

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:



BASES

alkaline or basic solutions

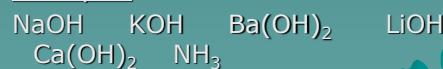
Properties

- 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



All water solutions are ACIDIC,
BASIC or NEUTRAL

All water solutions contain both H⁺ and OH⁻
 $\text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}^+_{(aq)} + \text{OH}^-_{(aq)}$
 $K_{eq} = K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$
 at 25°C

3 Types of Water Solutions

1) Neutral - pure water $[\text{H}^+] = [\text{OH}^-]$

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

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

2) Acidic $[\text{H}^+] > [\text{OH}^-]$ $[\text{H}^+] > 1.0 \times 10^{-7}$
 $[\text{OH}^-] < 1.0 \times 10^{-7}$

3) Basic $[\text{H}^+] < [\text{OH}^-]$ $[\text{H}^+] < 1.0 \times 10^{-7}$
 $[\text{OH}^-] > 1.0 \times 10^{-7}$

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

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

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

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

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

a) What is the $[\text{H}^+]$?

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

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

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

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

b) Is the solution acidic, basic or neutral?

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

pH Scale

- In most solutions $[\text{H}^+]$ is quite small
- Expressed in terms of pH

$$\text{pH} = -\log[\text{H}^+] = -\log[\text{H}_3\text{O}^+]$$

- logarithmic scale...

$[\text{H}^+]$ 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

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

pH > 7.00 solution is basic

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

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=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:

$$N_a \cdot V_a = N_b \cdot V_b$$

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?

$$\begin{aligned} N_a V_a &= N_b V_b \\ N_a (20.0\text{ml}) &= (.0154\text{M} \times 2)(27.4\text{ml}) \end{aligned}$$

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

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

(If the acid was H₂SO₄ 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 CO₂ produced during respiration



- ♦ 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) sodium bicarbonate + carbonic acid +

- ♦ 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.

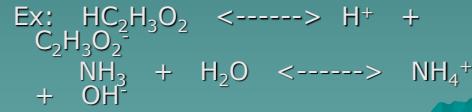
Strong Acids/Bases

- ♦ every molecule splits



Weak Acids/Bases

- ♦ only a few molecules split



Strong Acids

- ♦ dissociates 100% in solution
- ♦ HCl, HBr, HI, H₂SO₄, HNO₃, HClO₄



Weak Acids

- ♦ any other acid than those listed above



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, NH₃