

## ACID

*acidus*: Latin for "sour" or "tart"

Common acids: vinegar, battery acid, citrus fruits, carbonated beverages

### Properties

- 1) Have sour taste
- 2) Change color of certain organic dyes called indicators Ex: Litmus - Red
- 3) React with active metals to give hydrogen
- 4) Neutralize basic solutions (produce a salt and water)

### Because

- ◆ Acids increase the hydrogen ion concentration,  $[H^+]$ , in a solution
- ◆ Arrhenius: acids release hydrogen ions in water
- ◆ Bronsted-Lowry: acids release protons in water

### Examples:

HCl     $H_2SO_4$      $HNO_3$      $H_3PO_4$   
 $HC_2H_3O_2$      $H_2SO_3$

## BASES

alkaline or basic solutions

### Properties

- 1) Have slippery feel (caustic: corrodes living tissue)
- 2) Taste bitter
- 3) Change the color of indicators  
Litmus - Blue
- 4) React with  $Fe^{+3}$ ,  $Mg^{+2}$  to form a precipitate.
- 5) Neutralize acids (produce a salt and water)

### Because

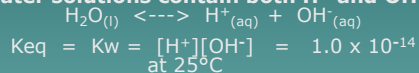
- ◆ Bases increase hydroxide ion concentration,  $[OH^-]$ , in a solution
- ◆ Arrhenius: bases release hydroxide ions in water
- ◆ Bronsted-Lowry: bases accept protons in water

### Examples

NaOH    KOH     $Ba(OH)_2$     LiOH  
 $Ca(OH)_2$      $NH_3$

All water solutions are ACIDIC,  
BASIC or NEUTRAL

All water solutions contain both  $H^+$  and  $OH^-$



### 3 Types of Water Solutions

1) Neutral - pure water  $[H^+] = [OH^-]$

$$[H^+] = 1.0 \times 10^{-7}$$

$$[OH^-] = 1.0 \times 10^{-7}$$

2) Acidic  $[H^+] > [OH^-]$

$$[H^+] > 1.0 \times 10^{-7}$$

$$[OH^-] < 1.0 \times 10^{-7}$$

3) Basic  $[H^+] < [OH^-]$

$$[H^+] < 1.0 \times 10^{-7}$$

$$[OH^-] > 1.0 \times 10^{-7}$$

Ex: A solution has a hydrogen ion concentration of  $1 \times 10^{-5}$  M. What is the concentration of hydroxide ions?

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

$$[OH^-] = \frac{1 \times 10^{-14}}{1 \times 10^{-5}}$$

$$[OH^-] = 1 \times 10^{-9}$$

Ex: A solution has a hydroxide ion concentration of  $2.0 \times 10^{-7}$ .

a) What is the  $[H^+]$ ?

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

$$[H^+] = \frac{1 \times 10^{-14}}{2 \times 10^{-7}}$$

$$[H^+] = 0.50 \times 10^{-7}$$

$$= 5.0 \times 10^{-8} \text{ M}$$

b) Is the solution acidic, basic or neutral?

$$[H^+] < [OH^-] \quad \text{Basic!}$$

## pH Scale

- In most solutions  $[H^+]$  is quite small
- Expressed in terms of pH
 
$$pH = -\log[H^+] = -\log[H_3O^+]$$
- logarithmic scale...
  - $[H^+]$  changes 10x
  - pH changes 1 unit

Most pH values are between 0 and 14

pH = 7.00 solution is neutral

pH < 7.00 solution is acidic

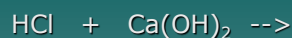
$$[H^+] > 1.0 \times 10^{-7}$$

pH > 7.00 solution is basic

$$[H^+] < 1.0 \times 10^{-7}$$

## NEUTRALIZATION

The reaction between equal amounts of  $H^+$  and  $OH^-$  to produce water  
Acidic and basic properties disappear



## TITRATION

Chemical procedure used to determine the concentration of an unknown solution using a standard solution

(a solution with a known concentration)

<https://www.youtube.com/watch?v=sEpECPTDv2w>

1. Fill a *clean* buret with the standard solution
  - a) Scrub buret with buret brush
  - b) Rinse well with distilled water
  - c) Rinse with small amount of titrant
  - d) Don't forget to run some through and fill tip of buret
2. Place a measured amount of the unknown solution in a flask  
Add a few drops of indicator
3. Slowly release standard solution into the flask until endpoint is reached
  - a) Swirl flask as you go
  - b) Occasionally rinse sides of flask with distilled water to wash down any standard solution on the sides
  - c) Endpoint is when **faint**, lingering color is noted

## Calculations

1. At the endpoint: moles  $OH^-$  = moles  $H^+$

2. Moles of each of these can be determined:

molarity \* number of  $H^+$  or  $OH^-$  \* volume (liters) = moles

Normality:

$$N_a * V_a = N_b * V_b$$

Ex: 27.4ml of standard  $Ba(OH)_2$  is added to 20.0 ml of unknown HCl solution. If the concentration of  $Ba(OH)_2$  is 0.0154M, what is the concentration of the unknown acid?

$$N_a V_a = N_b V_b$$

$$N_a(20.0ml) = (.0154M \times 2)(27.4ml)$$

$$N_a = \frac{(.0154M \times 2)(27.4ml)}{20.0ml} = 0.0422N$$

\*\*Since HCl has only one  $H^+$ , this is also 0.0422M

(If the acid was  $H_2SO_4$  the molarity would be 0.0211M)

## BUFFERS

The body needs to maintain a constant pH (7.35-7.45)  
 You constantly ingest large amounts of acid/base  
 Lots of acidity from the CO<sub>2</sub> produced during respiration



- ◆ **buffer** - a solution containing a **weak acid** + a **salt of that acid**

base

Ex: **citric acid** + **sodium citrate**  
 reacts with added base                      reacts with added acid

Ex: (blood buffer) **carbonic acid** + **sodium bicarbonate**

- ◆ **Indicators are organic dyes that change color in different pH ranges. The use of indicators to characterize acids and bases dates to the discovery of litmus in the 1500's.**

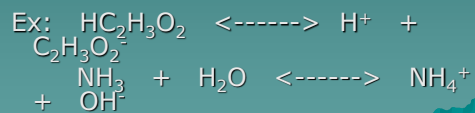
### Strong Acids/Bases

- ◆ every molecule splits



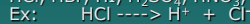
### Weak Acids/Bases

- ◆ only a few molecules split



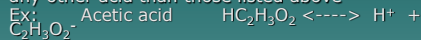
#### Strong Acids

- ◆ dissociates 100% in solution
- ◆ HCl, HBr, HI, H<sub>2</sub>SO<sub>4</sub>, HNO<sub>3</sub>, HClO<sub>4</sub>



#### Weak Acids

- ◆ any other acid than those listed above



#### Strong Bases

- ◆ dissociates 100% in solution
- ◆ hydroxides of alkalai metals (LiOH, NaOH, etc)
- ◆ Ca(OH)<sub>2</sub>, Sr(OH)<sub>2</sub>, Ba(OH)<sub>2</sub>

#### Weak Bases

- ◆ any other base than those listed above
- ◆ most common - ammonia, NH<sub>3</sub>