Acid/Base Definitions & Theories

Arrhenius Acids and Bases

Arrhenius Acids – produce **H⁺** and **an anion** in solution

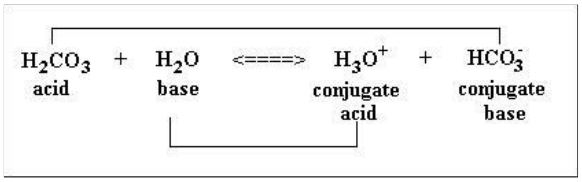
Ex: $HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$

Arrhenius Bases – produce **a cation** and **OH-** in solution

Ex: $NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$

Bronsted-Lowry Acids and Bases

Bronsted-Lowry Acid – is a proton (H+) **donor**
Bronsted-Lowry Base – is a proton (H+) **acceptor**



pH, pOH, [H+], [OH-]

- pH is a measure of how much H⁺ is in the solution, or how acidic the solution is

pH	pOH
[H+]	[OH-]

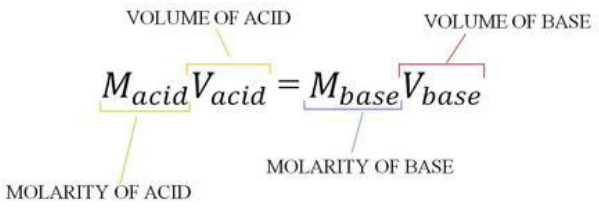
Blue Litmus paper **tuns red** in the presence of **an acid**.
Blue Litmus paper **stays blue** in the presence of **a base**.

Red Litmus paper **stays red** in the presence of **an acid**.
Red Litmus paper **turns blue** in the presence of **a base**.

$pH = -\log [H^+]$ **$10^{-pH} = [H^+]$** **$pOH = -\log [OH^-]$** **$10^{-pOH} = [OH^-]$**

An acid-base titration is a lab technique that allows you to determine the concentration of an unknown solution. (There are other types of titrations, but this is the most common.) Some terms you need to know:

- **Titrant**: a solution of known concentration (usually); usually the solution in the buret
- **Analyte**: the solution you are trying to determine the concentration of; usually the solution in the beaker or flask
- **Equivalence point:** the volume of titrant added to give **equal moles of acid and base** (in an acid/base titration)
- **End point:** the volume of titrant added to make the color of the indicator change
  **hopefully, the equivalence point and the end point happen at the same time!
- **Indicator:** a solution that changes color in varying pH ranges



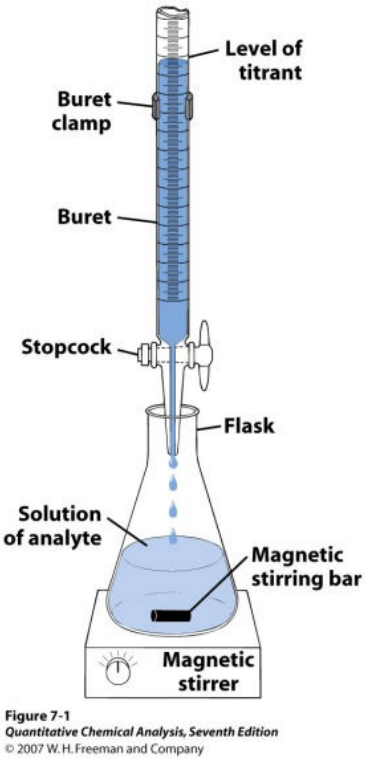


Figure 7-1
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Note: If you are using a diprotic or triprotic acid, then you need to add coefficients to the equation above. (C_a = coefficient of acid; C_b = coefficient of base. Coefficients are based on the # of moles required for a balanced chemical equation.)

$M_a C_b V_a = M_b C_a V_b$