Honors Chemistry
Acid/Base Summary Information:

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General equation for neutralization: Acid + Base → Salt + Water
Ex.

$HCl + NaOH \rightarrow NaCl + H_2O$

Acid-Base Characteristics

Acids:

- Turn litmus paper from blue to red
- React with active metals (e.g. $2HCl + Mg \rightarrow MgCl_2 + H_2$)
- Taste Sour
- React with bases to form water and salt

Bases:

- Turn litmus paper from red to blue
- Feel slippery or soapy on the skin
- Taste bitter
- React with acids to form water and salts.

Arrhenius Definition:

Acid - Produces H^+ ions in solution $HCl \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$

Base - Produces OH^- ions in solution $NaOH \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$

Dissociation = Is related to the ability of a substance to "break apart" in solution.

Bronsted-Lowry Definition:

Proton transfer direction (HCl → H_2O)

Acid - Proton donor: $HCl + H_2O \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

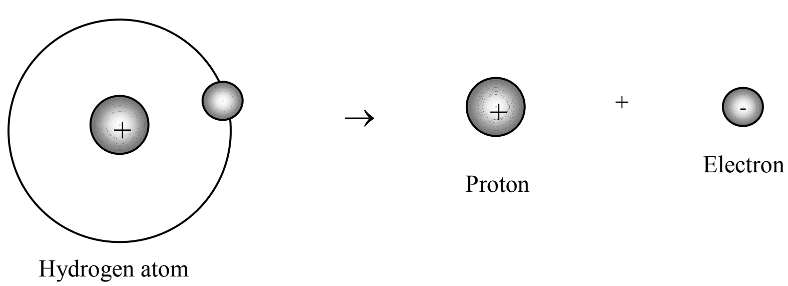
HCl: Proton donor; H_2O: Proton acceptor; H_3O^+: Hydronium

Proton transfer direction (H_2O → NH_3)

Base - Proton acceptor: $NH_3 + H_2O \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$

NH_3 (ammonia): Proton acceptor; H_2O: Proton donor; OH^-: Hydroxide

Consider a hydrogen atom which consists of one proton and one electron. If the electron is separated from the nucleus, you are left with a proton. This is the "proton" referred to in the Bronsted-Lowry definition



Usually H^+ and H_3O^+ (hydronium) refer to the same thing. The hydronium form shows the hydrogen ion's association with the water molecule in the solution.

The strength of an acid or base is dependent upon its ability to dissociate. You should be able to differentiate between
Strong vs. weak (based on the ability of the substance to dissociate) and
Concentrated vs. dilute (based on the concentration (e.g. molarity) of the solute in solution.

A strong acid (such as HCl) dissociates completely, as in shown in the right facing arrow.
$HCl \longrightarrow H^+ + Cl^-$

A weak acid (such as acetic acid, CH_3COOH) only dissociates to a very small degree
$CH_3COOH \leftrightarrow CH_3COO^- + H^+$
The two arrows show that the bulk of the molecules remain undissociated CH_3COOH
Make sure you are familiar with the 6 common strong acids
HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_4$
and the 8 common strong bases
LiOH, NaOH, KOH, RbOH, CsOH (hydroxides of the alkali metals) and
$Ca(OH)_2$, $Sr(OH)_2$ and $Ba(OH)_2$ (hydroxides of some of the alkaline earth metals. Note: Although these are considered strong bases, their solubility is actually limited)

You should also know the weak base, ammonia (NH_3) and the weak acid, acetic acid (CH_3COOH). These are a little tricky because NH_3 doesn't have OH^- in it. It produces OH^- by the way that it reacts with water

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

In addition CH_3COOH doesn't look like an acid because it is commonly written without a "leading hydrogen" in the formula which most acids have and it also looks like it has an OH at the end. In reality, only the H is removed from the molecule

$CH_3COOH \rightarrow CH_3COO^- + H^+$

Conjugate acid-base pairs (From Bronsted-Lowry)

In an acid/base reaction, the acid identified on the reactant side becomes the conjugate base on the product side (because it would be the proton acceptor in the reverse direction). The converse is true for bases (i.e. conjugate acid/base pairs differ by a hydrogen (proton))

Ex:

HF	+	NH_3	$\leftrightarrow$	F^-	+	NH_4^+
acid		base		conj base		conj acid

Ex:

H_2SO_4	+	H_2O	$\leftrightarrow$	H_3O^+	+	HSO_4^-
acid		base		conj acid		conj base

Note: The stronger the acid or base, the weaker its conjugate is.
Equilibrium lies on the side of the dissociation of the stronger counterpart.

Amphiprotic means able to act as an acid of a base [i.e. can gain or lose a proton]) and *amphoteric* means able to react with an acid or a base.
Ex:

$HPO_4^{2-} \rightarrow H^+ + PO_4^{3-}$ (here HPO_4^{2-} acts as an acid)

HPO_4^{2-} + H^+ → $H_2PO_4^-$ (here HPO_4^{2-} acts as a base)
Water is amphoteric and amphiprotic

The degree of dissociation of a weak acid or base is given by the ionization constant.

K_a = acid dissociation constant

$$K_a = \frac{[H^+]\,[\text{anion}]}{[\text{Undissociated acid}]}$$

Example: $[H^+][F^-]/[HF]$

Similarly,

$$K_b = \frac{[\text{cation}]\,[OH^-]}{[\text{Undissociated Base}]}$$

In water acids produce H_3O^+. Acids stronger than hydronium completely dissociate to hydronium ions in aqueous solutions. Since ionization is complete there is no ionization constant possible, therefore it is impossible to determine the relative strengths of these acids. This is called the *leveling effect*.
Ex:

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

To determine relative strengths for these stronger acids a solvent that is a weaker base than water (e.g. diethyl ether or acetone) must be used so that dissociation is not 100%.

pH
pH = The "power" of hydrogen
Related to the molarity of H^+ (H_3O^+) in solution
designated by $[H^+]$ (brackets mean molarity)

Definition: pH = -log of the hydronium ion concentration

$pH = -\log[H^+]$ $\qquad$ $[H^+] = 10^{-pH}$

$pOH = -\log[OH^-]$ $\qquad$ $[OH^-] = 10^{-pOH}$

Ex. Calculate the pH of a .003M nitric acid solution.
Answer: Given $HNO_3 \rightarrow H^+_{(aq)} + NO_3^-{}_{(aq)}$ $\quad$ pH = -log (.003) = 2.52

The pH scale

Based on the autoionization of water
$H_2O + H_2O \rightarrow H_3O^+ + OH^-$
Concentrations of both the hydronium and hydroxide are $1.0\text{x}10^{-7}$M in neutral water
The product of the hydronium and hydroxide is always $1.0\text{x}10^{-14}$, regardless of how basic or acidic the aqueous solution is. This should tell you that as the hydronium ion concentration goes up, the hydroxide ion concentration goes down (suppressed), and vice versa.

$pH = -\log[1.0\text{x}10^{-7}] = 7$ $\qquad$ which is considered neutral on the pH scale pH scale

<-0---------------7--------------14-->
7 is neutral, <7 is acidic, >7 is basic   (opposite for pOH)

[H+][OH-] = 1.0x10-14
pH + pOH = 14
NOTE: Each pH change of 1 corresponds to a ten-fold increase or decrease in the hydronium concentration.

Sample Question:
Calculate the pH of a solution formed from adding 50.0g of NaOH into enough water to produce 1500mL of solution.

Calculate the molarity of a nitric acid solution with a pOH of 8.3.
pH = 14-8.3 = 5.7

[H+] = 10-5.7 = $2.0\text{x}10^{-6}$M

What is the hydroxide ion concentration of the solution in the first problem?

Equations to remember:

Autoionization constant for water:

$K_w = 1.0\text{x}10^{-14} = [H_3O^+][OH^-]$ Remember brackets mean molarity

The product of $[H_3O^+][OH^-]$ ALWAYS equals $1.0\text{x}10^{-14}$

$pH = -\log[H_3O^+]$ $[H_3O^+] = 10^{-pH}$
$pOH = -\log[OH^-]$ $[OH^-] = 10^{-pOH}$
pH + pOH = 14

$K_aK_b = K_w$ For conjugates

Example:

You are given a .010M solution of HNO_3:

If you are asked,

1. **What is the H^+ ion concentration of the solution?**
   Solution: Since HNO_3 is a strong acid there is 100% dissociation, therefore the H^+ ion concentration is equivalent to the molarity of the acid itself.
   Answer: .010M
2. **What is the pH of the solution?**
   Solution: The pH calculated by pH = -log[.010] = 2
   Answer: pH = 2
3. **What is the hydroxide concentration of the solution?**
   Solution: Since $1.0\text{x}10^{-14} = [H_3O^+][OH^-]$ then $[OH^-] = 1.0\text{x}10^{-14}/.010$
   Answer: $[OH^-] = 1.0\text{x}10^{-12}$
4. **What is the pOH of the solution?**

Solution: There are two ways that this can be calculated

a. $pOH = -\log[OH^-] = -\log[1.0x10^{-12}]$

Answer: 12

Or

b. Knowing pH + pOH = 14 then pOH = 14-2 = 12 (where 2 was the pH calculated in question 2).

Answer: 12

Buffers:

Buffers resist changes in pH from the addition of strong acids and bases. Buffers are essential in many biological system (ex. blood) and are used in industry as well. They keep solutions from being "shocked" by excessive H^+ or OH^-

They are made from a weak acid and a salt of its conjugate base (or a weak base and a salt of its conjugate acid). For example HF and NaF could make a good buffer system (the cation of the salt is not too important as long as it is soluble).

The hydrogen ion of acids would be "consumed" as follows:

$$H^+ + F^- \rightarrow HF$$

The hydroxide of bases would be "consumed" as:

$$OH^- + HF \rightarrow H_2O + F^-$$

The H^+ and OH^- would have come from the addition of an acid or a base to the solution. Remember, as an undissociated molecule HF is not damaging.

Strong acids don't work because their conjugate base is so weak that no acids would be consumed.

The Henderson-Hasselbalch Equation for Buffers

$HA \leftrightarrow H^+ + A^-$

$K_a = [H^+][A^-]/[HA]$

$[H^+] = K_a [HA]/[A^-]$

$pH = pK_a + \log([A^-]/[HA])$ Henderson-Hasselbalch Equation

$pH = pK_a + \log(n_{A-}/n_{HA})$

Buffer Capacity: How much added acid or base can be added before being "exhausted".

Usually .1 to 1M solutions

Choosing the right buffer for a particular pH

Usually best when the ratio of $[A^-]/[HA] = 1$ (this gives the greatest range in protecting the solution against both added acids and bases.)

Determine whether the following could make good buffers.

a. CH_3COOH and $NaCH_3COO$
b. HNO_3 and KNO_3
c. HF and NaCl

Answer: a. yes, weak acid and conjugate base: b. No, strong acids are no good: c. No NaCl is not a conjugate of HF

Terms:

Amphoteric: A substance that can behave as both an acid and a base
$Al(OH)_3 + 3NaOH \rightarrow Na_3AlO_3 + 3H_2O$ (Acid)
$Al(OH)_3 + 3HCl \rightarrow AlCl_3 + 3H_2O$ (Base)

Amphiprotic: A substance that can gain or lost a proton (H^+)
e.g. H_2O
$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$
$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Acidic anhydride: Nonmetal oxide that in water forms an acid
$SO_3 + H_2O \rightarrow H_2SO_4$
Basic anhydride: Metal oxide that in water forms a base
$Na_2O + H_2O \rightarrow 2NaOH$
Anhydride means literally "without water"

Polyprotic: An acid having more than one hydrogen available to dissociate (e.g. H_2SO_4 (**diprotic**), H_3PO_4 (**triprotic**))

Electrolyte: A substance that conducts electricity in water.
Strong acids and bases are good electrolytes.

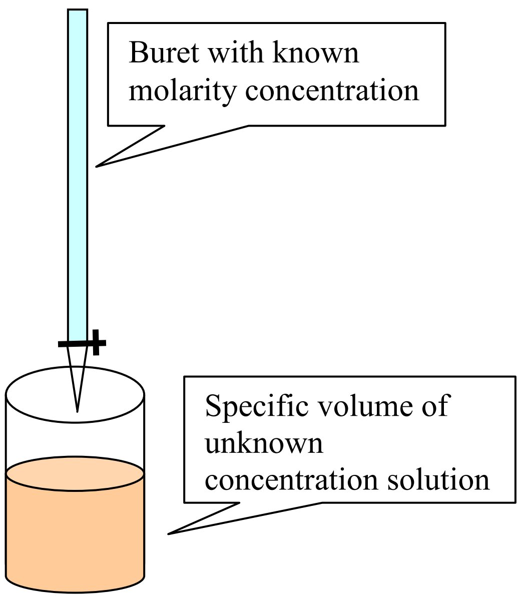
Indicator: A substance used in a titration that changes color depending on the pH of the solution. Normally used to visually indicate the equivalence point of the reaction.

Equivalence point: When equal numbers of hydronium and hydroxide have reacted in a neutralization process.

End point: The specific pH at which an indicator changes color

Titrations:

Titrating a solution is a method by which the unknown concentration (molarity) of a solution is determined by reacting it with a volume of a solution of known concentration (known as a standardized solution or titrant) to the point where the reaction has gone to completion (equivalence point).

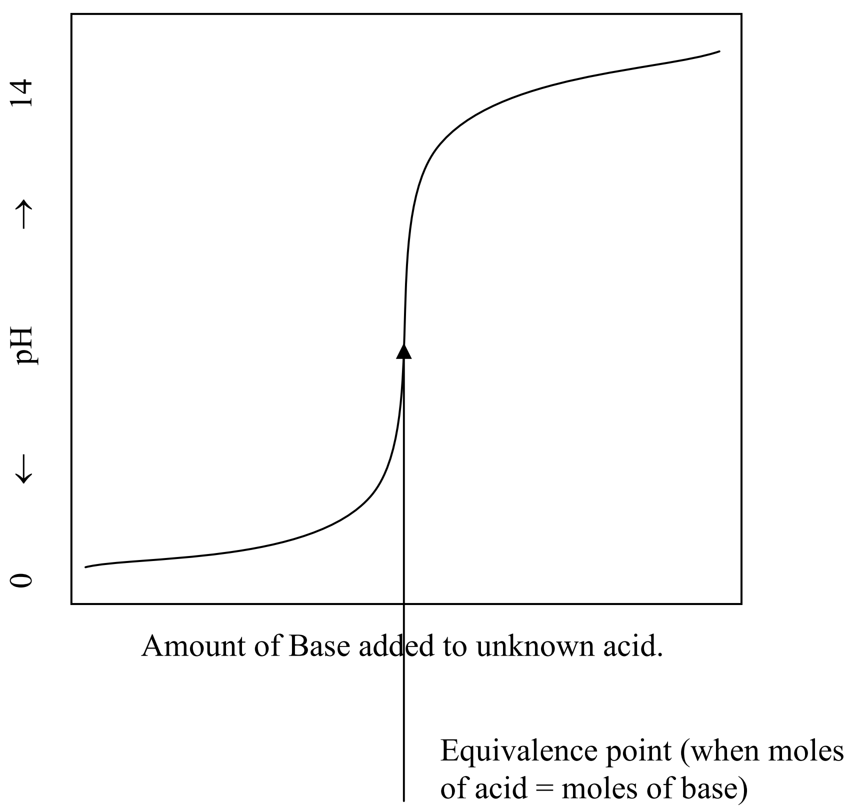


Example:
$HCl + NaOH \rightarrow NaCl + H_2O$

What is the molarity of an unknown NaOH solution if 50mL of a .1M HCl solution is required to neutralize 500mL of the unknown?

1. moles HCl = (.1M)(.050L)=.005mol HCl
2. Since the coefficients are 1:1 you titrated .005mol of NaOH
3. Concentration = (.005mol)/(.500L) = **.01M NaOH**

Titration Curve for strong acid and strong base:



The equivalence point is when equal numbers of moles of acid and base have reacted with one another.

In order to determine the exact equivalence point, a pH indicator (litmus, etc.) or pH meter must be used.

For titration curves involving a weak acid and/or a weak base, the titration curve will be skewed with the pH at the equivalence point (i.e. the salt) having a pH above or below 7 (see your book for details).

On Your Own:

If 250mL of a .25M LiOH solution is required to titrate 125mL of nitric acid (HNO_3)

a. What is the concentration (molarity) of the nitric acid solution?

b. Write a balanced equation for this reaction.