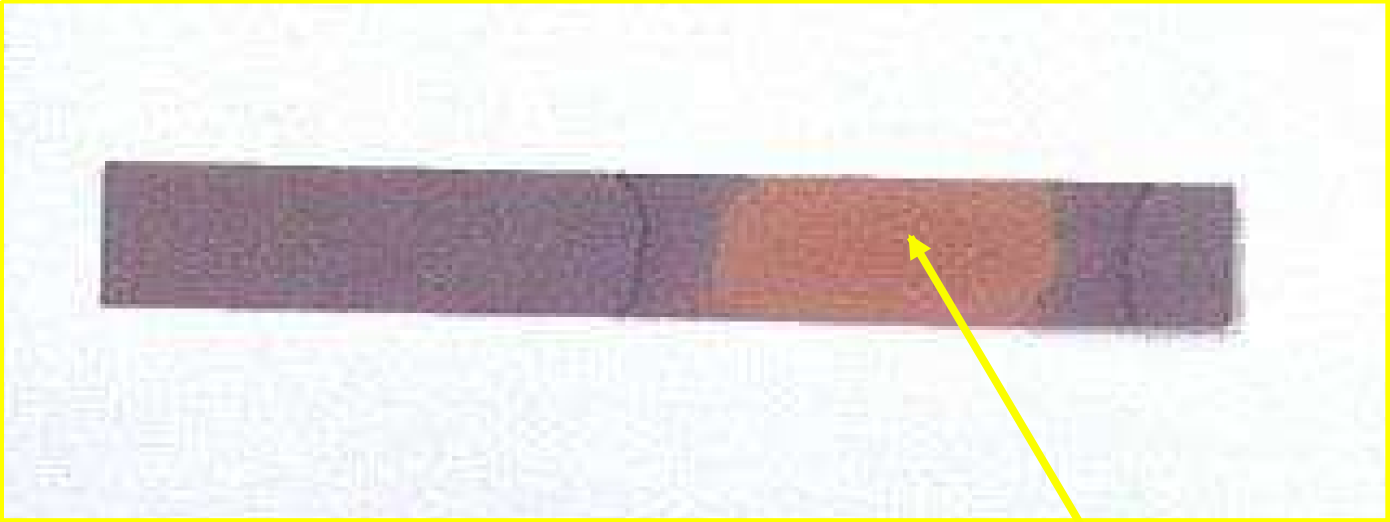
Properties of Acids

- They taste *sour* (don't try this at home).
- They can conduct electricity.
  - Can be strong or weak electrolytes in aqueous solution
- React with metals to form H_2 gas.
- Change the color of indicators (for example: blue litmus turns to red).
- React with bases (metallic hydroxides) to form water and a salt.

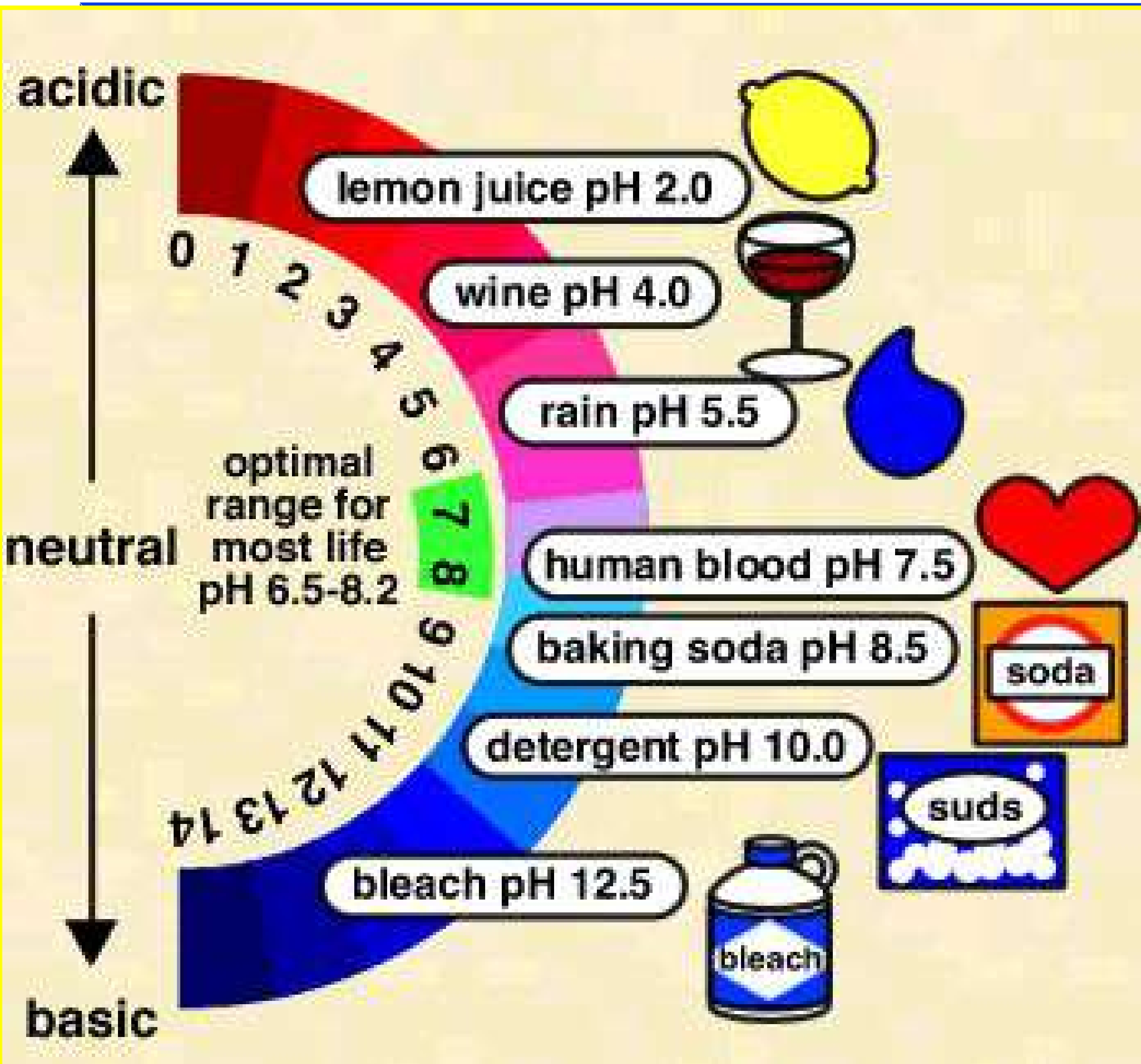
Properties of Acids

- They have a pH of less than 7 (more on this concept of pH in a later lesson)
- They react with carbonates and bicarbonates to produce a salt, water, and carbon dioxide gas
- How do you know if a chemical is an acid?
  - It usually starts with Hydrogen.
  - $\underline{H}Cl$, $\underline{H}_2SO_4$, $\underline{H}NO_3$, etc. (but not water!)

Acids Affect Indicators, by changing their color



Blue litmus paper turns red in contact with an acid **(and red paper stays red).**



Acids have a pH **less** than 7

Acids React with Active Metals

Acids react with active metals to form salts and hydrogen gas:

$$HCl_{(aq)} + Mg_{(s)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

This is a single-replacement reaction

Acids React with Carbonates and Bicarbonates

$HCl + NaHCO_3$

Hydrochloric acid + sodium bicarbonate

$$\downarrow$$

$NaCl + H_2O + CO_2$

salt + water + carbon dioxide

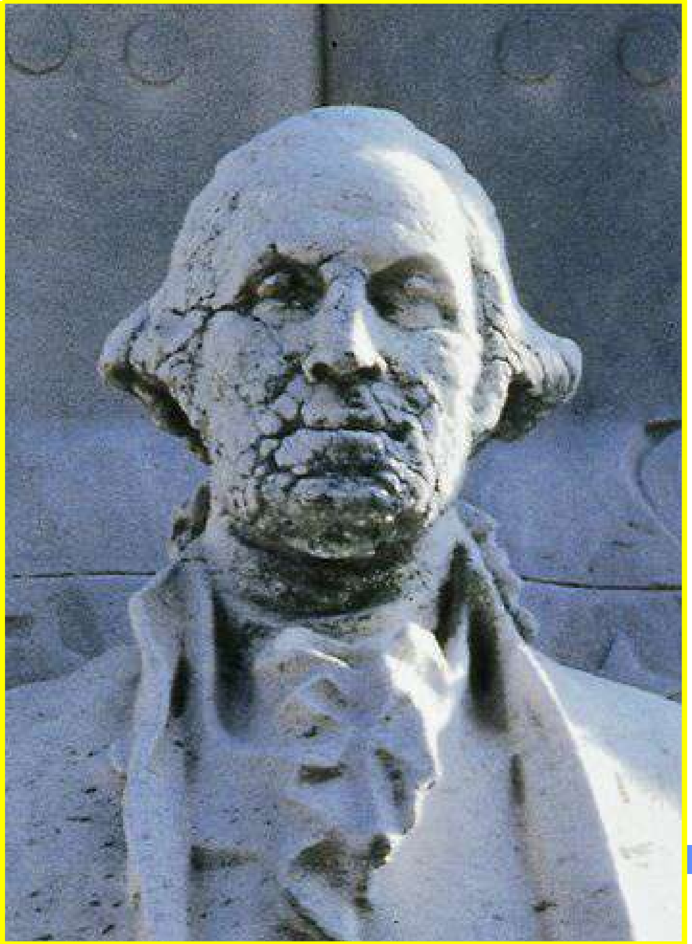
An old-time home remedy for relieving an upset stomach

Effects of *Acid Rain* on Marble

(marble is calcium carbonate)

George Washington: BEFORE acid rain

George Washington: AFTER acid rain



Acids *Neutralize* Bases

$HCl + NaOH \rightarrow NaCl + H_2O$

-Neutralization reactions ALWAYS produce a salt (which is an *ionic compound*) and water.

-Of course, it takes the *right proportion* of acid and base to produce a neutral salt

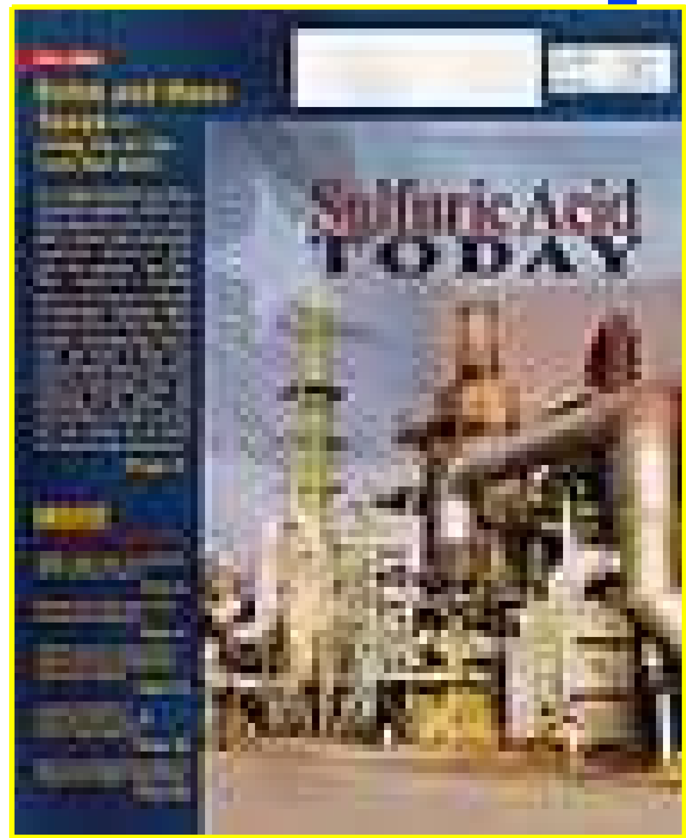
Sulfuric Acid = H_2SO_4

4 Highest volume production of ***any*** chemical in the U.S. (approximately 60 billion pounds/year)

4 Used in the production of paper

4 Used in production of fertilizers

4 Used in petroleum refining; auto batteries



Nitric Acid = HNO_3

- Used in the production of fertilizers
- Used in the production of explosives
- Nitric acid is a *volatile* acid – its reactive components evaporate easily
- Stains proteins yellow (including skin!)

Hydrochloric Acid = HCl

4 Used in the "pickling" of steel

4 Used to purify magnesium from sea water

4 Part of gastric juice, it aids in the digestion of proteins

4 Sold commercially as *Muriatic acid*

Phosphoric Acid = H_3PO_4

- A flavoring agent in sodas (adds "tart")
- Used in the manufacture of detergents
- Used in the manufacture of fertilizers
- *Not* a common laboratory reagent

Acetic Acid = $HC_2H_3O_2$

(also called Ethanoic Acid, CH_3COOH)

4 Used in the manufacture of plastics

4 Used in making pharmaceuticals

4 Acetic acid is the acid that is present in household *vinegar*

Properties of Bases (metallic hydroxides)

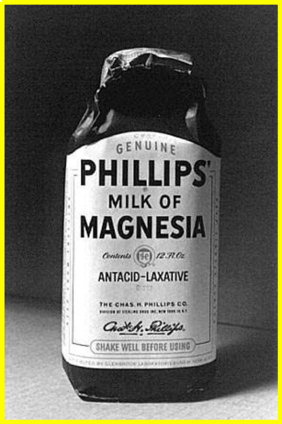
- React with acids to form water and a salt.
- Taste bitter.
- Feel slippery (don't try this either).
- Can be strong or weak electrolytes in aqueous solution
- Change the color of indicators (red litmus turns blue).

Examples of Bases
(metallic hydroxides)

- Sodium hydroxide, NaOH (*lye for drain cleaner; soap*)

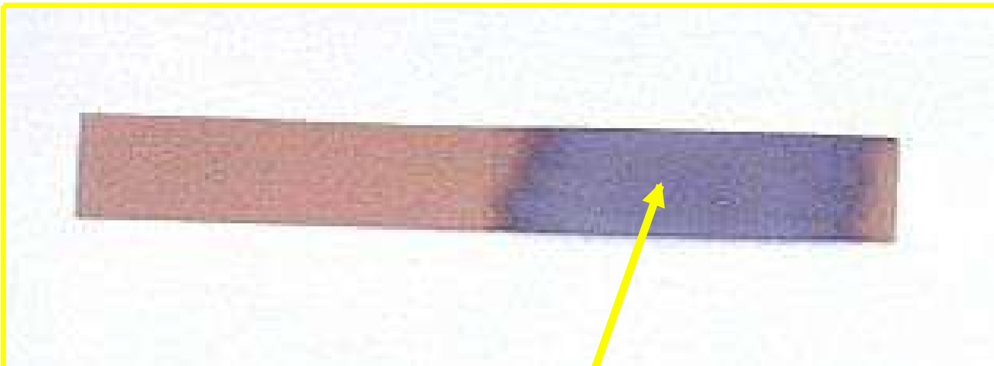
- Potassium hydroxide, KOH (*alkaline batteries*)

- Magnesium hydroxide, $Mg(OH)_2$ (*Milk of Magnesia*)



- Calcium hydroxide, $Ca(OH)_2$ (*lime*; *masonry*)

Bases Affect Indicators



Red litmus paper turns blue in contact with a base (and blue paper stays blue).

Phenolphthalein turns purple in a base.

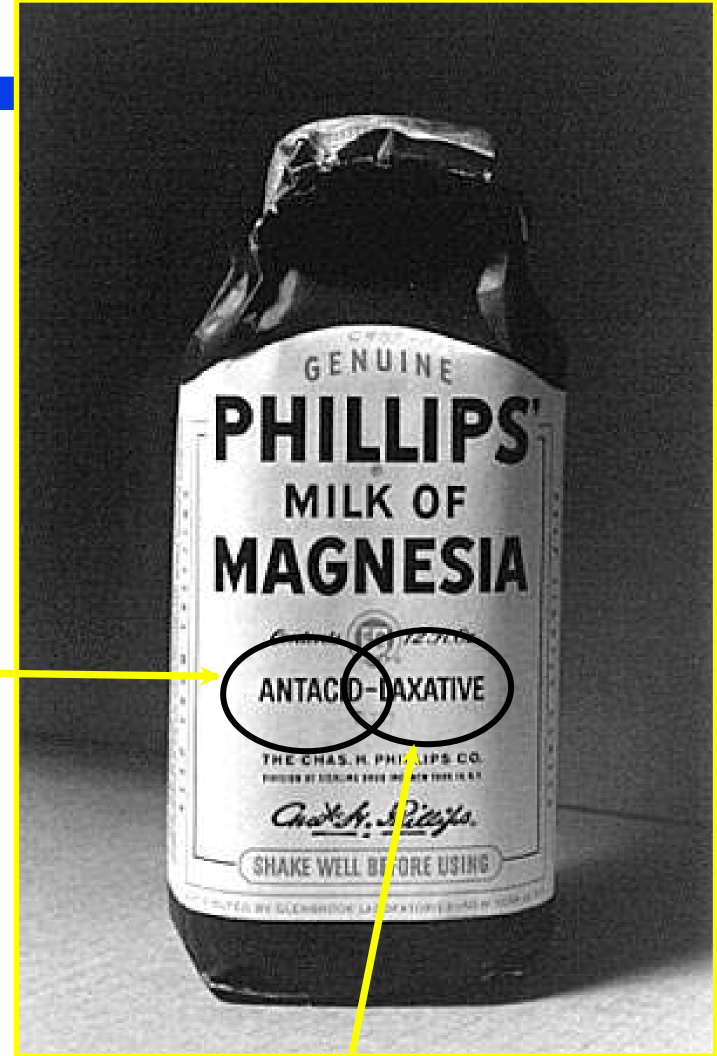
Bases Neutralize Acids

Milk of Magnesia contains magnesium hydroxide, $Mg(OH)_2$, which neutralizes stomach acid, HCl.

$$2\ HCl + Mg(OH)_2$$

$$\downarrow$$

$$MgCl_2 + 2\ H_2O$$



Magnesium salts can cause diarrhea (thus they are used as a laxative) and may also cause kidney stones.

Svante Arrhenius

- He was a Swedish chemist (1859-1927), and a Nobel prize winner in chemistry (1903)
- one of the first chemists to explain the chemical theory of the behavior of acids and bases
- Dr. Hubert Alyea (professor emeritus at Princeton University) was the last graduate student of Arrhenius.

1. Arrhenius Definition - 1887

- Acids produce hydrogen ions (H^{1+}) *in aqueous solution* ($HCl \rightarrow H^{1+} + Cl^{1-}$)
- Bases produce hydroxide ions (OH^{1-}) when dissolved *in water*.
  ($NaOH \rightarrow Na^{1+} + OH^{1-}$)
- Limited to **aqueous solutions**.
- Only one kind of base (hydroxides)
- NH_3 (ammonia) could not be an Arrhenius base: no OH^{1-} produced.

Polyprotic Acids?

- Some compounds have *more than one ionizable hydrogen* to release
- HNO_3 nitric acid - <u>mono</u>protic
- H_2SO_4 sulfuric acid - <u>di</u>protic - 2 H^+
- H_3PO_4 phosphoric acid - <u>tri</u>protic - 3 H^+
- Having more than one ionizable hydrogen does <u>**not**</u> mean stronger!

Acids

- Not all compounds that have hydrogen are acids. Water?
- Also, not all the hydrogen in an acid may be released as ions
  - only those that have *very polar bonds* are ionizable - this is when the *hydrogen is joined to a very electronegative element*

Arrhenius examples...

- Consider HCl = it is an acid!
- What about CH_4 (methane)?
- CH_3COOH (ethanoic acid, also called acetic acid) - it has 4 hydrogens just like methane does…?
- Table 19.2, p. 589 for bases, which are metallic hydroxides

Organic Acids (those with carbon)

Organic acids all contain the ***carboxyl*** group, (-COOH), sometimes several of them.

CH_3COOH – of the 4 hydrogen, only 1 ionizable

(due to being bonded to the highly electronegative **Oxygen**)

The carboxyl group is a poor proton donor, so **ALL organic acids are <u>weak acids</u>.**

2. Brønsted-Lowry - 1923

- A broader definition than Arrhenius
- Acid is hydrogen-ion donor (H^+ or proton); base is hydrogen-ion acceptor.
- Acids and bases always come in pairs.
- HCl is an acid.
  - When it dissolves in water, it gives it's proton to water.

$$HCl_{(g)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

- Water is a base; makes hydronium ion.

Why Ammonia is a Base

- Ammonia can be explained as a base by using Brønsted-Lowry:

$NH_{3(aq)} + H_2O_{(l)} \leftrightarrow NH_4^{1+}{}_{(aq)} + OH^{1-}{}_{(aq)}$

Ammonia is the hydrogen ion acceptor (base), and water is the hydrogen ion donor (acid).

This causes the OH^{1-} concentration to be greater than in pure water, and the ammonia solution is basic

Acids and bases come in pairs

- A "**conjugate base**" is the remainder of the original acid, after it donates it's hydrogen ion
- A "**conjugate acid**" is the particle formed when the original base gains a hydrogen ion
- Thus, a conjugate acid-base pair is related by the *loss or gain* of a single hydrogen ion.
- Chemical Indicators? They are weak acids or bases that have a different color from their original acid and base

Acids and bases come in pairs

- General equation is:

$HA_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + A^-_{(aq)}$

- Acid + Base $\leftrightarrow$ Conjugate acid + Conjugate base
- $NH_3 + H_2O \leftrightarrow NH_4^{1+} + OH^{1-}$

  base acid c.a. c.b.
- $HCl + H_2O \leftrightarrow H_3O^{1+} + Cl^{1-}$

  acid base c.a. c.b.
- Amphoteric – a substance that can act as both an acid and base- as water shows

3. Lewis Acids and Bases

- Gilbert Lewis focused on the donation or acceptance of a pair of electrons during a reaction
- Lewis Acid - electron pair acceptor
- Lewis Base - electron pair donor
- Most general of all 3 definitions; acids don’t even need hydrogen!
- **Summary**: Table 19.4, page 592

Hydrogen Ions from Water

- Water ionizes, or falls apart into ions:

  $$H_2O \leftrightarrow H^{1+} + OH^{1-}$$

- Called the "*self ionization*" of water
- Occurs to a *very small extent*:

  $$[H^{1+}] = [OH^{1-}] = 1 \times 10^{-7}\ M$$

- Since they are equal, a neutral solution results from water

$$K_w = [H^{1+}] \times [OH^{1-}] = 1 \times 10^{-14}\ M^2$$

- K_w is called the "*ion product constant*" for water

Ion Product Constant

- $H_2O \leftrightarrow H^{1+} + OH^{1-}$
- K_w is constant in every aqueous solution: $[H^+] \times [OH^-] = 1 \times 10^{-14}\ M^2$
- If $[H^+] > 10^{-7}$ then $[OH^-] < 10^{-7}$
- If $[H^+] < 10^{-7}$ then $[OH^-] > 10^{-7}$
- If we know one, other can be determined
- If $[H^+] > 10^{-7}$, it is acidic and $[OH^-] < 10^{-7}$
- If $[H^+] < 10^{-7}$, it is basic and $[OH^-] > 10^{-7}$

The pH concept – from 0 to 14

- pH = *pouvoir hydrogene* (Fr.)
  "hydrogen power"
- definition: **$pH = -\log[H^+]$**
- in neutral pH = $-\log(1 \times 10^{-7})$ **= 7**
- in acidic solution $[H^+] > 10^{-7}$
- pH **<** $-\log(10^{-7})$
  - pH < 7 (from 0 to 7 is the acid range)
  - in base, pH > 7 (7 to 14 is base range)

Calculating pOH

- $pOH = -\log [OH^-]$
- $[H^+] \times [OH^-] = 1 \times 10^{-14}\ M^2$
- $pH + pOH = 14$
- Thus, a solution with a pOH less than 7 is basic; with a pOH greater than 7 is an acid
- Not greatly used like pH is.

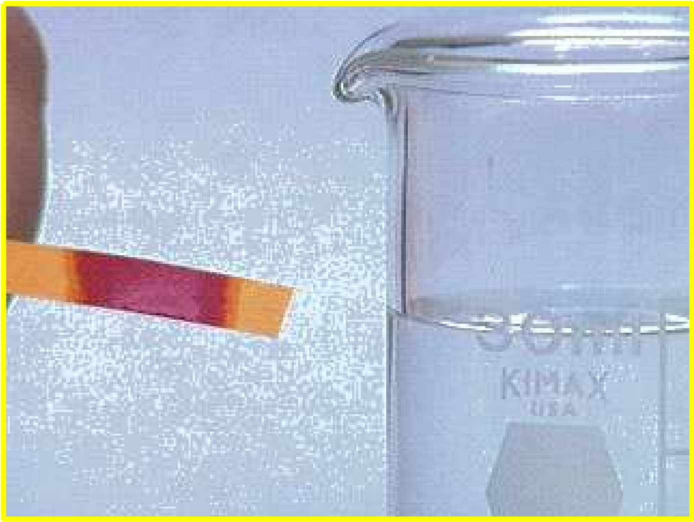
pH and Significant Figures

- For pH calculations, the hydrogen ion concentration is usually expressed in scientific notation
- $[H^{1+}] = 0.0010\ M = 1.0 \times 10^{-3}\ M$, and $0.00\underline{10}$ has 2 significant figures
- the pH = $3.\underline{00}$, with the two numbers to the right of the decimal corresponding to the two significant figures

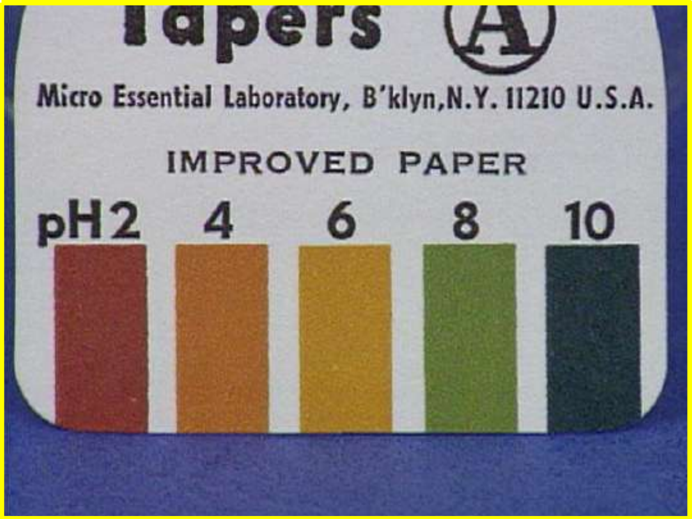
Measuring pH

- Why measure pH?
  - *Everyday solutions* we use - everything from swimming pools, soil conditions for plants, medical diagnosis, soaps and shampoos, etc.
- Sometimes we can use *indicators*, other times we might need a *pH meter*

How to measure pH with wide-range paper



1. Moisten the pH indicator paper strip with a few drops of solution, by using a stirring rod.



2.Compare the color to the chart on the vial – then read the pH value.

Acid-Base Indicators

- Although useful, there are *limitations* to indicators:
  - usually given for a certain temperature (25 °C), thus may change at different temperatures
  - what if the solution already has a color, like paint?
  - the ability of the human eye to distinguish colors is limited

Acid-Base Indicators

- A pH meter may give more definitive results
  - some are **large**, others portable
  - works by measuring the *voltage* between two electrodes; typically accurate to within 0.01 pH unit of the true pH
  - Instruments need to be *calibrated*
  - Fig. 19.15, p.603

Strength

- Acids and Bases are classified acording to the degree to which they ionize in water:
  - **Strong** are *completely ionized* in aqueous solution; this means they ionize 100 %
  - **Weak** *ionize only slightly* in aqueous solution
- ***Strength*** is very different from ***Concentration***

Strength

- Strong – means it forms ***many*** ions when dissolved (100 % ionization)
- $Mg(OH)_2$ is a strong base- it falls completely apart (nearly 100% when dissolved).
  - But, *not much dissolves-* so it is not concentrated

Measuring strength

- Ionization is reversible:

$HA + H_2O \leftrightarrow H^+ + A^-$

- This makes an equilibrium

(Note that the arrow goes both directions.)

- Acid dissociation constant = K_a
- $K_a = \frac{[H^+][A^-]}{[HA]}$

(Note that water is **NOT** shown, because its concentration is constant, and built into K_a)

- Stronger acid = more products (ions), thus a larger K_a (Table 19.7, page 607)

What about bases?

- Strong bases dissociate completely.
- $MOH + H_2O \leftrightarrow M^+ + OH^-$ (M = a metal)
- Base dissociation constant = K_b
- $K_b = \dfrac{[M^+][OH^-]}{[MOH]}$
- Stronger base = more dissociated ions are produced, thus a larger K_b.

Strength vs. Concentration

- The words *concentrated* and *dilute* tell how much of an acid or base is dissolved in solution - refers to the number of moles of acid or base in a given volume
- The words *strong* and *weak* refer to the extent of ionization of an acid or base
- Is a *concentrated, weak* acid possible?

Practice

- Write the K_a expression for HNO_2

1) Equation: $HNO_2 \leftrightarrow H^{1+} + NO_2^{1-}$

2) $K_a = \dfrac{[H^{1+}] \times [NO_2^{1-}]}{[HNO_2]}$

- Write the K_b expression for NH_3 (as NH_4OH)

Acid-Base Reactions

- **Acid + Base ← Water + Salt**
- Properties related to every day:
  - antacids depend on neutralization
  - farmers adjust the soil pH
  - formation of cave stalactites
  - human body kidney stones from insoluble salts

Acid-Base Reactions

n <u>Neutralization Reaction</u> - a reaction in which an acid and a base react in an aqueous solution to produce a salt and water:

$HCl_{(aq)} + NaOH_{(aq)} \leftarrow NaCl_{(aq)} + H_2O_{(l)}$

$H_2SO_{4(aq)} + 2KOH_{(aq)} \leftarrow K_2SO_{4(aq)} + 2$

$H_2O_{(l)}$

– Table 19.9, page 613 lists some salts

Titration

- Titration is the process of adding a known amount of solution of known concentration to determine the concentration of another solution
- Remember? - a **balanced equation** is a ***mole ratio***
- The *equivalence point* is when the moles of hydrogen ions is ***equal*** to the moles of hydroxide ions (= neutralized!)

Titration

- The concentration of acid (or base) in solution can be determined by performing a neutralization reaction
  - An ***indicator*** is used to show when neutralization has occurred
  - Often we use *phenolphthalein*- because it is colorless in neutral and acid; turns pink in base

Steps - Neutralization reaction

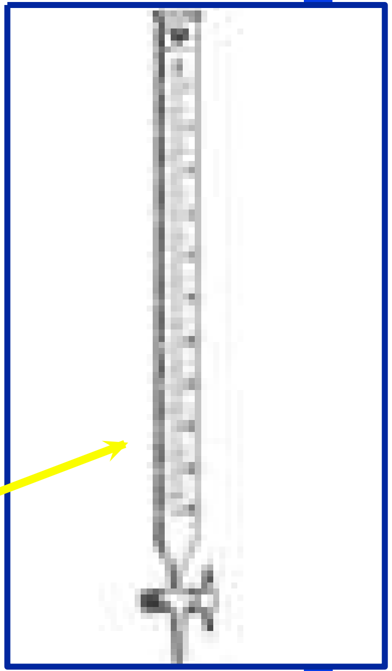
#1. A measured volume of acid of unknown concentration is added to a flask

#2. Several drops of indicator added

#3. A base of known concentration is slowly added, until the indicator changes color; measure the volume

– Figure 19.22, page 615

Neutralization



- The solution of known concentration is called the <u>standard solution</u>
  - added by using a <u>buret</u>
- Continue adding until the indicator *changes color*
  - called the "***end point***" of the titration
  - Sample Problem 19.7, page 616

Salt Hydrolysis

- A **salt** is an ionic compound that:
  - comes from the anion of an acid
  - comes from the cation of a base
  - is formed from a neutralization reaction
  - some neutral; others acidic or basic
- "Salt hydrolysis" - a *salt* that reacts with water to produce an acid or base

Salt Hydrolysis

- Hydrolyzing salts usually come from:
  1. a strong acid + a weak base, or
  2. a weak acid + a strong base
- Strong refers to the *degree of ionization*

  A strong Acid + a strong Base = Neutral Salt
- How do you <u>know</u> if it's strong?
  - Refer to the handout provided

Salt Hydrolysis

- To see if the resulting salt is acidic or basic, check the **"parent" acid and base** that formed it. Practice on these:

$HCl + NaOH \leftarrow$ NaCl, a neutral salt

$H_2SO_4 + NH_4OH \leftarrow (NH_4)_2SO_4$, acidic salt

$CH_3COOH + KOH \leftarrow CH_3COOK$, basic salt

Buffers

- Buffers are solutions in which the pH remains relatively constant, even when small amounts of acid or base are added
  - made from a *pair of chemicals*: a *weak acid and one of it's salts*; or a *weak base and one of it's salts*

Buffers

- A buffer system is better able to resist changes in pH than pure water
- Since it is a <u>pair</u> of chemicals:
  - one chemical neutralizes any *acid* added, while the other chemical would neutralize any additional *base*
  - AND, they ***<u>produce each other</u>*** in the process!!!

Buffers

- Example: Ethanoic (acetic) acid and sodium ethanoate (also called sodium acetate)
- Examples on page 621 of these
- The buffer capacity is the amount of acid or base that can be added before a significant change in pH

Buffers

- The two buffers that are crucial to maintain the pH of human blood are:
  1. carbonic acid (H_2CO_3) & hydrogen carbonate (HCO_3^{1-})
  2. dihydrogen phosphate ($H_2PO_4^{1-}$) & monohydrogen phoshate (HPO_4^{2-})
  - Table 19.10, page 621 has some important buffer systems
  - Conceptual Problem 19.2, p. 622