Acids

There are two types of acids:
1. Binary Acid
2. Oxyacid

Binary acids are acids that are composed of hydrogen and one other element alone. An example of this is H2S.
When naming binary acids, the following steps are used:

1. The name begins with hydro-
2. The root of the non-hydrogen ion is used
3. The name ends with -ic
4. The name ends with acid

So, the name for H2S would be:
Hydro - sulfur - ic - acid
Hydrosulfuric acid

Binary Acids

- Acids which contain H and another non-metallic element
- Naming -- to the root name of the non-metallic element:
  - add the prefix *hydro-*
  - add suffix *-ic acid*
    - $HF_{(aq)}$ hydrofluoric acid
    - $HBr_{(aq)}$ hydrobromic acid
    - $HCl_{(aq)}$ hydrochloric acid

Note!

Naming Oxyacids

An oxyacid is comprised of hydrogen and a polyatomic ion (also has oxygen in it). An example of this is H2SO4. When naming oxyacids, the following steps are used:

1. Take the root of the polyatomic ion
2. Add the suffix -ic if the polyatomic ion is the -ate version (such as sulfate, SO4) or add the suffix -ous if the polyatomic ion is the -ite version (such as sulfite, SO3)
3. End with acid

So, H2SO4 would be:
Sulfur - ic - acid
Sulfuric acid

Naming Oxyacids

The names of oxyacids are related to the names of the corresponding oxyanions:

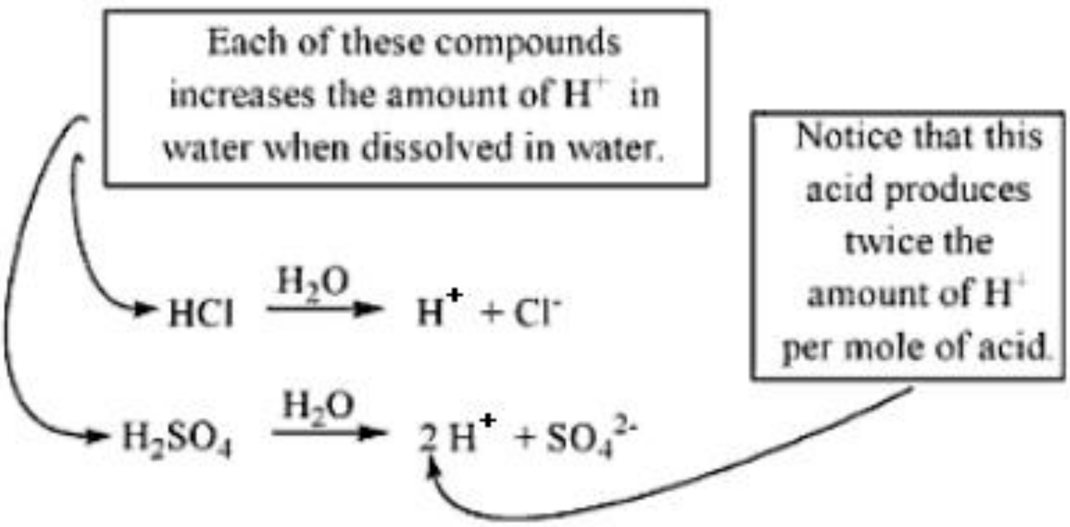
SO_4^{2-} sulf*ate*	H_2SO_4 (*aq*) $\Rightarrow$ sulfur*ic acid*
SO_3^{2-} sulf*ite*	H_2SO_3 (*aq*) $\Rightarrow$ sulfur*ous acid*
NO_3^- nitr*ate*	HNO_3 (*aq*) $\Rightarrow$ nitr*ic acid*
NO_2^- nitr*ite*	HNO_2 (*aq*) $\Rightarrow$ nitr*ous acid*

Defining Acids and Bases

What makes an acid an acid and a base a base? Well, according to different scientists in different time periods, there are several determining factors

Arrhenius Definition

According to Arrhenius's definition, an acid dissociates i water to make H+ ions, and a bas dissociates in water to make OH- ions.



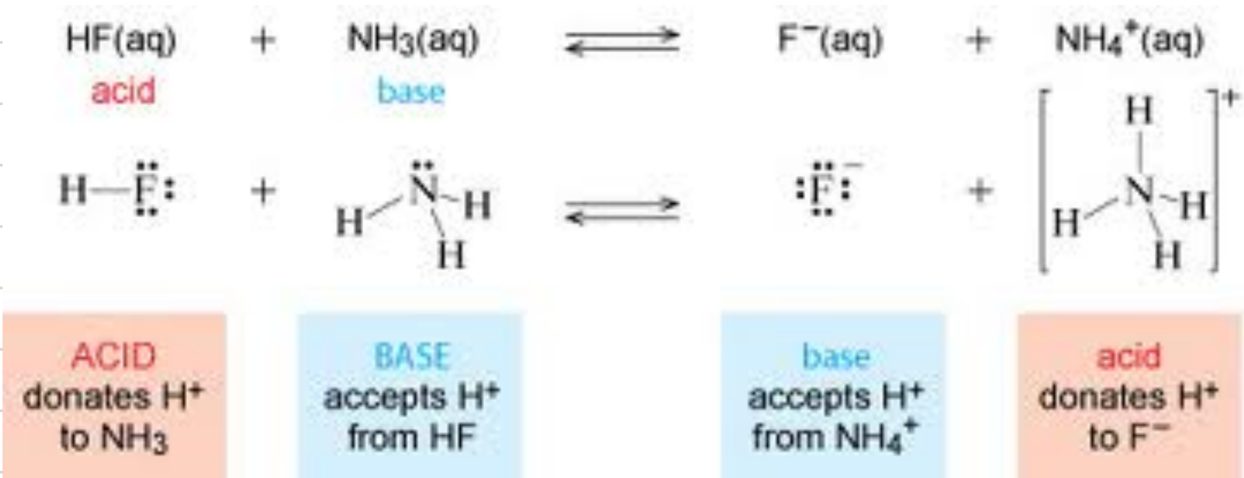
To identify Arrhenius acids look for compounds that dissociate or ionize in water forming H^+

Bronsted/Lowry Definition

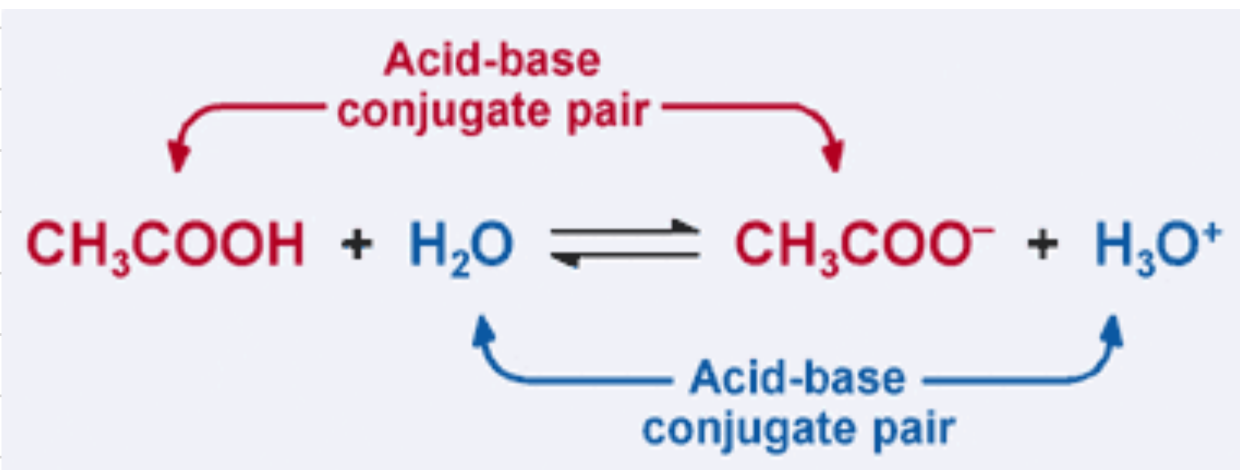
According to the Bronsted/Lowry definition, an acid is any compound that donates an H+ ion, and a base is any compound that accepts an H+ ion.

This definition is more versatile because it eliminates the need of OH- as well as the need for the compounds to be in water. This definition therefore expands the definition of bases.

Keep in mind that H+ is simply a proton (1 proton, no electrons).



Since these reactions are reversible, it can be confusing telling acid/base is being talked about and whether you are talking about the forward or reverse reaction. This is where conjugate acid/base pairs come into play



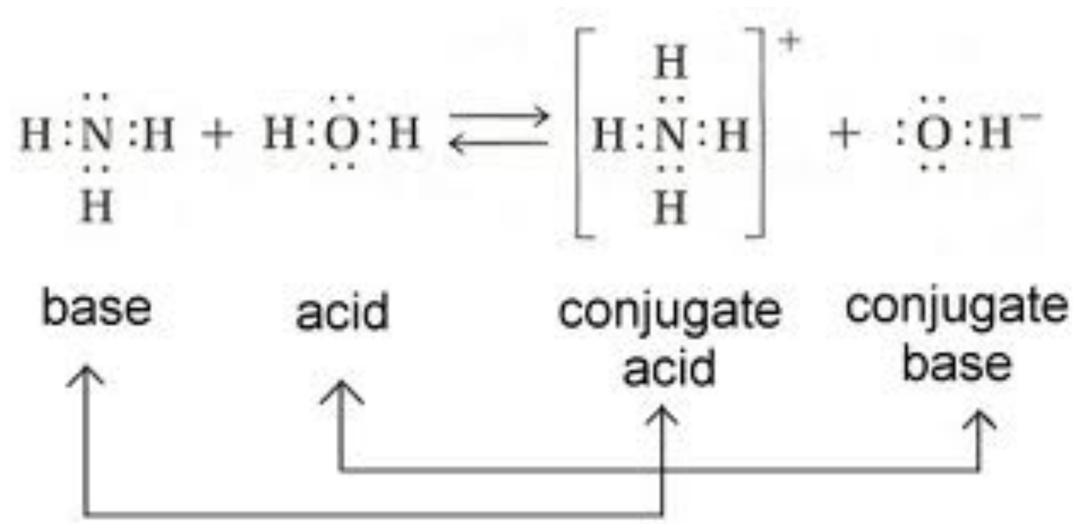
Conjugate Acid/Base Pairs

Amphoteric - This is a term used to describe something that can act as either an acid or a base (an example of this is H2O. It can give up a proton to become OH- or accept a proton to become H3O+)

To distinguish between reactants and products in acid/base reactions, products are renamed based on their behavior.

Since acids donate an H+ ion in the forward reaction, in the reverse reaction they will accept an H+ ion. Therefore, in the forward reaction, it is an acid, and in the reverse, it is a conjugate base.

The same can be said for bases. They accept H+ ions in the forward reactions, so they give up H+ ion in the reverse reactions. Therefore, in the forward reaction, they are bases, but in the reverse, they are conjugate acids.



Weak Acid Dissociation

Just like equilibrium constants, acid dissociation can be measured using a Ka (acid dissociation constant). This value measures the relative strength of an acid.

$$HA + H_2O \rightleftarrows H_3O^+ + A^-$$

A = generic acid

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

This is the generic layout for any acid reaction to find the relative strength of the acid.

When Ka < 1, the acid is weak
When Ka > 1, the acid is strong.

Essentially, the Ka is:

$$Ka = \frac{\text{Concentration of products}}{\text{Concentration of acid}}$$

When looking at the chemical equation of the reaction, if the products coefficients are 1 to 1, then the concentration of the products will be equal.
If the concentration of the acid is given at the beginning, subtract the concentration of one of the products to find the end concentration of the acid.

Weak Base Dissociation

Like weak acid dissociation, there is a generic weak base constant, Kb:

$$B + H_2O \rightleftarrows HB^+ + OH^-$$

B = generic base

$$K_b = \frac{[HB^+][OH^-]}{[B]}$$

With weak acids and bases problems, when you are given the Ka/b and are told to find x, since the dissociation is very small, [A/B] - [x] ~~ [A/B].

This is because weak acids do not dissociate much at all, so not much product is made, meaning not much reactant is used.

Self Ionization of Water

Water as a solute is constantly dissociating and associating.

$$2H_2O \rightleftharpoons H_3O^+ + OH$$

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

Kw is a constant used to determine the acidity of a solution.

Pure water has a Kw of 1.0

If a solution's Kw > 1, then it is acidic

If a solution's Kw < 1, then it is basic.

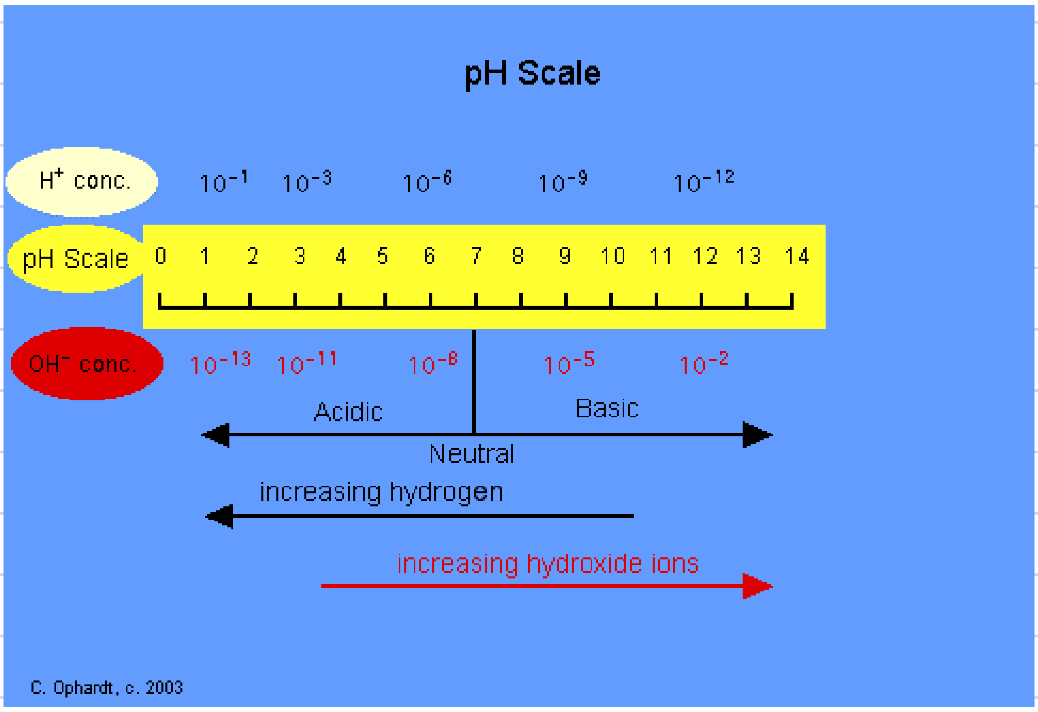
Measuring H30+ concentration

There are two ways to measure the concentration of H30+ in a solution:

1. Indicators - a chemical that changes color when it gains or loses H+

2. pH meter - an instrument that is sensitive to H30+

Using [H30+], we can calculate the pH of a solution.



pH = - log [H30+]

[H30+] is in moles.

As far as the pH scale is concerned:
0-3 are strong acids
4-6 are weak acids
7 is neutral
8-10 are weak bases
11-14 are strong bases

Lewis Concept of Acids and Bases

Lewis Acid - forms a covalent bond by accepting electron pairs from another species

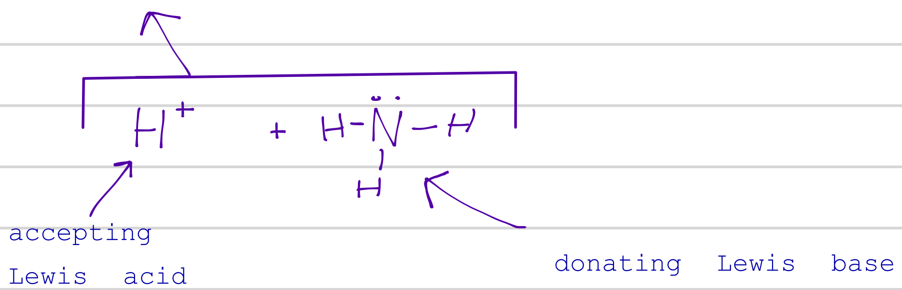
Lewis Base - forms a covalent bond by donating electron pairs to another species

Electron dot notations are very important for this method.

Strategy for Lewis Problems

1. Make electron dot models
2. Use location of electrons to determine the donor and aceptor

$$NH_4^+ + Cl^- \rightarrow NH_4Cl$$



Determining Acid/Base Strength

Concentrated & Dilute deals with the amount of water and acid in a solution.

Strong - dissociates almost completely in water
Weak - does not readily dissociate in water

Acid - donates H+ to produce H3O+
Base - accepts H+ to produce OH-

There are 4 general rules for determining the strength of an acid or base; 2 rules for binary acids and 2 rules for oxyacids.

1. For Binary acids, as the atom size increases the strength of the acid increases
*This is for elements in the same group
ex) Following acids are in order from weakest to strongest: HF, HCl, HI

2. For Binary acids, as electronegativity increases, acid strength increases.
ex) H2O vs. HF --> HF is stronger

Binary Acids

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Note!

Oxyacid Rules

3. For oxyacids with the same ratio of O to another atom Y (not H), the strength increases with the electronegativity of Y.
ex) HClO vs. HBrO --> HClO is stronger

4. For oxyacids, strength increases with the number of oxygens.

There are also 2 factors for Acid Strength:

1. Polarity of the bon
* the more polar the bond, the more easily the proton is removed and the greater the acid strength

2. Strength of the bond
* Bigger atoms hold protons weaker, making stronger acids (They break apart and dissociate more readily)

Neutralization

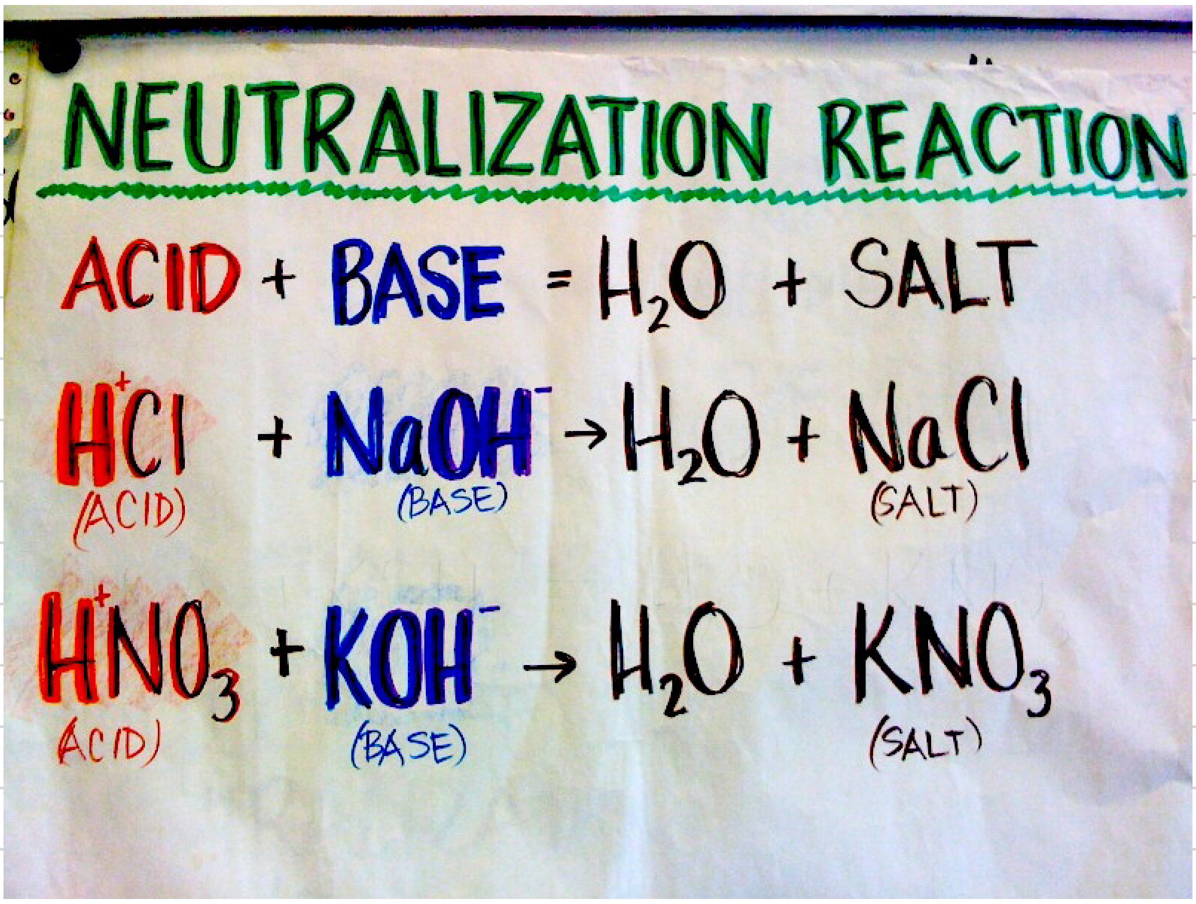
Neutralization - complete reaction of an acid and a base.

Antacids - contain a basic chemical that can neutralize stomach acid.

Neutralization always creates water and salt as a product.

$$HCl + NaOH \rightarrow \boxed{H_2O + NaCl}$$

salt - compound made from a negative ion from an acid and a positive ion from a base.
Common examples:
CaCl, NaHCo3, NaNo3, MgSo4



Redox Reactions

Redox is short for Reduction and Oxidation
Reduction - a reaction where an element gains electrons

$O + 2e^- \rightarrow O^{-2}$ Reduction half reaction

Oxidation - a reaction where an element loses electrons
*When an element is oxidized, it picks up a (+) charge

$Zn \rightarrow Zn^{+2} + 2e^-$ Oxidation Half Reaction

These definitions can be remember with the mnemonic device OIL RIG:

Oxidation
Is
Loss (of electrons)

Reduction
Is
Gain (of electrons)

Oxidation number - charge an atom would have if electrons in each bond belonged to the more electronegative element.