removed from the solution but OH^- will remain.

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- Though the Arrhenius understanding of acids and bases explained a lot of things, such as neutralization– It did not explain others.
- To explain such discrepancies a better explanation of acids and bases was needed.

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- Arrhenius description of acids and bases explanation though useful, is very limited. It has a few significant problems.

We now know that H^+ ion is so reactive it does not exist as such in aqueous solutions. Instead H^+ reacts with water to give hydronium, H_3O^+ ion.

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Another Problem with the Arrhenius Definitions

There are many bases that do not contain OH^- ions. These react with water to produce OH^- ions.

For example the base NH_3 itself does not contain OH^- ion but reacts with water to produce OH^- ion.

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

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Brϕnsted - Lowry concept of Acids and Bases

Bronsted-Lowry Acids:

Any substance that is able to give a hydrogen ion, H^+, to another ion or molecule is referred to as a Bronsted-Lowry Acid.

H^+ ions are actually *protons*. Therefore, acids are those substances that can donate protons.

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Bases

Bronsted-Lowry bases:

Any substance that is able to accept a hydrogen ion (H^+) from an acid is referred to as a **Bronsted-Lowry base**.

A base can be neutral or negatively charged, for example, ammonia, NH_3 and hydroxide ion, OH^-

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Applying the Brφnsted – Lowry Concept of Acids and Bases

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

The HCl donates the H^+ it is the acid
The H_2O acceptes the H^+ is is the base.

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Under this definition of acids the following tend behave as acids.

Some Common acids

H_2SO_4	Sulfuric Acid
HCl	Hydrochloric Acid
HNO_3	Nitric Acid
H_3PO_4	Phosphoric Acid
CH_3COOH	Acetic Acid

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Changing from Arrhenius to B-L definition expands the list of common bases

Some common Bases

NaOH	Sodium hydroxide	(Caustic soda)
$Ca(OH)_2$	Calcium hydroxide	(Lime)
$Mg(OH)_2$	Magnesium hydroxide	(Milk of magnesia)
NH_3	Ammonia	
$NaCH_3COO$	Sodium Acetate	

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Acids differ in their relative ability to give up a proton. Similarly, bases differ in their ability to accept a proton. Acids and bases are classified as follows:

Strong acid: An acid that gives up H^+ easily and is essentially **100% dissociated** (splits to produce H^+ and an anion) in water.

Weak acid: An acid that gives up H^+ with difficulty and is **less than 100%** dissociated in water.

Strong base: A base that has a high affinity for H^+ and holds it tightly.

Weak base: A base that has only a slight affinity for H^+ and holds it weakly.

The stronger the acid, the weaker its conjugate base; the weaker the acid, the stronger its conjugate base.

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Background Info:

Telling Reactants from Products

- So far we have seen chemical equations written with a single arrow pointing forward from reactants to products.

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$$

- But often this not quite right. Sometimes it is obvious that products are reacting with each other to reform the reactants.
- Double arrows are used when the reaction appears to happen in both directions.

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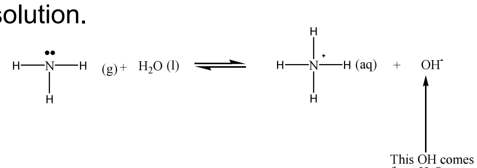
Background Info:

Double Arrows

- When ammonia reacts with water all of the ammonia does not become the ammonium ion. No matter how long a water/ammonia solution sits there is still be ammonia in solution.



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Background Info:

Why the reaction "stops"

- The number of ammonia molecules becoming ammonium ions = the number of ammonium ions becoming ammonia molecules. This means the number of each does not change.
- When an escalator has equal numbers of people going up and down. There is no change in the number of people on each floor.

Background Info:

Using Double Arrows

- Using Double arrows just indicates that we cannot treat the reaction as though only products were present once the reaction occurs.
- How much product and how much reactant are present depends on properties of the reactants and products.
- In reactions between acids and bases double arrows are common.

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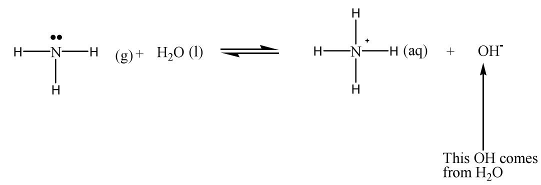
Forward and Reverse Acids and Bases



If an acid-base reaction occurs in both directions, by the B-L model the base on the left hand side of the equation becomes an acid once it gets a proton.

In the equation above ammonia is a base and ammonium an acid.

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Weak Acids and Bases

- Many of the acids listed in the list of acids dissociate completely in water and it is appropriate to only use a single arrow when you describe this or the interaction of these acids with bases.
- Similarly many of the hydroxides that were bases under Arrhenius's definition will be completely protinated under most circumstances.
- However, most acids are not strong acids and most bases are not strong bases, when the L-B definition of acids and bases are used.

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Weak Acids and Bases

- When we start considering solutions that contain weak acids and bases, we need to consider both the acidic and basic form of a molecule.
- The acidic and basic forms of a molecule are called acid/base conjugate pairs.
- When ammonia reacts with water ammonia and ammonium ion are a conjugate acid/base pair.

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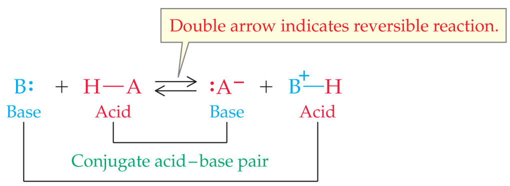
Conjugate Acid Base Pairs



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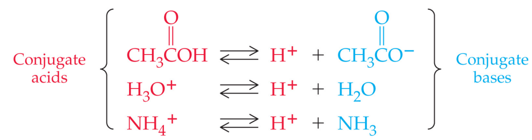
Conjugate Acid Base Pairs



- This image is simplified because it only show the acids becoming bases and leaves off the complimentary bases that must be there.

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Buffers

- Buffers are solutions that are able to keep a nearly constant proton concentration, when an acid or base is added to the solution.
- Buffers are made when a solution is prepared that has nearly equal amounts of an acid and its conjugate base. The acid and its conjugate base are present in amounts much greater than the amount of acid or base they are buffering against.
- When an acid is added to a buffered solution it reacts with the weak base, which can then react with the weak acid. The pH is determine by the balance of forward and back reaction of the weak acid/weak base pair.

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Buffer Solutions

- Buffers are commonly made with acetic acid and sodium acetate.
- CH_3COOH and $NaCH_3COO$
- For example, a solution that contains 0.10 M acetic acid and 0.10 M acetate ion can be utilized as a buffer.
- If a small amount of OH- is added to the above buffer solution, the pH will increase, but not very much because acetic acid of the buffer mixture will neutralize the added OH-.

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Buffers are incredibly important to organisms

- Most of the chemical reactions that our bodies use require a fairly level proton concentration.

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Buffers in the Body

The pH of body fluid is maintained by the following three major buffer systems:

Carbonic acid – bicarbonate system

Dihydrogen phosphate – hydrogen phosphate system

Proteins acting as either proton acceptor or proton donors at different pH values.

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Acidity of a Solution

- Many slides ago we considered how the acidic and basic properties a solution made with water depends on the concentration of hydrogen ions and hydroxide ions in the solution.
- Acidity increases as the hydrogen ion concentration increases, and decreases when the hydrogen ion concentration decreases.
- In basic solutions the hydrogen ion concentration is quite small. (why?)
- The pH scale is a scale that quantifies the acidity of a solution. It is based on the hydrogen ion concentration in a solution.

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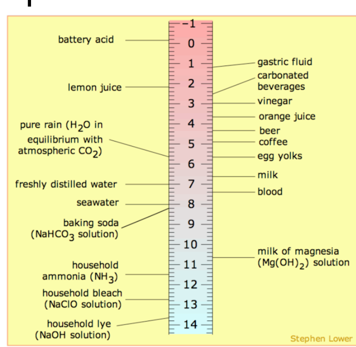
pH Scale – A Measure of Acidity



pH indicates whether a solution is acidic or basic

pH < 7 acidic

pH > 7 basic

pH=7 neutral

pH indicates how basic or acidic the solution is

↑pH - more basic

↓pH -more acidic

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pH definition

- pH is defined in terms of the H+ ion concentration in the solution.
- pH=-log$[H^+]$
- $[H^+]$ is the H^+ concentration in M
- This relation means that as $[H^+]$ increases pH decreases (There is that negative sign in front of the log.)
- As $[H^+]$ decreases pH increases.

pH definition-2

- pH=-log$[H^+]$
- When a concentration is written in scientific notation the pH approximates the value of the exponent.
- For example:
  - $[H^+] = 1.0 \times 10^{-9}$M pH=9
  - $[H^+] = 1.0 \times 10^{-2}$M pH = 2
  - [H+] = 2.45 x 10^{-3M} pH = 2.6

<u>Laboratory Determination of Acidity</u>

The simplest but least accurate way to measure the pH of a solution is to use an acid-base indicator that changes color depending on the pH of the solution.

For example:

- **Litmus** is red below pH 4.8 but blue above pH 7.8.
- **Phenolphthalein** is colorless below pH 8.2 but red above pH 10;
- A mixture of indicators known as **universal indicator** gives approximate pH measurements in the range 2-10 pH.

pH indicators-Red Cabbage Juice

pH indicators act as acids or bases

- pH indicators are dyes that change color when they gain or lose an electron
- The color of a molecule depends on its electronic structure.
- When a molecule loses a proton it is almost as though it gains a lone pair.
- When the proton is returned its old color returns.

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pH indicators for labs, urine, medical labs, pools

pH meters

- For more accurate pH measurements, an **electronic pH** meter can be used.
- pH meters have and electrode that is submerged in the solution

Seagraves

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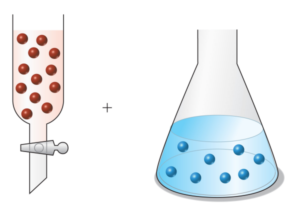
Neutralization

At times neutralization I very carefully in order to better understand he properties of a solution.

•Determine how much of an acid or base are present in the solution
•Determine properties of the acid or base in the solution.
•In the old days it was used to identify acids, in particular amino acids.



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Titration

- Acid-Base titrations use the neutralization reaction to learn about solutions that contain acids or bases.
- You will be doing an acid-base titration in lab when you come back from break.
- Sometimes in titrations indicators are used to monitor the reaction, as are pH meters.

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Summary

- Introduction of Acids and Bases
- Qualitative – (sour, etc.)
- Arrhenius definition–give H+ or OH-
- Bronsted-Lowry definition
- Weak and Strong Acids and Bases
- Conjugate acid/base pairs
- Buffers
- pH
- Titrations

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