ACID-BASE CHEMISTRY

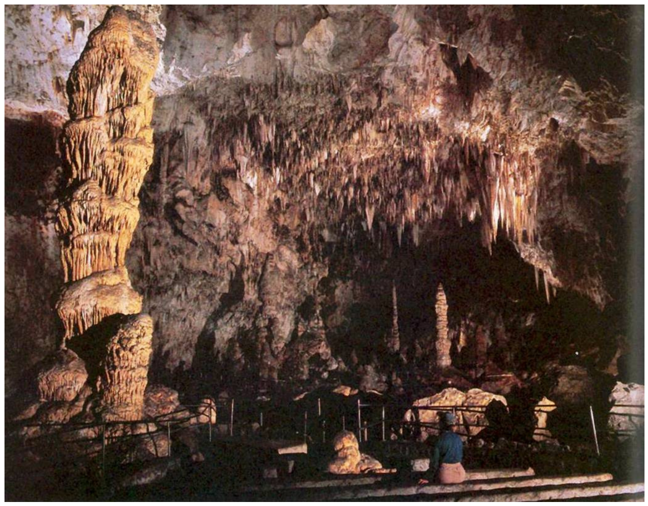
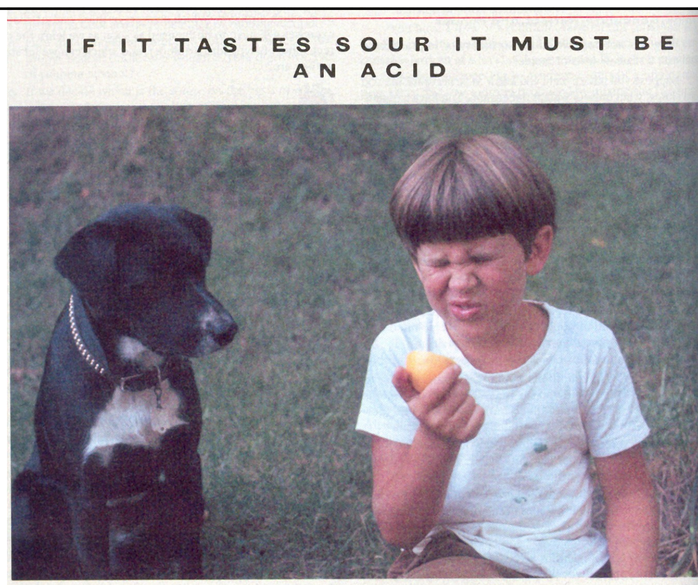


Image available at C. Snyder, “The Extraordinary Chemistry of Ordinary Things,” 4th ed. Wiley, 2003.



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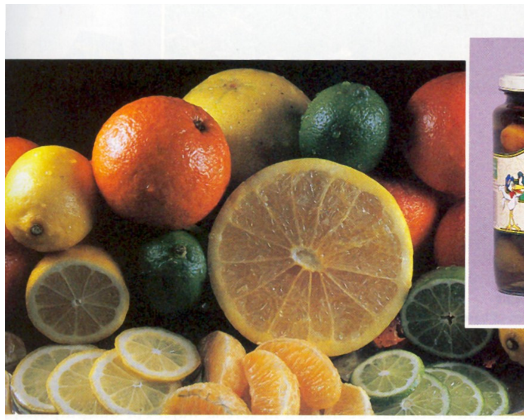
Common acid-containing materials



(b)

(c)

(d)

Examples of acids. (a) Citrus fruits – ***ascorbic*** and ***citric acids***, (b) vinegar used in cooking and preserving food – ***acetic acid***; (c) toilet bowl cleaners like Lysol – ***hydrochloric acid*** and (d) carbonated drinks – ***carbonic*** and ***phosphoric acids***.

Image available at C. Snyder, "The Extraordinary Chemistry of Ordinary Things," 4th ed. Wiley, 2003.

Acids in Everyday Life

Material	**Acid** (Name)
❖ Citrus Ex. Oranges, lemons, grapefruits	*Citric acid; Ascorbic acid or Vitamin C*
❖ Toilet bowl cleaners Ex. Lysol	*Hydrochloric acid*
❖ Vinegar Ex. Pickle juice Sweet & sour sauce	*Acetic acid*
❖ Carbonated drinks	*Carbonic acid ; Phosphoric acid*

Common base-containing materials

Examples of bases. (a) Baking soda – ***sodium bicarbonate***, (b) wood ash – ***potassium carbonate***; (c) bar soap and (d) Drain clog remover – ***sodium hydroxide.***

Image available at C. Snyder, "The Extraordinary Chemistry of Ordinary Things," 4th ed. Wiley, 2003.

Bases in Everyday Life

Material	Base (Name)
Baking powder	Sodium bicarbonate
Ash	Potassium carbonate
Glass cleaners Ex. Windex	Ammonia
Drano or Liquid Plumr Removes clogs in drains	Sodium hydroxide

What differentiates acids from bases?

Acid

Physical properties

- Tastes sour

Chemical properties

- Turns blue litmus red

  BRA => **b**lue to **r**ed = **a**cid

- Reacts with some ***metals***

  => $\mathbf{H_2}$ gas released

- React with ***carbonate*** materials

  => $\mathbf{CO_2}$ gas released

Base

Physical properties

- Tastes bitter
- Feels slippery or slimy

Chemical properties

- Turns *red* litmus *blue*

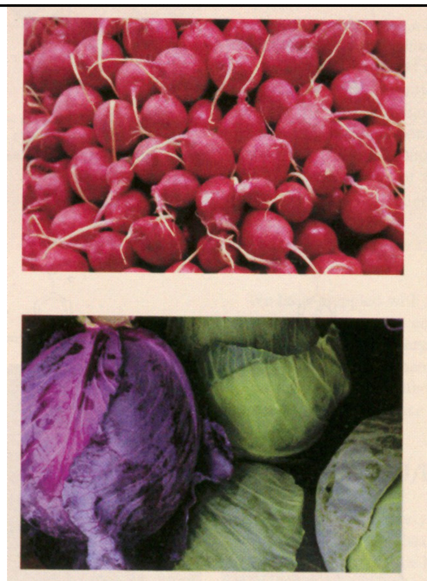
Classroom demos:
Acid-base chemistry

Acid-base indicators

Indicators = substances (like natural dyes) that *change colors* in acidic or basic (alkaline) solutions

Examples:

- Litmus
- Anthocyanins *Anthos* = flower; *Cyan* = blue
  - Red cabbage
  - Cranberries
  - Radishes



Anthocyanins give many fruits and flowers their ***stunning color*** and ***acid-base behavior***.

Image available at P. Kelter, J. Carr and A. Scott, “Chemistry: A World of Choices.” Boston: McGraw-Hill, 1999. (p. 288)

Acidic soil

Alkaline soil

Figure 11.17 Hydrangeas. These flowers are blue when grown in _____ soil and pink when grown in _____ _____ *(Diane Hirsch/Fundamental Photographs)*

Image available at C. Baird and W. Gloffke, “Chemistry In Your Life.” New York: Freeman, 2003. (p. 437)

Explaining the difference in properties of acids and bases

ACIDS = substances that ***produce H_3O^+*** (H^+ bonded to H_2O, thus simplified as **H^+**) ions in water [**Arrhenius definition**]

ACIDS = substances that ***donate H^+*** in water [**Bronsted-Lowry definition**]

Strong acids are 100 % dissociated in water

$$HA + H_2O \longrightarrow H_3O^+ + A^-$$

Or simply

$$HA_{(aq)} \longrightarrow H^+ + A^-$$

(One-sided arrow means 100 % conversion to products)

➢ Note that H_2O is omitted in the simplified dissociation, and H_3O^+ is simplified as H^+

Weak acids are only partially dissociated in water

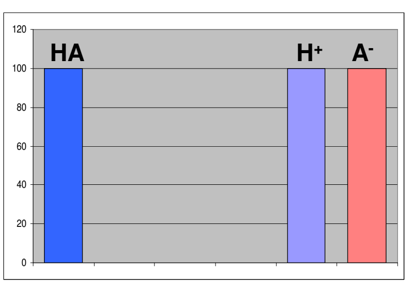
$$HA_{(aq)} \rightleftarrows H^+ + A^-$$

(Double-sided arrow means partial conversion to products)

Dissociation of strong acids and weak acids in water

Before dissociation

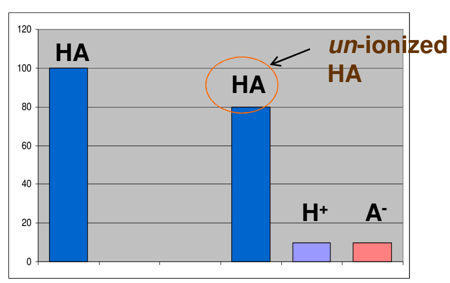
After dissociation



Strong acid

Before dissociation

After dissociation



Weak acid

➢ *Greater tendency to dissociate (ionize) = stronger acid*

Explaining the difference in properties of acids and bases

Bronsted-Lowry (B-L) Concept

At the submicroscopic level:

aqua **= water**

Acids are ***proton (H^+)*** *donors* in aqueous solutions

Hydrogen ions, H^+

Example:

$$H-Cl \;+\; H-O-H \;\Longrightarrow\; Cl^- \;+\; H-OH^+(-H)$$

Hydrochloric acid + Water produces Chloride ions + Hydronium ions

A B-L acid

(Hydrogen from the acid is donated as H^+ ion to the other species in solution)

Makes the solution acidic

Bronsted-Lowry (B-L) Concept *(Cont.)*

Bases are proton (H^+) *acceptors* in *aqueous* solutions

Example:

$$H-O-H \;+\; NH_3 \;\Longrightarrow\; NH_3^+(-H) \;+\; OH^-$$

Water + Ammonia produces Ammonium ions + Hydroxide ions

A *B-L* **acid** A *B-L* **base**

(The base accepts a H^+ ion from the other species in solution)

Explaining the difference in properties of acids and bases

BASES = substances that ***produce OH^-*** ions in water [**Arrhenius definition**]

BASES = substances that ***accept H^+*** in water [**Bronsted-Lowry definition**]

Similarly, **strong bases** are 100 % dissociated while **weak bases** are only partially dissociated in water

How can we tell if a given species is strong or weak?

STRONG ACIDS and STRONG BASES

There are only **seven (7) strong acids**: General formula HA, where A is an anion

HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_3$, $HClO_4$

First ionization *only* (H_2SO_4)

Chloric and **perchloric acid** ($HClO_3$, $HClO_4$)

There are only **seven (7) strong bases**: General formula $M(OH)_n$, where M is a Group I or II metal with charge n+

$\underbrace{LiOH, NaOH, KOH, RbOH}_{\text{Group I metal hydroxides}}$, $\underbrace{Ca(OH)_2, Sr(OH)_2, Ba(OH)_2}_{\text{Group II metal hydroxides}}$

- If it is NOT in the list above, it is weak

Review of Concepts

Exercise: Identify each of the following species as either a strong acid (SA), strong base (SB), weak acid (WA) or weak base (WB)

Species	ID
HCN	WA
$Mg(OH)_2$	WB
H_2CO_3	WA
HI	SA
NH_3	WB
$HC_2H_3O_2$	WA
NH_4^+	WA

Note: In both examples, water behaved as an acid or a base. A species that can act as an acid or a base is called ***amphoteric***.

Illustrate on the board the B-L concept

Drill: Identify the ***acid*** and the ***base*** in each of the following.

1) HNO_3 + CO_3^{2-} => NO_3^- + HCO_3^-

acid base

2) HPO_4^{2-} + H_2O => $H_2PO_4^-$ + OH^-

base acid

The Self-Dissociation (*Autoionization*) of Water

- Water is amphiprotic (it can act as a B-L acid or a B-L base)
- A molecule of H_2O can donate H^+ to another H_2O molecule

$$H_2O + H_2O \underset{K_w}{\overset{K_w}{\rightleftharpoons}} H_3O^+ + OH^-$$

Autoionization of water

Transfer of H^+

$K_w = [H_3O^+]\,[OH^-]$ At 25 ^{0}C, $K_w = 1.0 \times 10^{-14}$

The pH Scale

- The **pH scale** is a numeric scale that is used as a measure of acidity or basicity of solutions.
- Mathematically, pH is the negative logarithm (base 10) of $[H_3O^+]$

$$\mathbf{pH = -\log [H_3O^+]}$$

Recall: At 25 ^{0}C, $K_w = 1.0 \times 10^{-14}$. What is the pH of pure water at 25 ^{0}C?

Since $[H_3O^+] = 1.0 \times 10^{-7}$ M in neutral solutions like pure H_2O:

$pH = -\log [H_3O^+] = -\log (1.0 \times 10^{-7}) = -(-7.00)$

pH = 7.00 (In neutral solutions and in pure H_2O)

Relationship between $[H^+]$ and $[OH^-]$

- In most aqueous solutions, $[H_3O^+]$ and $[OH^-]$ are not equal.
- What happens to $[OH^-]$ when $[H_3O^+]$ is increased? Decreased?

Since $[H_3O^+][OH^-] = 1.0 \times 10^{-14} = K_w$ at 25 ^{0}C:

$\uparrow[H_3O^+]$, $[OH^-]$ must decrease so their product equals 1.0×10^{-14}

Thus, when $[H_3O^+]$ is very high, $[OH^-]$ is very low, and vice versa.

- In **acidic solution**, $\mathbf{[H_3O^+] > [OH^-]}$
- In **neutral solution**, $\mathbf{[H_3O^+] = [OH^-]}$
- In **basic solution**, $\mathbf{[H_3O^+] < [OH^-]}$

Relationship between pH and $[H_3O^+]$

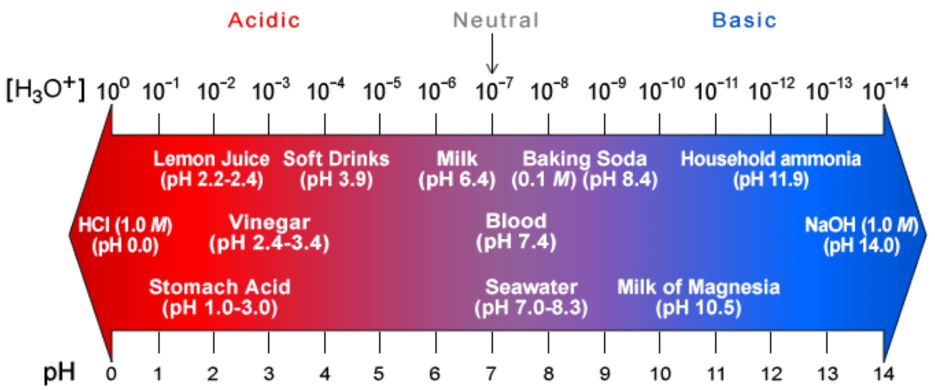


Image available at http://www.kmacgill.com/documents/pH_scale.gif

Notes: (1) The lower the pH, the more acidic the solution. (2) For every factor of ten change in $[H_3O^+]$, pH changes by one unit.

pH and pOH

pH = - log $[H_3O^+]$ **pOH = - log $[OH^-]$**

- Since $[H_3O^+][OH^-] = 1.0 \times 10^{-14} = K_w$ at 25 ^{0}C, it follows that

  $\boxed{pH + pOH = 14.00}$ at 25 ^{0}C

pH scale (cont.)

Less than 7 = Acidic
0-1 Highly acidic; highly *corrosive*
5-6 Weakly acidic
pH 7 = Neutral
Greater than 7 = Basic
8-9 Weakly basic
13-14 Strongly basic; highly *corrosive*

pH and everyday life

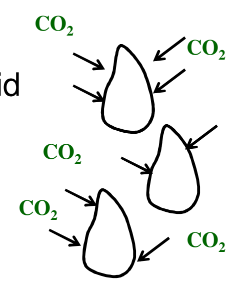
Rainwater is naturally acidic

WHY?

➢ Carbon dioxide (from the air) dissolves in rainwater, producing ***carbonic acid***, a weak acid

$$CO_2\,\textbf{(gas)} + H_2O \longrightarrow H_2CO_3\,\textbf{(aqueous)}$$

Carbonic acid



Rain drops

• Rain is "carbonated water"

• The ***normal pH*** of rainwater is about **5.6**

Acid rain is even more acidic

Acid rain is rain having a ***pH less than 5***.

➢ Forms when air-polluting gases dissolve in rainwater

- Sulfur dioxide (SO_2)
- Nitrogen dioxide (NO_2)

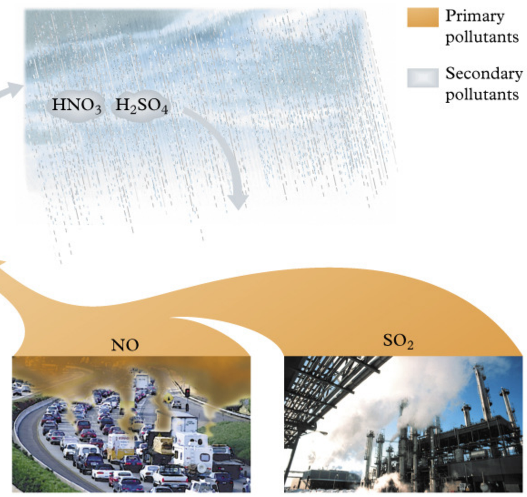
➢ Further lowers the pH of rain. WHY?

➢ These gases form stronger acids* than *carbonic acid* in rainwater → Lowers the pH to less than 5

***Sulfuric acid** and **nitric acid**

Formation of acid rain

Image available at C. Baird, “Chemistry in Your Life”. 2nd ed., New York: W.H. Freeman, 2006.



Like photochemical smog, **acid rain** is a ***secondary air pollutant*** (i.e. it forms from further reaction of primary pollutants)

Formation of acid rain

- From the *oxidation* of 1^0 pollutants sulfur dioxide and nitrogen dioxide.

The **overall reactions** for these multi-step reactions are given below:

$$SO_{2\,(g)} + 1/2\ O_{2\,(g)} + H_2O_{\,(l\ or\ g)} \longrightarrow H_2SO_{4\,(aq)}$$

Sulfur dioxide **Sulfuric acid**

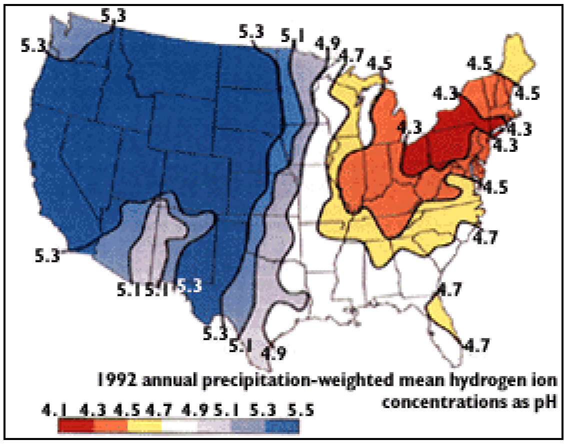
$$NO_{2\,(g)} + 1/2\ O_{2\,(g)} + H_2O_{(l\ or\ g)} \longrightarrow 2HNO_{3\,(aq)}$$

Nitrogen dioxide **Nitric acid**

Acid Rain and Geography

➢ Acidic precipitation is more abundant in the **northeastern U.S.** - a consequence of air mass movement (and geographic location)

➢ Even areas downwind from where acid-producing gases are generated can be burdened by the effects of acid rain



http://pubs.usgs.gov/gip/acidrain/2.html

Effects of Acid Rain

Damages lakes and aquatic life

http://www.sciencemaster.com/jump/earth/acid_rain.php

In some sensitive lakes and streams, ***acidification has completely eradicated fish species***, such as the **brook trout**, leaving these bodies of water barren. In fact, hundreds of the lakes in the **Adirondacks** surveyed in the NSWS have acidity levels indicative of chemical conditions unsuitable for the survival of sensitive fish species.

Effects of Acid Rain

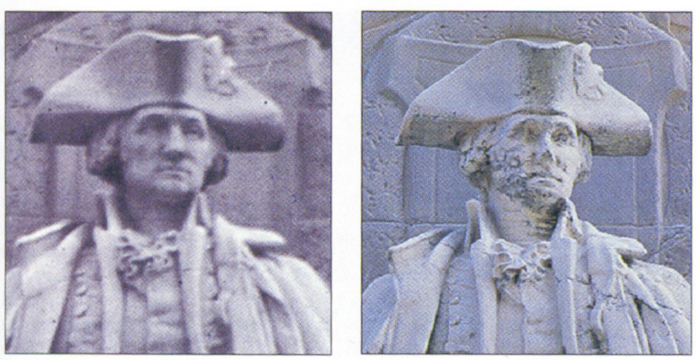
Damage to vegetation (destruction of sensitive forests)

Image available at http://www.sciencemaster.com/jump/earth/acid_rain.php

Effects of Acid Rain

Damage to limestone statues

Neutralizing the Threat of Acid Rain



In 1944 — At present

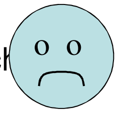
Figure 6.15

Acid rain can damage limestone statuary. This statue of George Washington was first put outside in New York City in 1944. During the next 58 years, acid rain caused significant damage to the statue.

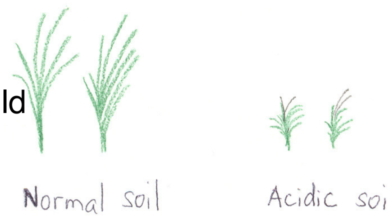
Source: American Chemical Society, "Chemistry In Context." 4th ed. C. Stanitski et al (Eds.) Boston: McGraw-Hill, 2003 (p. 263)

What if things get out of control?

- Hyperacidic stomach



- Highly acidic soil => low crop yield



- Air pollutants => acid rain
  - Acidic lakes/rivers => fish kills; erosion of statues; vegetation dies

➢ ***There must be a way to control pH***

Ways of Controlling pH

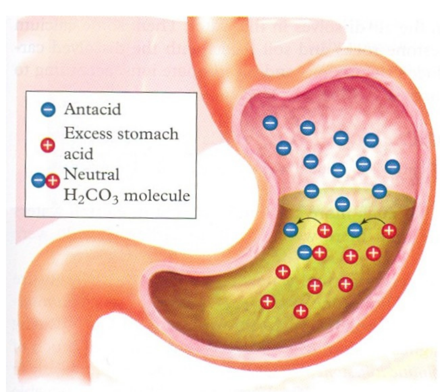
1) Through ***neutralization*** reactions = acid-base reactions

Ex. **HCl** + **antacid** (a carbonate) = > **salt** + **H_2O**

Excess stomach acid

A base

Always produced in neutralization



Source: C. Baird and W. Gloffke, "Chemistry In Your Life." New York: Freeman, 2003. (p. 427)

Neutralization Reactions

HCl + $NaOH$ ⟶ $\underline{NaCl}$ + $\underline{HOH}$ (or H_2O)
salt

HNO_3 + KOH ⟶ $\underline{KNO_3}$ + $\underline{H_2O}$
salt

HCN + $NaOH$ ⟶ $\underline{NaCN}$ + $\underline{H_2O}$
salt

Notes: 1) The metal from the base is always *written* (or *named*) *first*.

2) The metal from the base ***replaces*** the hydrogen of the acid.

Ways of Controlling pH (Cont.)

2) Through the action of ***buffers***

- Substances that resist drastic changes in pH. HOW?
- They consist of a mixture of a ***weak acid*** and its conjugate ***base*** (i.e an acid-base pair).

Related in structure

Ex. The pH of blood is maintained by a biological buffer, a mixture of ***carbonic acid*** and ***carbonate ions***.

acid component   **base component**

Ways of Controlling pH: *Action of Buffers* (Cont.)

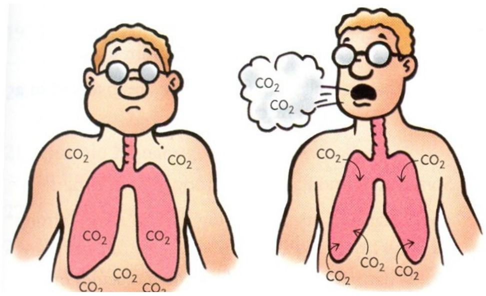


Image available at C. Snyder, "The Extraordinary Chemistry of Ordinary Things," 4th ed. Wiley, 2003.

- During **acidosis** (blood pH drops), the base component (carbonate) neutralizes the excess acid and restores the pH to around 7.4.
- During alkalosis (blood pH rises) the acid component (carbonic acid) of our *biological buffer* neutralizes the excess base and restores the pH to around 7.4, the normal pH of blood.