Acids, Bases and Salts

A. Properties

1. Acids

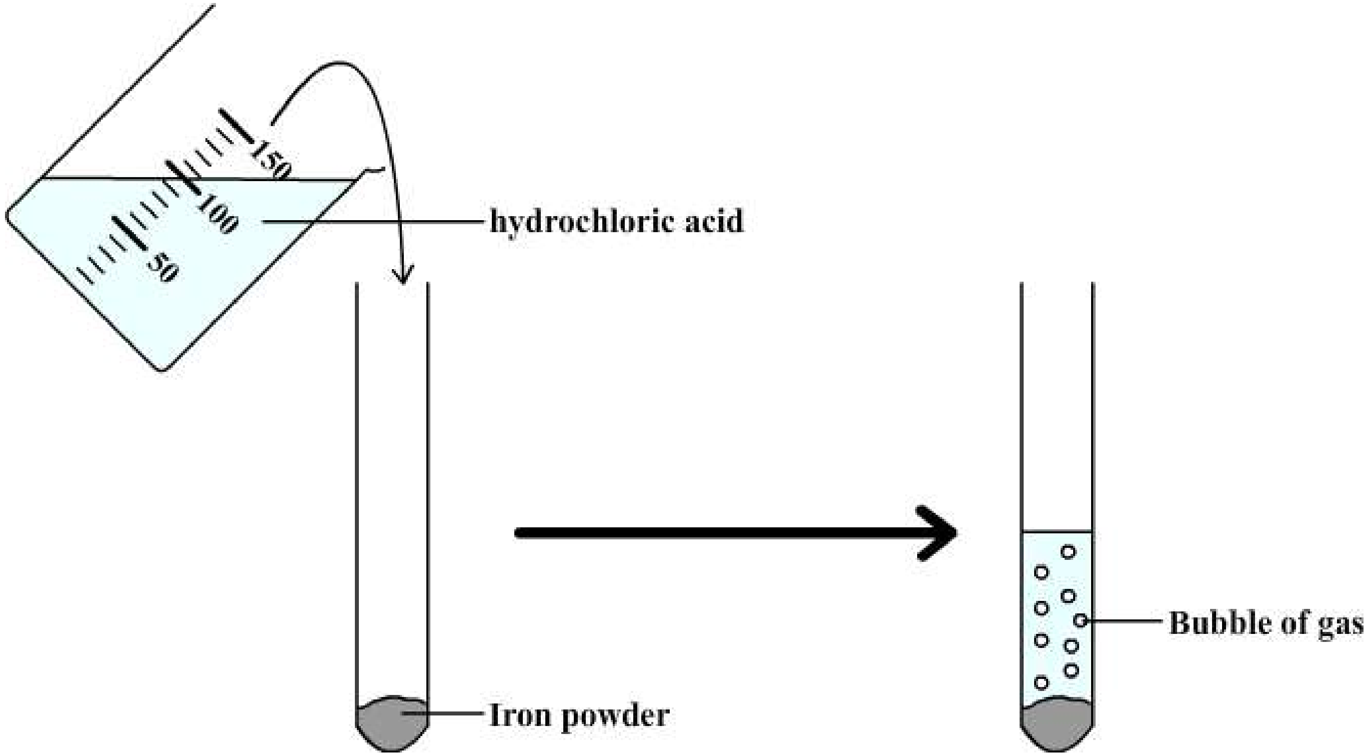
a. Donates H^+

b. Sour taste

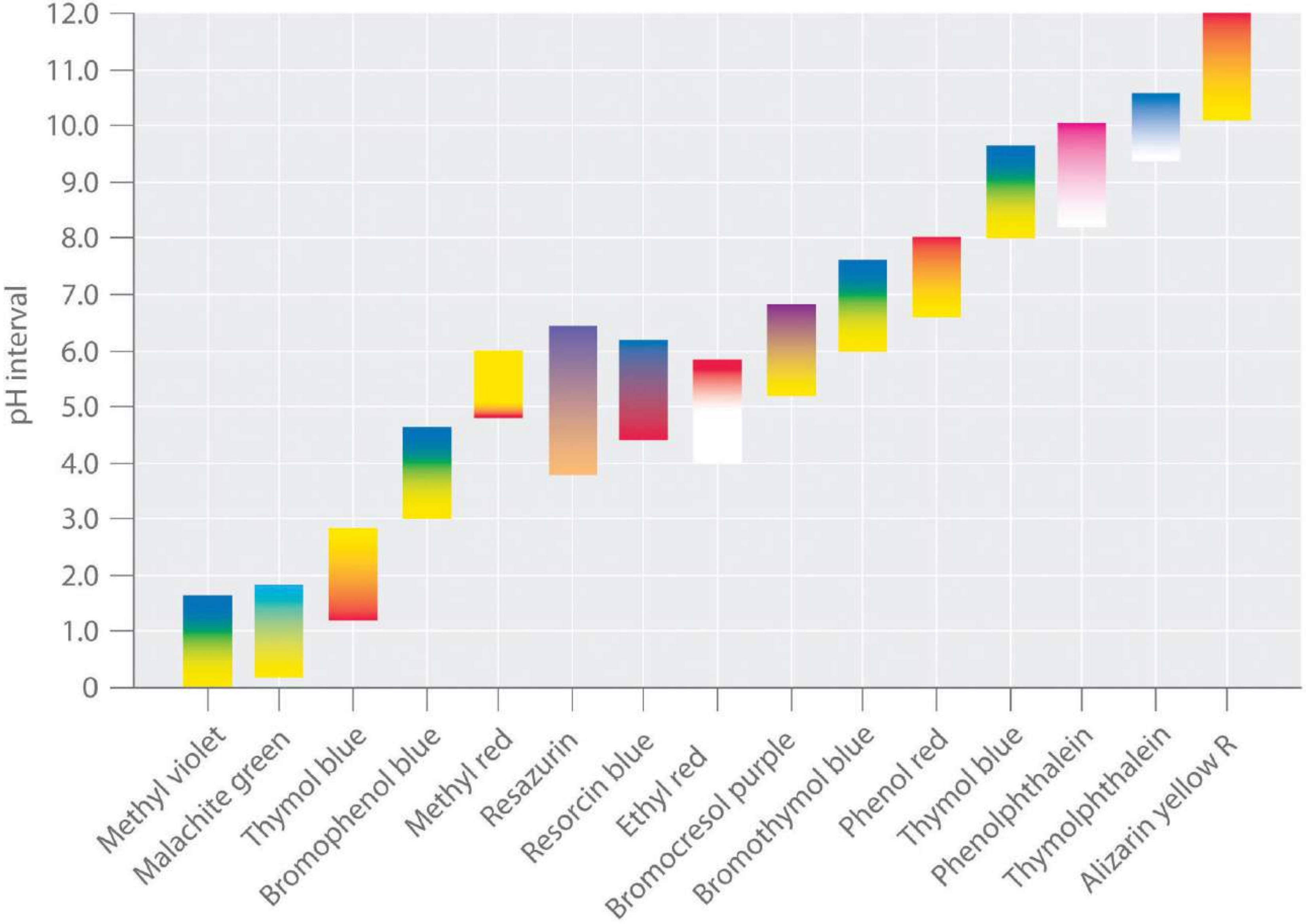
c. Reacts with metals

d. Turns blue litmus red, UI red, phth clear

$$HCl \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$



Indicator	pK_{in}	Color change
Alizarin yellow R	11.0	yellow to red
Thymolphthalein	9.9	colorless to blue
Phenolphthalein	9.5	colorless to pink
Thymol blue	9.2	yellow to blue
Phenol red	7.4	yellow to red
Bromothymol blue	7.3	yellow to blue
Bromocresol purple	6.4	yellow to purple
Ethyl red	5.4	colorless to red
Resorcin blue	5.3	red to blue
Resazurin	5.1	orange to violet
Methyl red	5.0	red to yellow
Bromophenol blue	4.1	yellow to blue
Thymol blue	1.7	red to yellow
Malachite green	1.3	yellow to turquoise
Methyl violet	0.8	yellow to blue



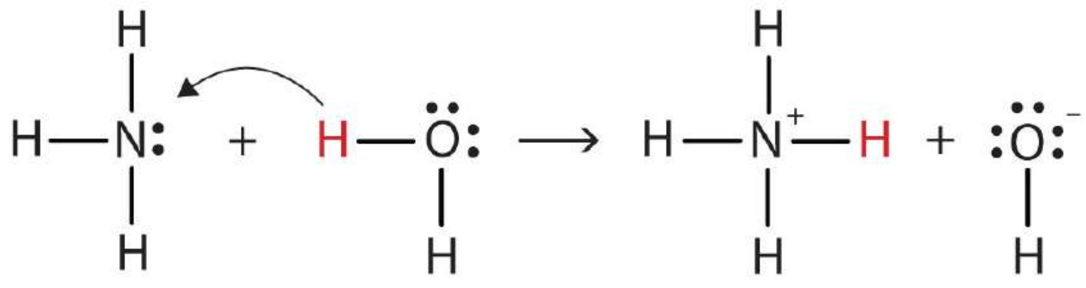
2. Bases

a. Accepts H^+

b. Tastes bitter

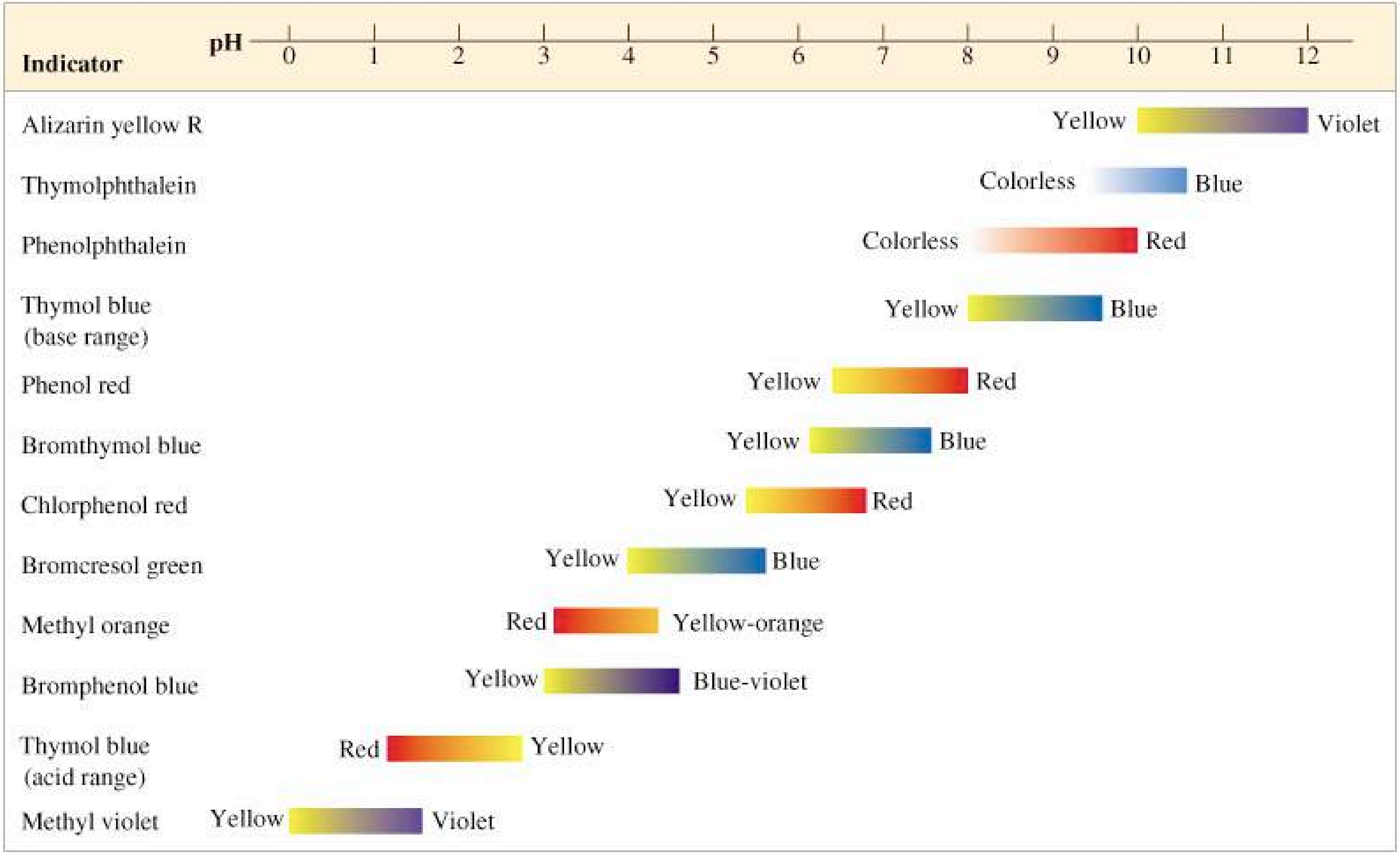
c. Feels slippery

d. Turns red litmus blue, UI purple, phth pink



Hydrogen ion acceptor: B-L base

Hydrogen ion donor: B-L acid


Indicator
pH
0
1
2
3
4
5
6
7
8
9
10
11
12
Alizarin yellow R
Yellow
Violet
Thymolphthalein
Colorless
Blue
Phenolphthalein
Colorless
Red
Thymol blue
(base range)
Yellow
Blue
Phenol red
Yellow
Red
Bromthymol blue
Yellow
Blue
Chlorphenol red
Yellow
Red
Bromcresol green
Yellow
Blue
Methyl orange
Red
Yellow-orange
Bromphenol blue
Yellow
Blue-violet
Thymol blue
(acid range)
Red
Yellow
Methyl violet
Yellow
Violet

B. Definitions

1. Arrhenius (1887) proposed that acids are compounds containing hydrogen that yield / donate that hydrogen (a proton) in solution. Bases are compounds containing hydroxide that yield / donate that hydroxide in solution.

Arrhenius Acids:

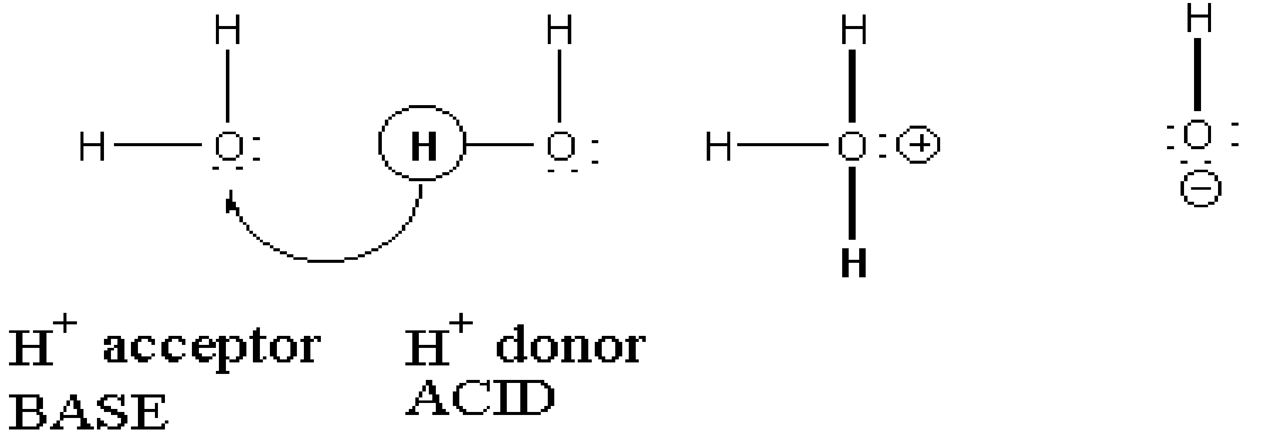
$$HA\,(aq) \rightleftharpoons H^{+}\,(aq) + A^{-}\,(aq)$$

Arrhenius Bases:

$$BOH\,(aq) \rightleftharpoons B^{+}\,(aq) + OH^{-}\,(aq)$$

2. Bronsted-Lowry (1923) broadened the Arrhenius definition. They proposed that acids are compounds that donate hydrogen (protons). This is the same as the Arrhenius definition. But they define bases as compounds that will accept hydrogen ions. There is no mention of hydroxide. Conjugate pairs are created by this giving and acceptance of the hydrogen.

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$



$$\underset{\text{Acid}_1}{HNO_2} + \underset{\text{Base}_2}{H_2O} \rightarrow \underset{\text{Conj Base}_1}{NO_2^-} + \underset{\text{Conj Acid}_2}{H_2O^+}$$

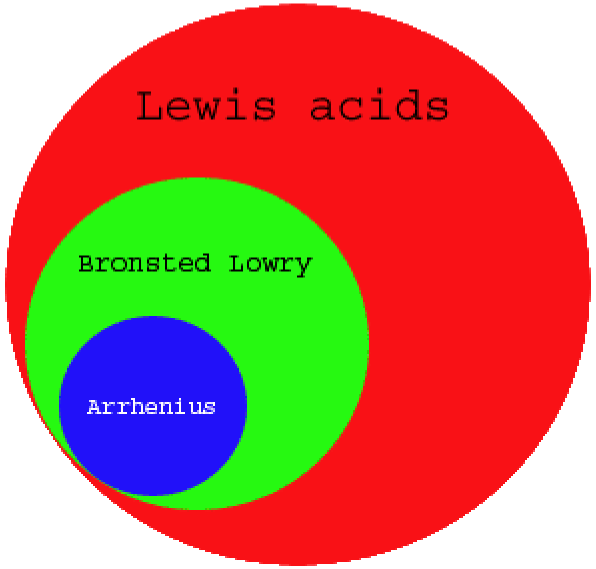
$$\underset{\text{Base}_1}{HPO_4^{2-}} + \underset{\text{Acid}_2}{NH_4^+} \rightarrow \underset{\text{Conj Acid}_1}{H_2PO_4^-} + \underset{\text{Conj Base}_2}{NH_3}$$

Conjugate acid-base pair

$$\underset{\text{Weak acid}}{CH_3COOH} + \underset{\text{Weak base}}{H_2O} \rightleftharpoons \underset{\text{Strong acid}}{H_3O^+} + \underset{\text{Strong base}}{CH_3COO^-}$$

Conjugate acid-base pair

3. Lewis (1930's) broadened the definition even more. He proposed that acids are substances that accept electron pairs, while bases are substances that donate them.

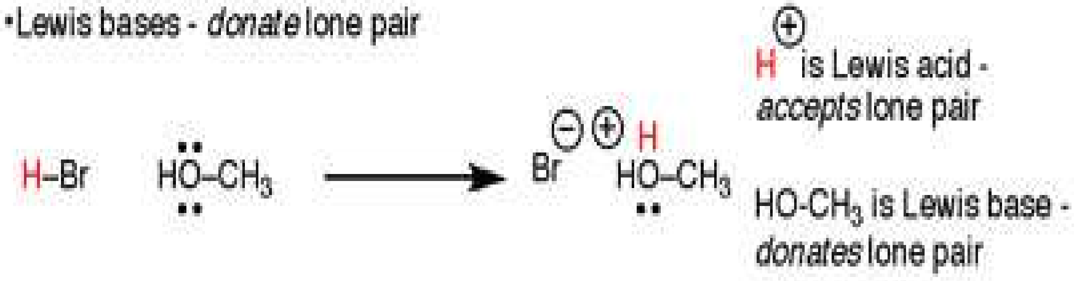


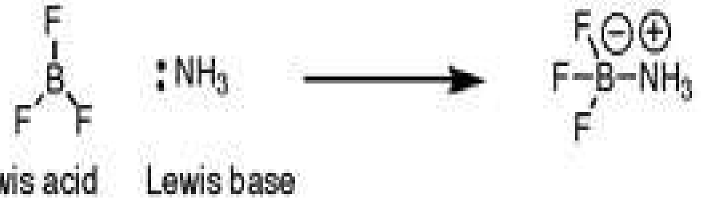
This ammonia molecule has four pairs of electrons around its central atom and only three of the pairs are shared. Therefore, this molecule has another pair of electrons that it can share and that makes it a Lewis base.

This hydroxide ion has four pairs of electrons around its central atom and only one of the pairs is shared. Therefore, this ion has unshared pairs of electrons available to share and that makes it a Lewis base.

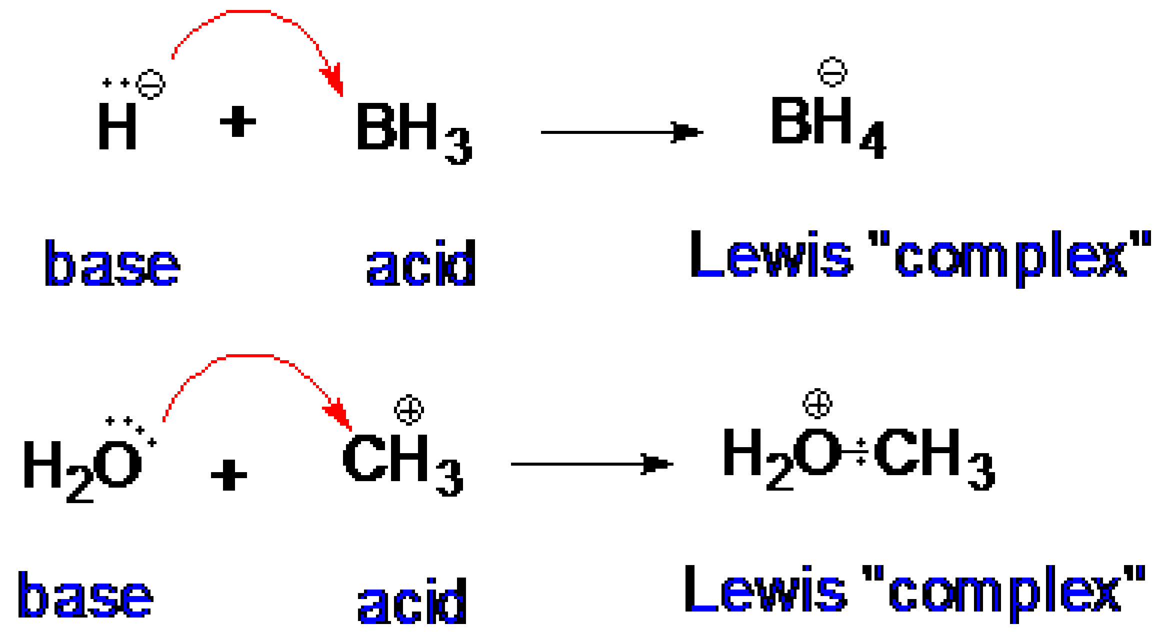
Lewis acids and bases

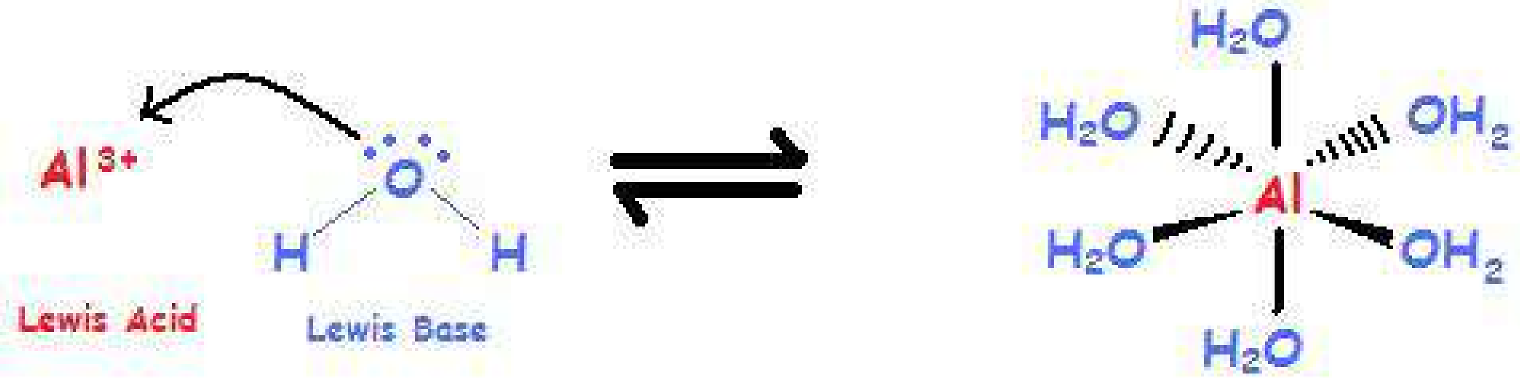
- Lewis acids - *accept* lone pair
- Lewis bases - *donate* lone pair

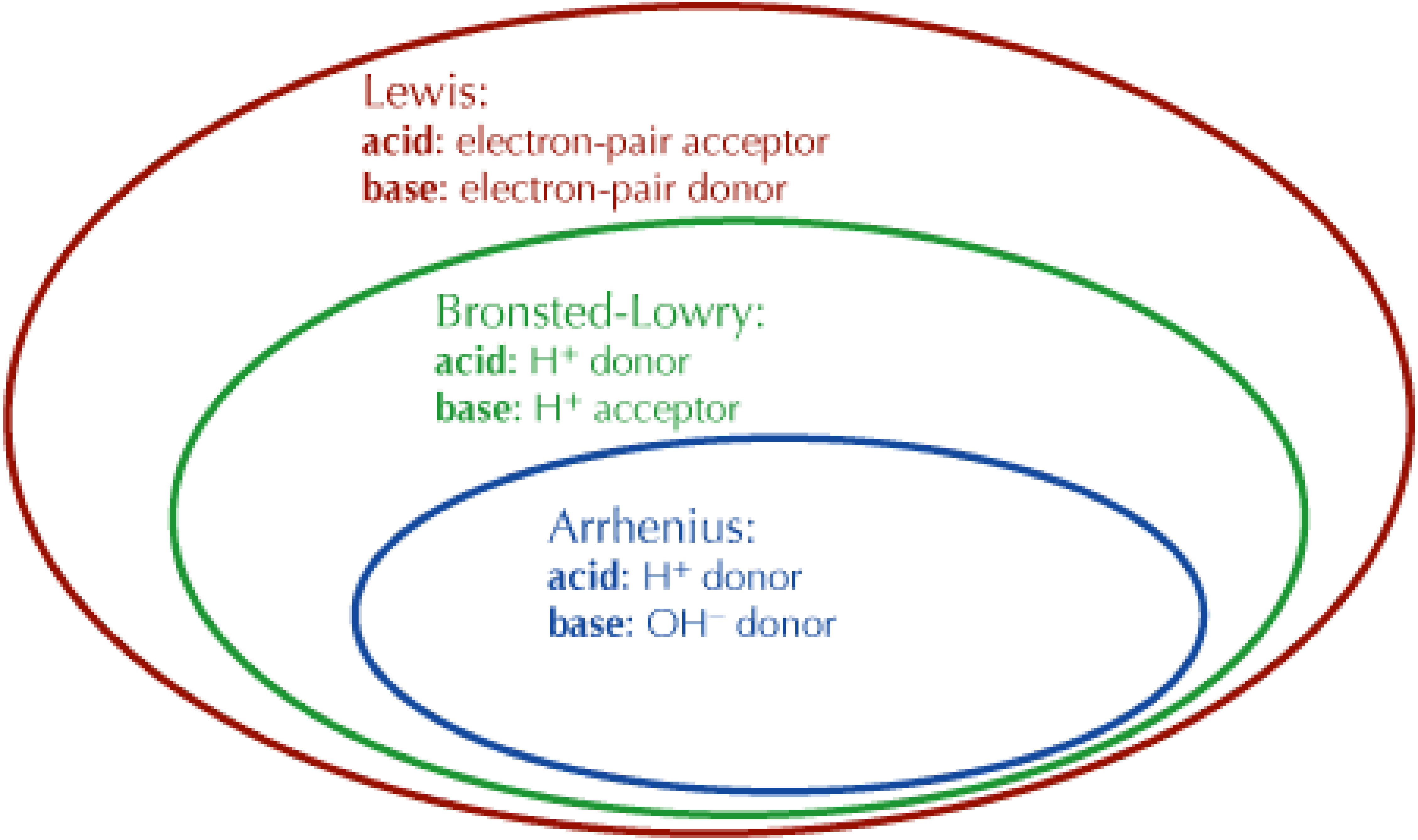




- More general definition of acidity/basicity
- All Brönsted acids/bases are also Lewis acids/bases






Lewis:
acid: electron-pair acceptor
base: electron-pair donor
Bronsted-Lowry:
acid: H+ donor
base: H+ acceptor
Arrhenius:
acid: H+ donor
base: OH− donor

C. Naming acids and bases

1. Rules for naming bases

a. Name the cation

b. Name the anion. This will always be OH, hydroxide

c. No prefixes, no changing the suffix.

Metal hydroxides:

name of metal ion

+

"hydroxide"

Inorganic base	Name of inorganic base
NaOH	Sodium hydroxide
Ba(OH)2	Barium hydroxide
Fe(OH)2	Iron(II) hydroxide
Fe(OH)2	Iron (III) hydroxide
Mg(OH)2	Magnesium hydroxide
Ca(OH)2	Calcium hydroxide
Al(OH)3	Aluminium hydroxide
NH_4OH	Ammonium hydroxide
KOH	Potassium hydroxide
CsOH	Caesium hydroxide
LiOH	Lithium hydroxide

2. Rules for naming acids

a.When the anion comes from the periodic table, start with the prefix hydro-. EX. HF is called **hydro-**. Attach the root. This is the name of the element without the last syllable. EX. HF is called hydro**fluor-**. Now attach the suffic –ic to the end. EX. HF is called hydrofluor**ic**. Lastly, add the word acid to the name. EX. HF is called hydrofluoric **acid.**

b. When the anion is polyatomic, you do not start with the prefix hydro-. Determine the root. This is the name of the polyatomic ion. There is a difference between the 2 kinds of suffixes in polyatomic ions. EX. $HClO_2$ starts out as chlor**ite.** $HClO_3$ starts out as chlor**ate.** Now attach the suffix –ous to the root, *if* the root ends in -ite. EX. $HCLO_2$ (chlorite) becomes chlor**ous.** Attach the suffix –ic to the root, *if* the root ends in -ate. EX. $HClO_3$ (chlorate) becomes chlor**ic.** Lastly, add the word acid to the name. EX. $HClO_2$ finishes as chlorous **acid.** $HClO_3$ finishes as chloric **acid.**

TABLE 6.6 Compounds that are acids in water solution and their anions

Acid formula	*Oxidation number of nonmetal*	*Name in aqueous solution*	*Name and formula of anion*
*HNO_3	+5	nitric acid	nitrate, NO_3^-
HNO_2	+3	nitrous acid	nitrite, NO_2^-
*H_2SO_4	+6	sulfuric acid	sulfate, SO_4^{-2}
H_2SO_3	+4	sulfurous acid	sulfite, SO_3^{-2}
*H_3PO_4	+5	phosphoric acid	phosphate, PO_4^{-3}
H_2CO_3	+4	carbonic acid	carbonate, CO_3^{-2}
$HClO_4$	+7	perchloric acid	perchlorate, ClO_4^-
*$HClO_3$	+5	chloric acid	chlorate, ClO_3^-
$HClO_2$	+3	chlorous acid	chlorite, ClO_2^-
#$HClO$	+1	hypochlorous acid	hypochlorite, ClO^-
HCl	-1	hydrochloric acid	chloride, Cl^-

* These acids are the most common for a particular nonmetal.

Although only chlorine is shown, similar compounds are formed by the other halogens and would be named the same way as are these chlorine-containing compounds.

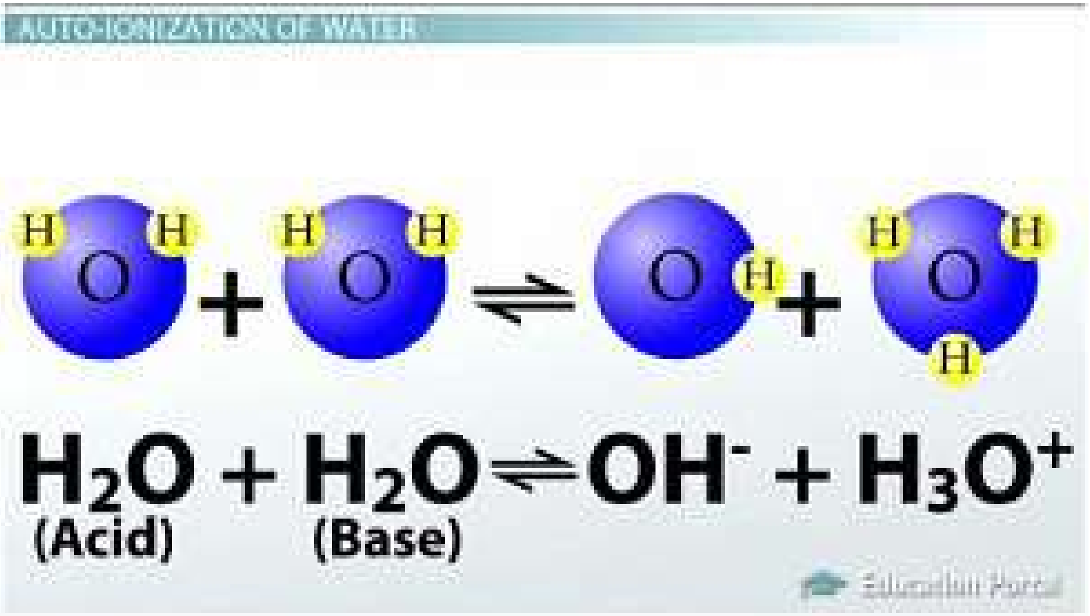
anion	anion name	acid	acid name
Cl^-	chloride ion	HCl	hydrochloric acid
CO_3^{2-}	carbonate ion	H_2CO_3	carbonic acid
NO_2^-	nitrite ion	HNO_2	nitrous acid
NO_3^-	nitrate ion	HNO_3	nitric acid
SO_3^{2-}	sulfite ion	H_2SO_3	sulfurous acid
SO_4^{2-}	sulfate ion	H_2SO_4	sulfuric acid
CH_3COO^-	acetate ion	CH_3COOH	acetic acid

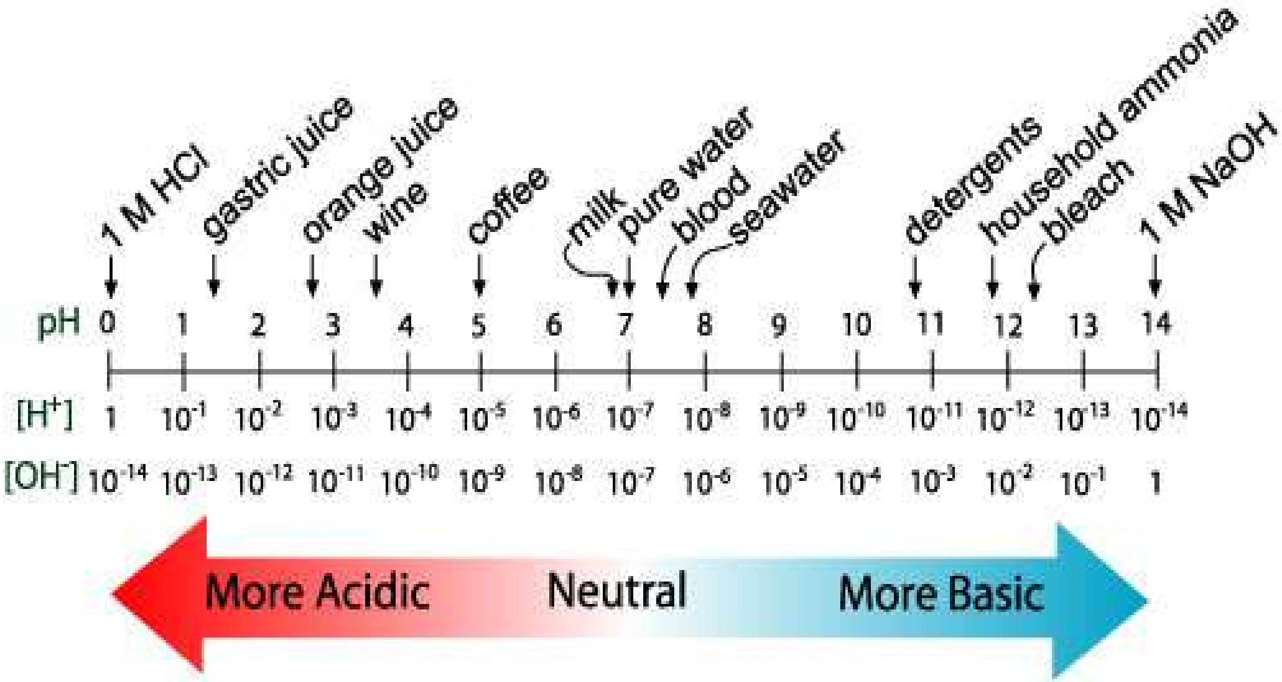
D. The pH concept

1. Plain water will react with itself to become the 2 ions OH^- (hydroxide ion) and H_3O^+ (hydronium ion)

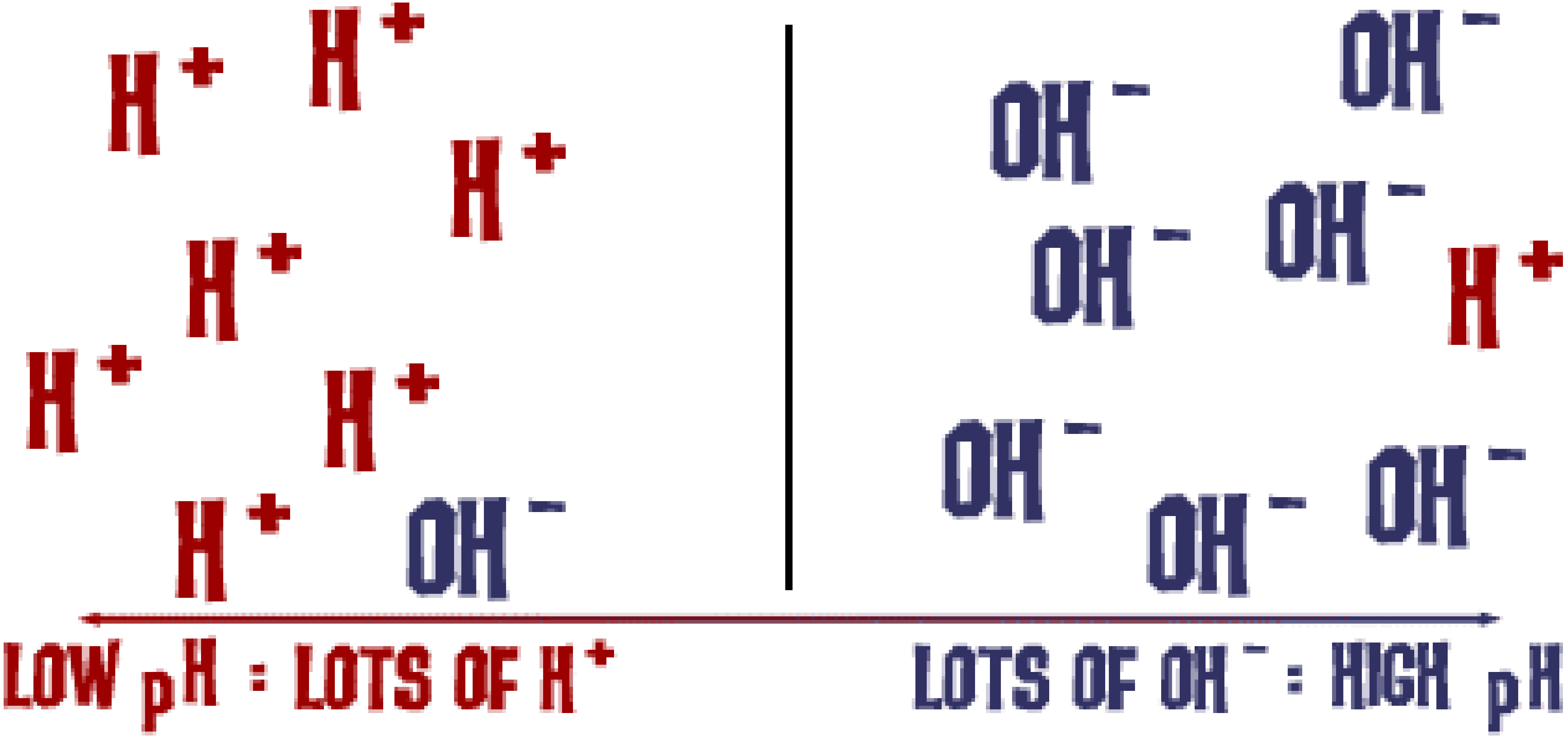
$$\text{EX. } H_2O + H_2O \rightarrow OH^- + H_3O^+$$

2. The product of the concentrations of the hydroxide and hydronium ions is always equal to a total of 1×10^{-14} mol/l.
3. Neutral solutions are such that the number of OH^- and H_3O^+ is the same. Both equal 1×10^{-7} mol/l.

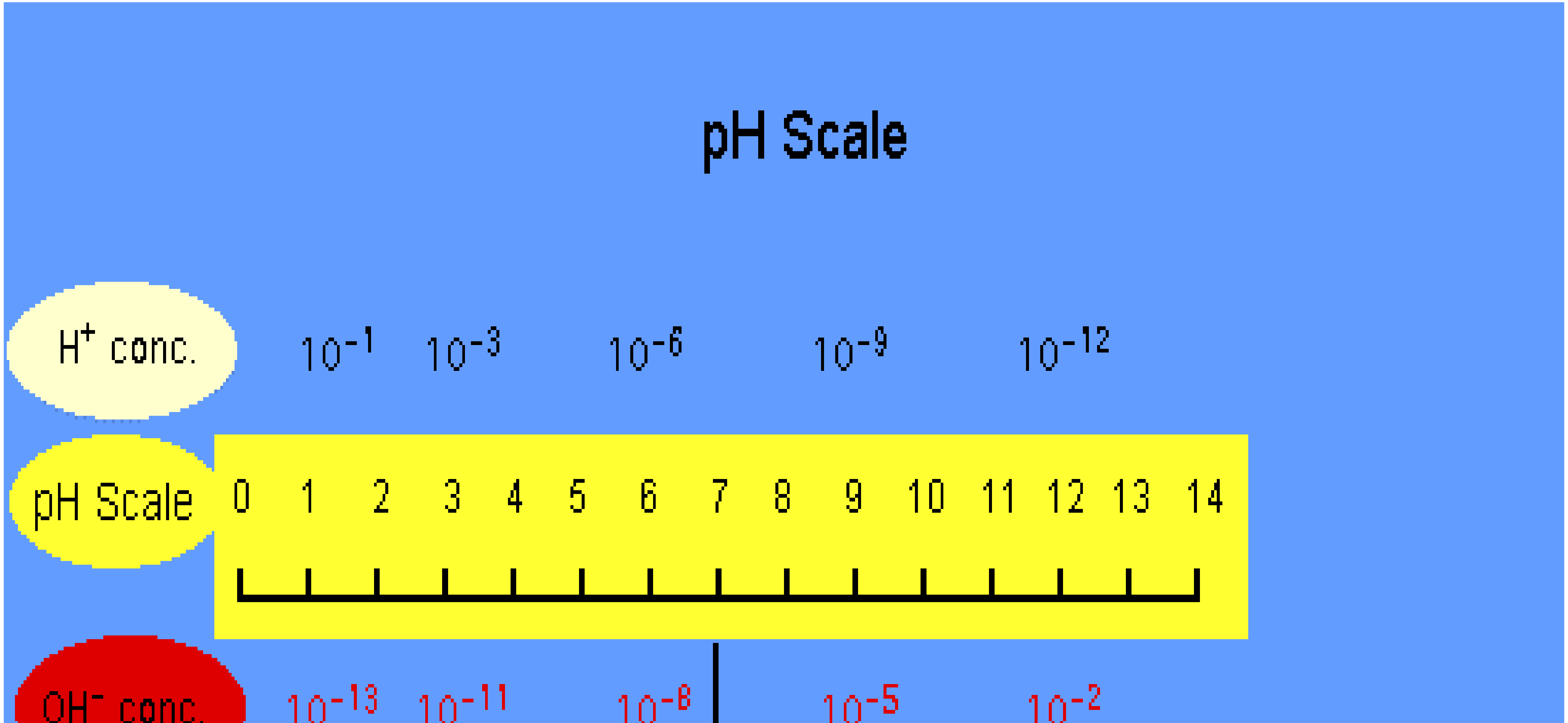




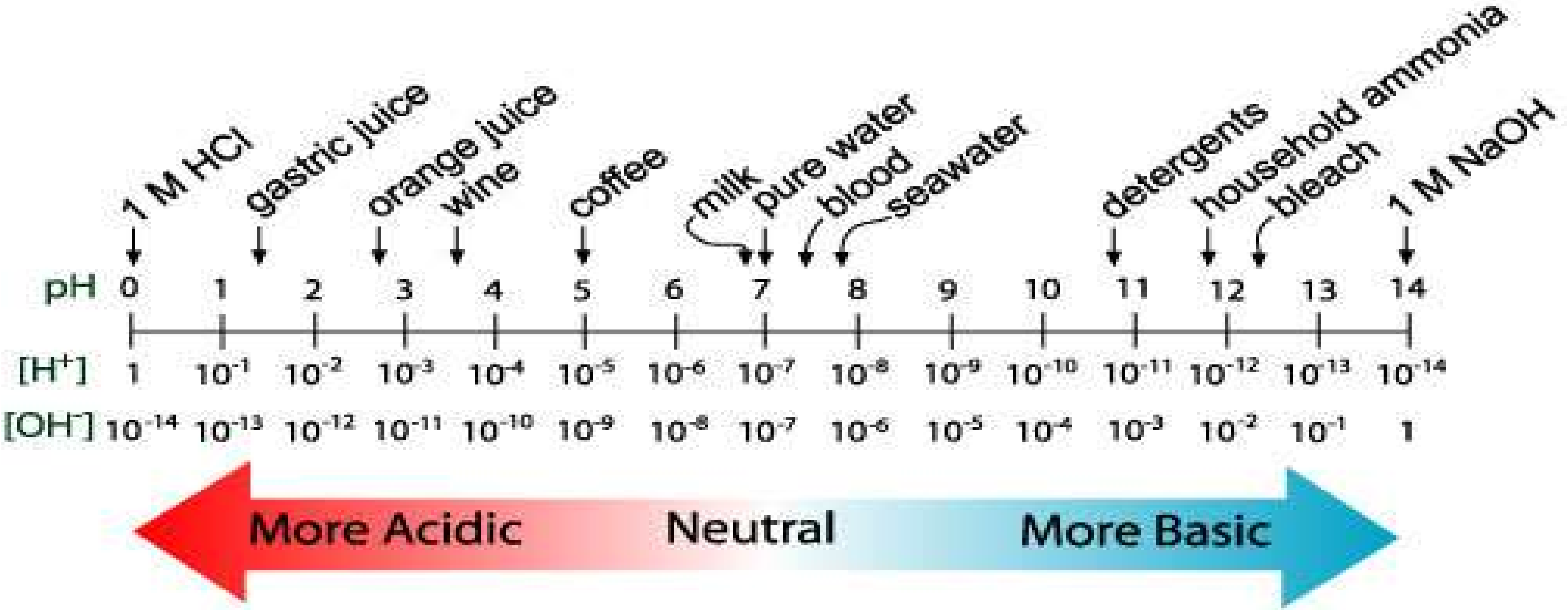
4. An acidic solution is one in which the H_3O^+ concentration is greater than the OH^- concentration and a basic solution has an OH^- concentration greater than the concentration of H_3O^+.



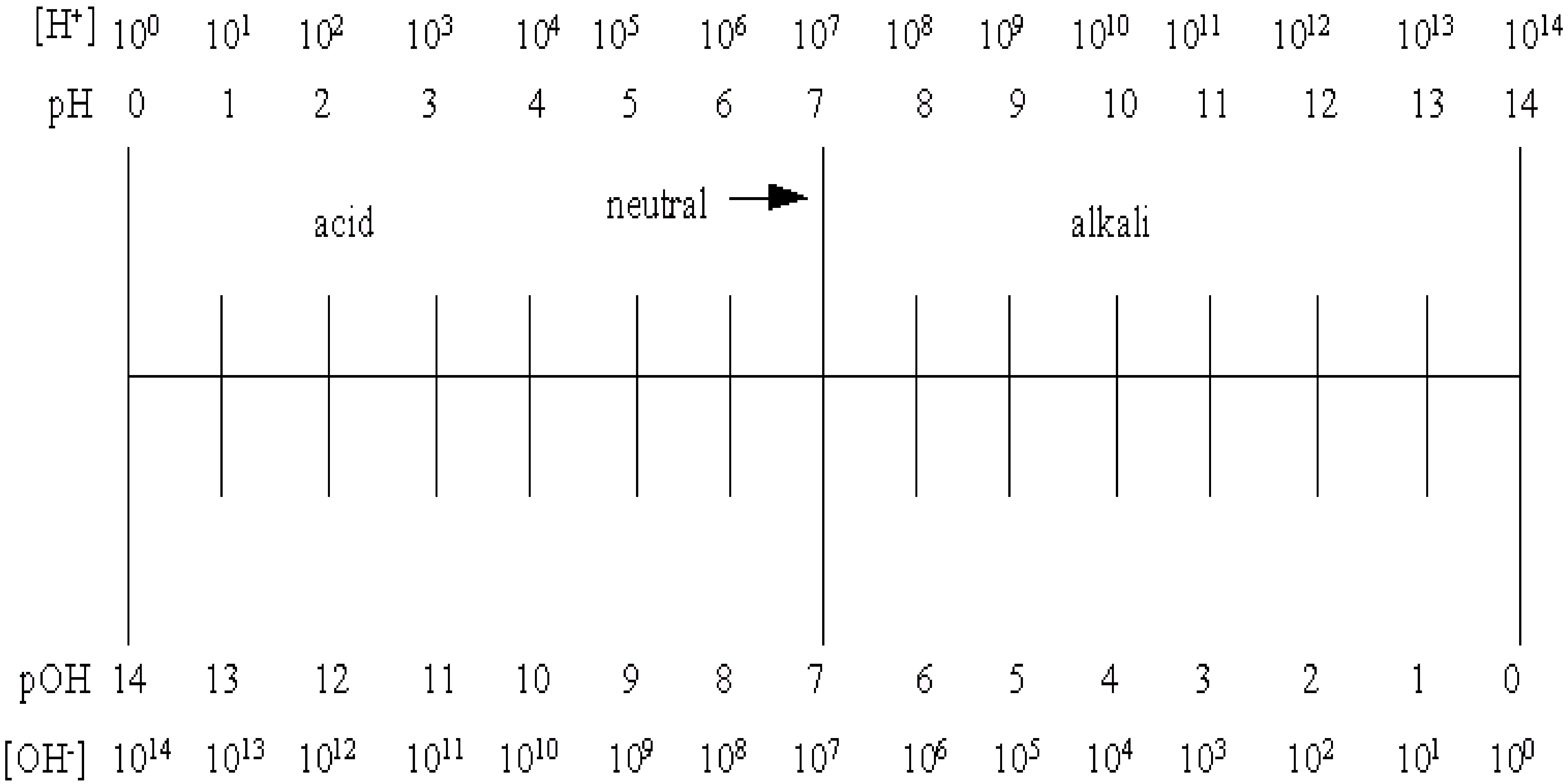
5. Because the nomenclature of writing concentrations in scientific notation (EX. 1×10^{-4}) is so cumbersome, the pH scale was constructed. The power of the 10 becomes a whole number on the scale.



6. The scale runs from 0 – 14. Acid solutions range from 0 – 6. Basic solutions range from 8 – 14. Neutral solutions are at 7. The farther the number moves away from the 7, the stronger the solution. The strongest acid is pH 0, while the strongest base is pH 14. The weakest acid is pH 6, while the weakest base is pH 8.



7. It is called the pH scale because the number on the scale indicates the power (**p**H) of the hydronium (p**H**) ion.
8. There is also a hydroxide scale. This runs oppositely to the pH scale number wise, but equally in ion concentrations.



[H+] 10^0 10^1 10^2 10^3 10^4 10^5 10^6 10^7 10^8 10^9 10^10 10^11 10^12 10^13 10^14
pH 0 1 2 3 4 5 6 7 8 9 10 11 12 13 14
acid
neutral
alkali
pOH 14 13 12 11 10 9 8 7 6 5 4 3 2 1 0
[OH-] 10^14 10^13 10^12 10^11 10^10 10^9 10^8 10^7 10^6 10^5 10^4 10^3 10^2 10^1 10^0

$$pH = -\log[H^+] \qquad pOH = -\log[OH^-]$$

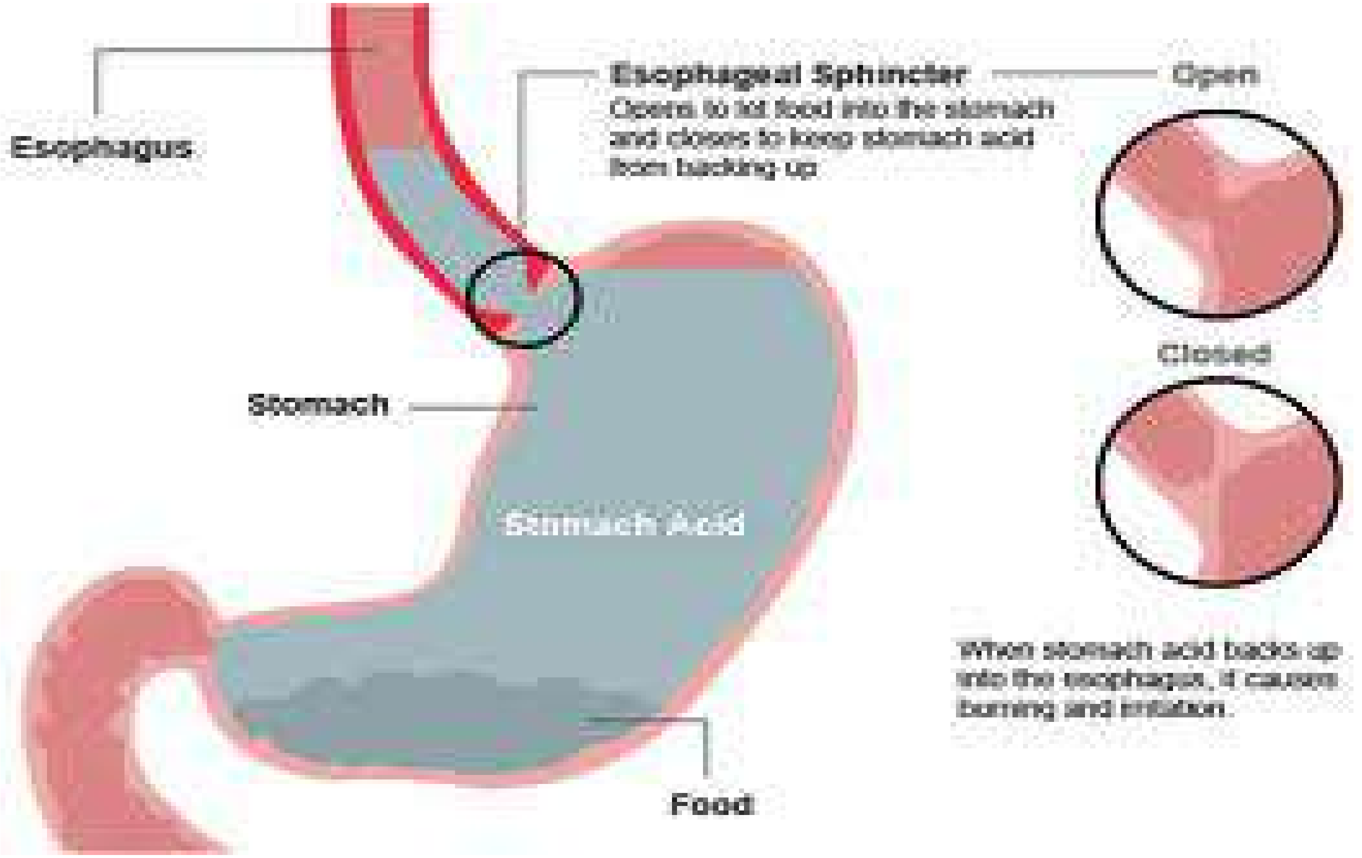
$$[H^+][OH^-] = 10^{-14} \qquad pH + pOH = 14$$

E. Acid – base reactions

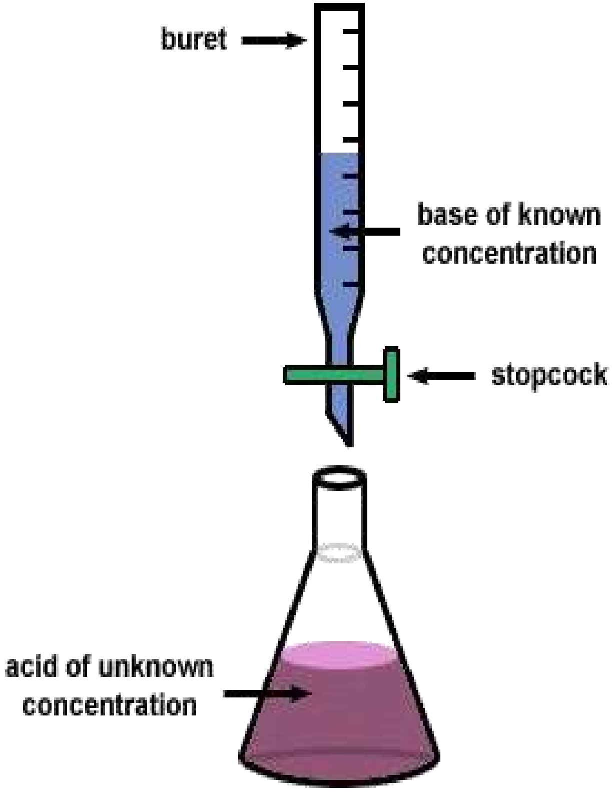
1. Neutralization is the term used to describe the reaction between an acid and a base. This reaction will always yield a salt and water.

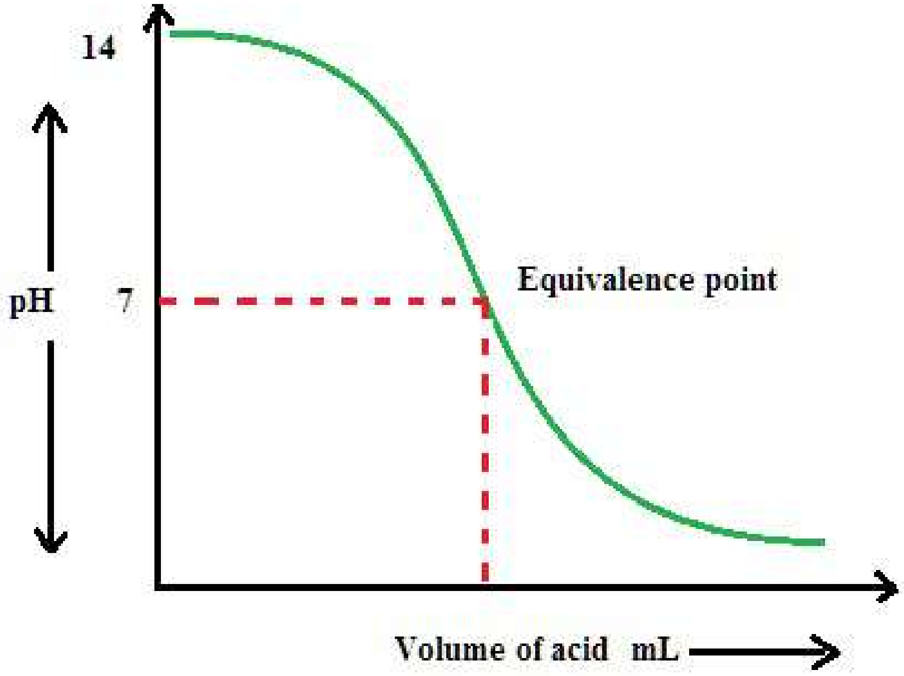
   EX. $HCl + NaOH \rightarrow NaCl + H_2O$

2. This is the reaction that occurs when you take an antacid to settle an upset, ACIDIC stomach.



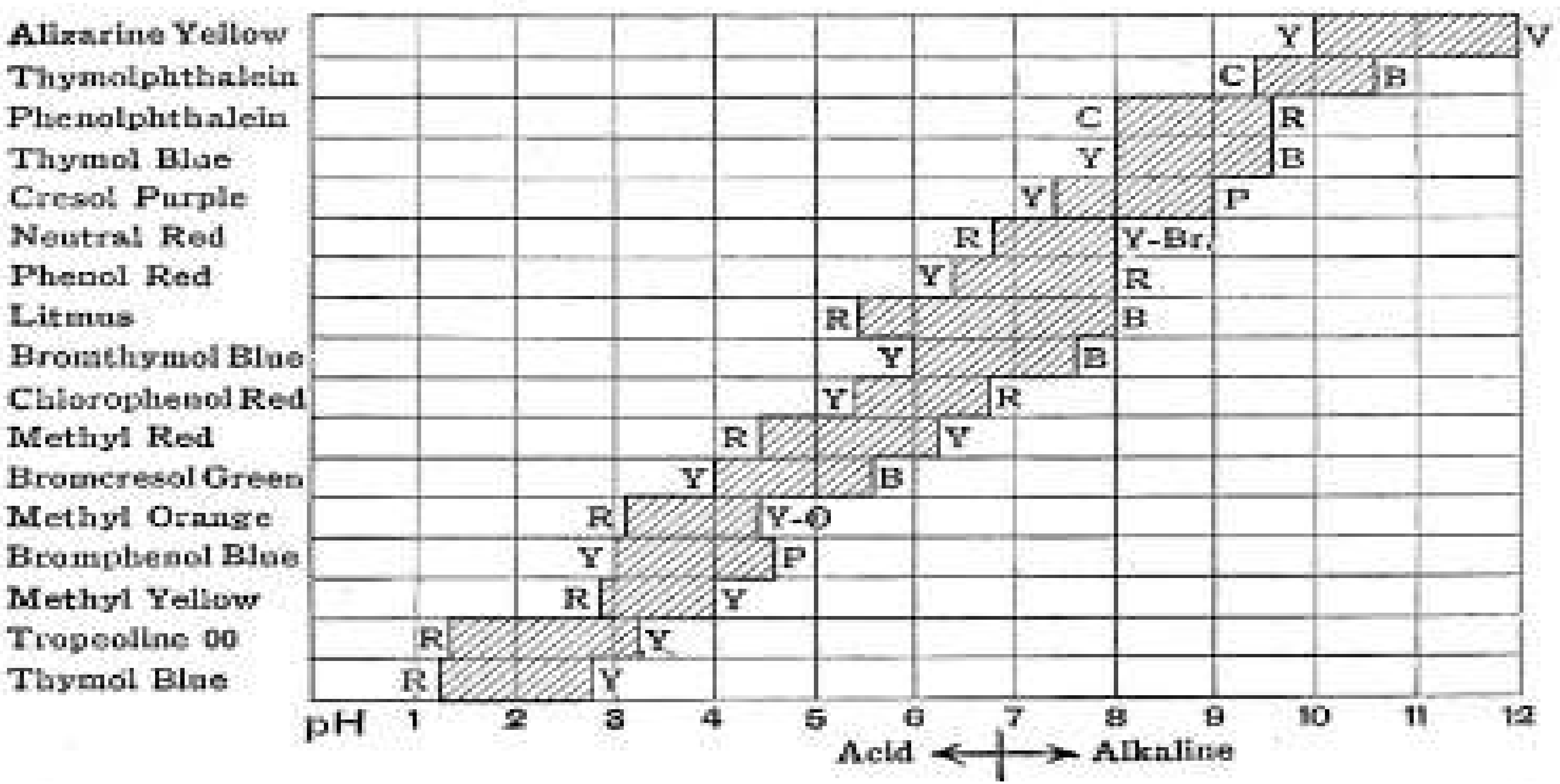
3. How can you determine the concentration of an acid or base? A process called **titration** is used. One substance (acid or base) is added to a known base or acid until an **equivalence point** is reached. A color change will indicate when this happens. This is the point at which the original substance has been completely neutralized. Knowing the concentration of the substance you used and the amount, you can determine the concentration of the unknown.





D. Measuring pH

1. **Indicators** are chemical dyes that change color depending on the pH of the substance

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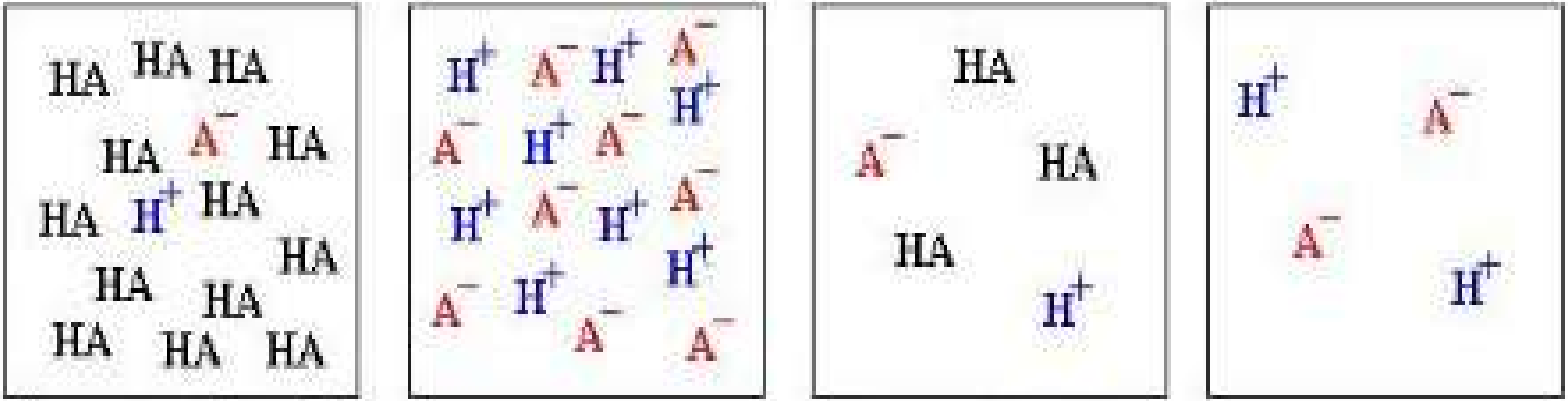
2. **Strong acids** completely dissociate/ionize in aqueous solutions, yielding many hydronium ions. **Weak ones** only do so slightly.



Concentrated weak acid - a lot present, but little dissociation of acid

Concentrated strong acid - a lot present with a lot of dissociation to form many hydrogen ions

Dilute weak acid - little acid present with little dissciation of acid

Dilute strong acid - much acid present with a high degree of dissociation (a)The **acid dissociation constant (Ka)** is the ratio of the concentrations of the dissociated form to the undissociated form of an acid. The larger the Ka, the stronger the acid.

EX. $HF + H2O \longrightarrow H3O^{+} + F^{-}$ $\quad Ka = \dfrac{[H3O^{+}][F^{-}]}{[HF]}$

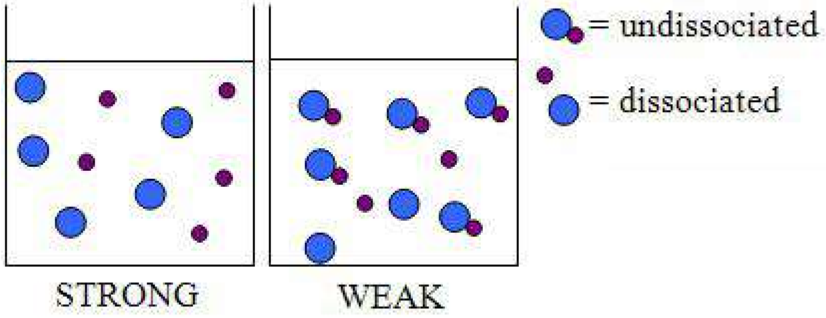
If 10 HF's broke down into 7 $H3O^{+}$ and 7 F^{-} and the last 3 HF did not break down, the Ka would be set up $Ka = \dfrac{[7][7]}{[3]} = 49/3 = 16$

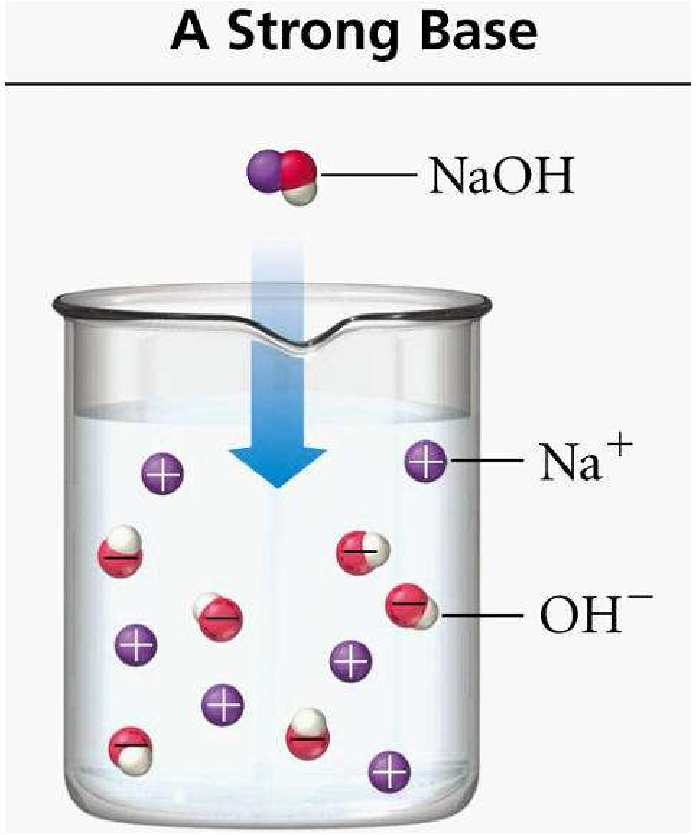
EX. $HCl + H2O \longrightarrow H3O^{+} + Cl^{-}$ $\quad Ka = \dfrac{[H3O^{+}][Cl^{-}]}{[HCl]}$

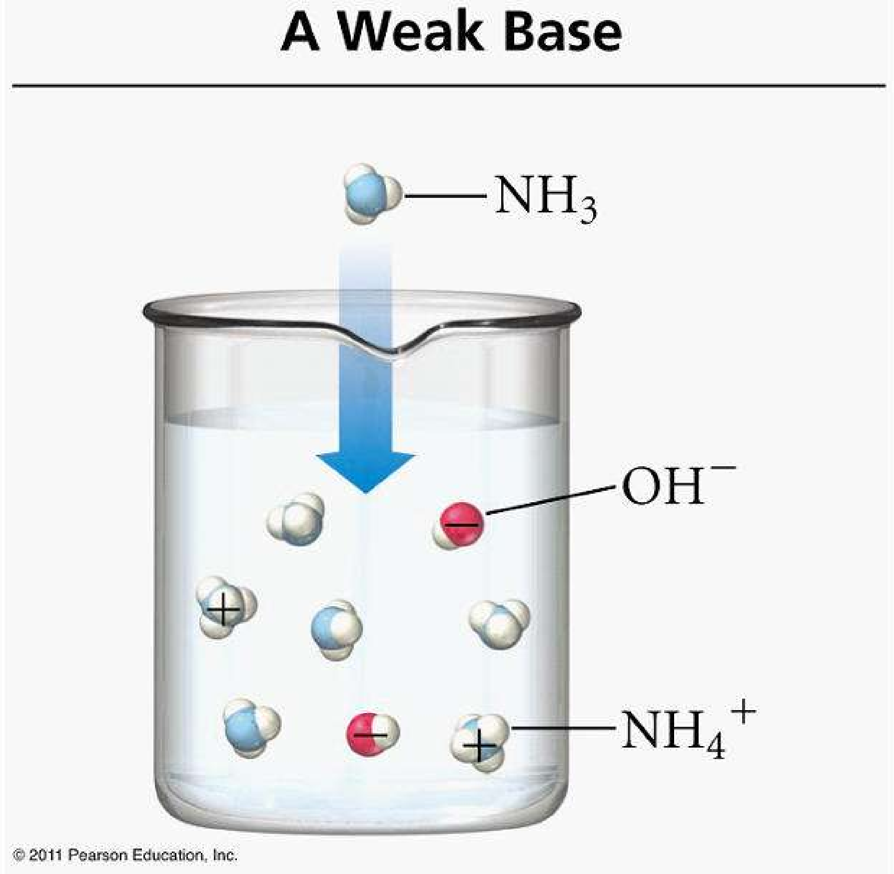
If 10 HCl's broke down into 3 $H3O^{+}$ and 3 Cl^{-} and the last 7 HCl did not break down, the Ka would be set up $Ka = \dfrac{[3][3]}{[7]} = 9/7 = 1.3$

In the 2 examples above, the HF has more dissociation (7 instead of 3), more H^{+} / H_3O^{+} (7 instead of 3) production, a higher Ka (16 instead of 1.3), so is termed a stronger acid than the HCl.

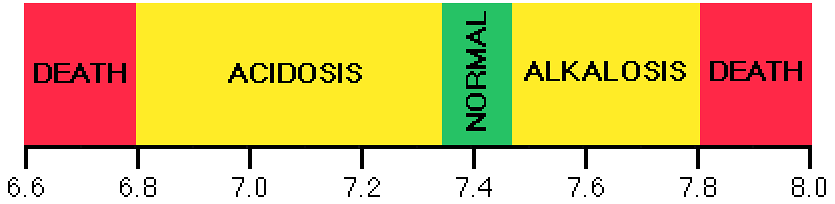
3. Just like acids, **strong bases** dissociate completely in aqueous solution while **weak ones** react less.

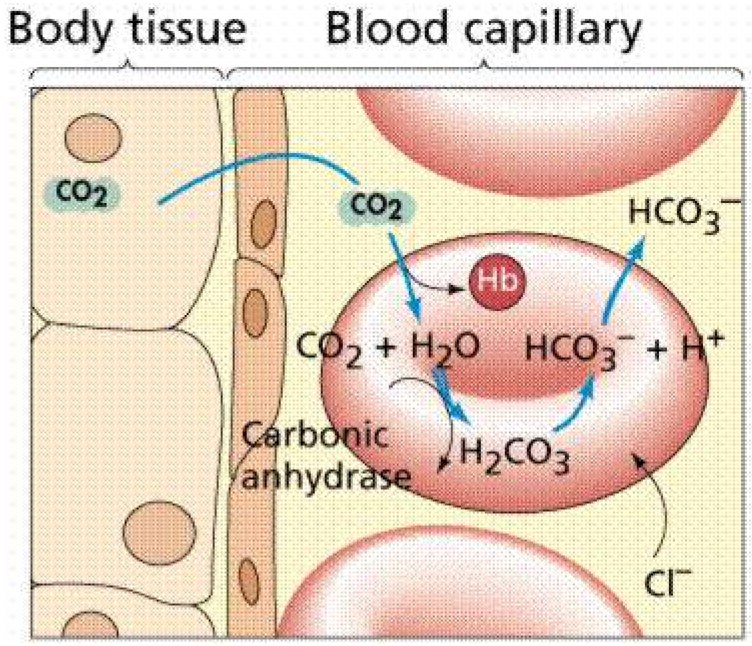






4. **Buffers** are solutions that resist chance in pH even when acids / bases are added to it. This is important in body chemistry since a strict pH range must be observed in order for necessary chemical reactions to be carried out. Buffering is accomplished using a **conjugate acid-base pair**. In blood, this pair is H_2CO_3 and CO_3^{-2}. When a person starts to hyperventilate, their blood gets too basic, the pH too high. There is not enough H+ in it. The H_2CO_3 will dissociate to become HCO_3^{-1} or CO_3^{-2}, thus releasing some H+ to make the blood more acidic. This will bring the pH back down. When a person holds their breath, their blood gets too acidic. The pH is too low. There is too much H+ in the blood. The CO_3^{-2} will pick up the excess H+ to become HCO_3^{-1} or H_2CO_3. This makes the blood more acidic, raising the pH back up.





How it Works:

$$A^- \rightarrow HA$$

H_3O^+ **Buffer** H_2O

Remember pH = Conc. of H_3O^+

Your blood	Rxn: $HCO_3^- \rightleftarrows H_2CO_3$
Excess H_3O^+	$H_3O^+ + HCO_3^- \rightarrow H_2CO_3 + H_2O$ $\rightarrow CO_2 + H_2O$
Excess OH^-	$OH^- + H_2CO_3 \rightleftarrows HCO_3^- + H_2O$