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^-

$H_2O_{(l)}$ <---> $H^+_{(aq)}$ + $OH^-_{(aq)}$

Keq = Kw = $[H^+][OH^-]$ = 1.0×10^{-14} at 25°C

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 1×10^{-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 2.0×10^{-7}.

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

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

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

$$[H^+] = 0.50 \times 10^{-7}$$
$$= 5.0 \times 10^{-8}\ M$$

b) Is the solution acidic, basic or neutral?

$[H+] < [OH-]$ 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

$HCl + NaOH \rightarrow H_2O + NaCl$

$HCl + Ca(OH)_2 \rightarrow$

$HNO_3 + KOH \rightarrow$

$H_3PO_4 + LiOH \rightarrow$

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=sFpFCPTDv2w

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?

$$N_aV_a = N_bV_b$$

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

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

**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_2 produced during respiration

CO_2 + H_2O -----> H_2CO_3
carbonic acid

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

Ex: HCl ----> H^+ + Cl^-

Weak Acids/Bases
- only a few molecules split

Ex: $HC_2H_3O_2$ <------> H^+ + $C_2H_3O_2^-$
NH_3 + H_2O <------> NH_4^+ + OH^-

Strong Acids
- dissociates 100% in solution
- HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_4$
 Ex: HCl ----> H^+ + Cl^-

Weak Acids
- any other acid than those listed above
 Ex: Acetic acid $HC_2H_3O_2$ <----> H^+ + $C_2H_3O_2^-$

Strong Bases
- dissociates 100% in solution
- hydroxides of alkalai metals (LiOH, NaOH, etc)
- $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

Weak Bases
- any other base than those listed above
- most common - ammonia, NH_3