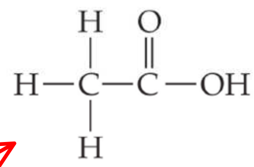
Chapter 14: Acids and Bases

Properties of Acids

Sour taste
React with some metals
Turns blue litmus paper red
React with bases

Some Common Acids

HCl, hydrochloric acid
H_2SO_4, sulfuric acid
HNO_3, nitric acid
$HC_2H_3O_2$, acetic acid
(common name)



What is the IUPAC name of this carboxylic acid?

Properties of Bases

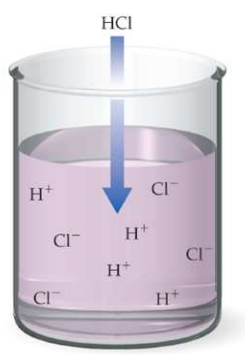
Taste bitter
Solutions feel slippery
Turns red litmus paper blue
React with acids

Some Common Bases

NaOH, sodium hydroxide
KOH, potassium hydroxide
$NaHCO_3$, sodium bicarbonate

Arrhenius Acid-Base Theory

Arrhenius Acid - a hydrogen containing compound that ionizes to produce hydrogen ions (H^+) when dissolved in water



$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

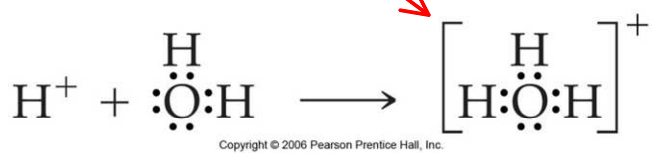
Because **molecular acids are not made of ions**, they cannot dissociate.

They must be pulled apart, or **ionized**, by the water.

Arrhenius Acids and Bases

Hydrogen ions released from acids do not float about freely in solution. Instead, they attach to water molecules to form

HYDRONIUM IONS



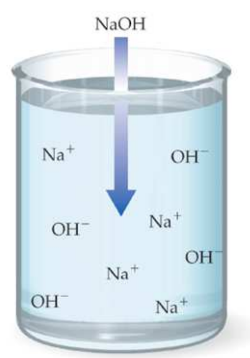
$= H_3O^+$

In the formula for an acid, the *hydrogen atom that ionizes* is many times <u>written first</u> - ex. $HC_2H_3O_2$

This hydrogen atom is the one that ionizes

These hydrogen atoms do not ionize in water

<u>Arrhenius Base</u> - a hydroxide containing compound that dissociates to produce OH^- when dissolved in water



Bases that contain OH^- are ionic compounds.

Ionic substances **dissociate** in water.

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

Arrhenius Acid–Base Reactions

The H^+ from the acid combines with the OH^- from the base to make a molecule of H_2O

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

(acid + base → salt + water)

The cation from the base combines with the anion from the acid to make a salt (*an exchange reaction*)

Bronsted-Lowry Acid-Base Theory

Although widely used, the Arrhenius acid-base theory has limitations. (For example, some compounds act as bases even though they do not contain OH^-).

- Bronsted-Lowry theory broader definition than Arrhenius theory

Bronsted-Lowry Acid - any substance that can donate a proton (H+)

Bronsted-Lowry Base - any substance that can accept a H^+

A Brønsted-Lowry *acid–base reaction* is any reaction in which an **H^+ is transferred**.

- The acid molecule donates an H^+ to the base molecule:

$$H–A \quad + \quad :B \longrightarrow \quad :A^- \quad + \quad H–B^+$$

Base structure must contain an atom with an unshared pair of electrons to bond to H^+.

Bronsted-Lowry Acid-Base reactions don't have to take place in aqueous solution.

Examples:

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

Proton donor Proton acceptor

$$HCl_{(aq)} \quad + \quad NH_{3(aq)} \longrightarrow NH_4^+{}_{(aq)} \quad + \quad Cl^-_{(aq)}$$

Proton donor Proton acceptor

Conjugate Acid-Base Pairs

Identify the Bronsted-Lowry acid and base in the following reactions:

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

$$HCO_3^-{}_{(aq)} + H_2O_{(l)} \longrightarrow CO_3^{2-}{}_{(aq)} + H_3O^+{}_{(aq)}$$

Conjugate Acids and Bases

A conjugate acid-base pair is *two species that differ from each other by one proton*

A = acid; CB = conj. base

B = base; CA = conj. acid

$$\underset{A}{HCl_{(aq)}} + \underset{B}{H_2O_{(l)}} \longrightarrow \underset{CB}{Cl^-{}_{(aq)}} + \underset{CA}{H_3O^+{}_{(aq)}}$$

$$\underset{B}{NH_2^-{}_{(aq)}} + \underset{A}{H_2O_{(l)}} \longrightarrow \underset{CA}{NH_{3(aq)}} + \underset{CB}{OH^-{}_{(aq)}}$$

Identify the conjugate acid-base pairs in the following reactions:

$$HF_{(aq)} + H_2O_{(l)} \longrightarrow F^-{}_{(aq)} + H_3O^+{}_{(aq)}$$

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

We will not cover Section 14.5 "Reactions of Acids and Bases" **(we already discussed part of this in Ch.7)** **or Section 14.6 "Acid-Base Titration"**

Note: Water is an amphoteric molecule - it can function as an acid or a base

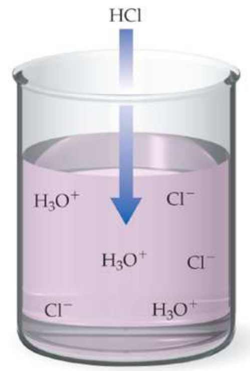
Strong and Weak Acids and Bases

A **strong acid** is an acid that transfers 100% (or nearly 100%) of its acidic hydrogens to water

100% transfer

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

Hydrochloric acid: **a strong acid**



Common Strong Acids
HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$

A **weak acid** is an acid that transfers only a small amount of its acidic hydrogens to water

$$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

Acetic acid: **a weak acid**

Common Weak Acids
$HC_2H_3O_2$, HF, H_2CO_3, H_3PO_4, H_2SO_3

When acetic acid is dissolved in water only **0.42%** of its acidic protons are transferred to H_2O. 99.58% of the molecules remain un-ionized.

Bases can also be **strong** or *weak*

100% dissociation

Only small amount of dissociation

Common Strong Bases
$LiOH$, $NaOH$, KOH, $Ca(OH)_2$

$$NaOH \rightarrow Na^+ + OH^-$$

Dissociates 100%

Common Weak Bases

NH_3, HCO_3^-, CH_3NH_2

(ammonia, bicarbonate, methyl amine)

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

When NH_3 reacts with H_2O only a **small amount of OH^- produced**

Self Ionization of Water

A VERY SMALL PERCENTAGE (about 1 out of every 10 million) of water molecules can dissociate into H_3O^+ and OH^-

$$H_2O \rightleftharpoons H^+ + OH^- \quad \text{(Arrhenius)}$$

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^- \quad \text{(Bronsted-Lowry)}$$

All aqueous solutions contain both H_3O^+ and OH^-

At 25°C, there are equal numbers of hydronium and hydroxide ions:

$$\boxed{\begin{array}{c}[H_3O^+] = [OH^-] = 1 \times 10^{-7} M \\ \text{at } 25\ °C \text{ in pure water}\end{array}}$$

Note: [brackets] mean molar concentration in H_2O

$[H_3O^+] \times [OH^-]$ = Ion Product for Water (K_w)

$$\boxed{[H_3O^+] \times [OH^-] = 1 \times 10^{-14}\ M}$$

The product of the H_3O^+ and OH^- concentrations is constant: **1×10^{-14} M**

If **H_3O^+ ions are added** to water, there will be a **decrease in $[OH^-]$**

If **OH^- ions are added** to water, there will be a **decrease in $[H_3O^+]$**

Sufficient acid is added so that the $[H_3O^+]$ is now 6.52×10^{-4} M. What is the $[OH^-]$ in this solution?

Sufficient base is added so that the $[OH^-]$ reaches 5.41×10^{-6}M. What is the new $[H_3O^+]$?

The pH Scale

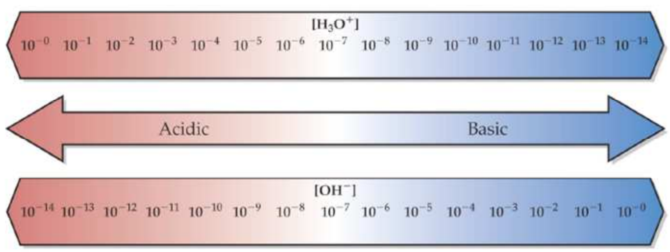
Acidic Solutions: **$[H_3O^+] > [OH^-]$**

Basic Solutions: **$[H_3O^+] < [OH^-]$**

Neutral Solutions: **$[H_3O^+] = [OH^-]$**

A **strongly acidic** solution has a relatively high concentration of __________.

A **strongly basic** solution has a relatively high concentration of __________.



The **pH scale** is a 0-14 scale which measures the acidity or basicity of solution.

- The pH scale was derived because *very small concentrations* (such as 4.22 x 10^{-11} M and 3.12 x 10^{-12}M) *are difficult to compare.*

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

The pH of a solution is the **negative logarithm** of the hydronium ion concentration

pH = 2.0 means $[H_3O^+]$ = 1 x 10^{-2} *M*

pH = 3.0 means $[H_3O^+]$ = 1 x 10^{-3}M

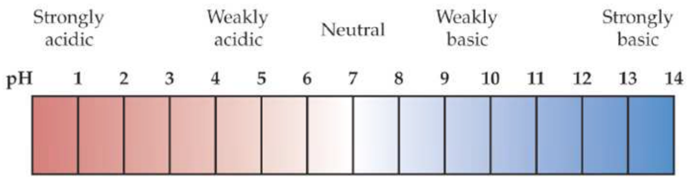
If $[H_3O^+]$ = 1 x 10^{-9} *M*, then pH = _____

Which is **more acidic**, pH = 2.0 or pH = 3.0?

pH Calculations

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

pH < 7 acidic solution
pH > 7 basic solution
pH = 7 neutral solution



The *lower the pH*, the more acidic the solution; the *higher the pH*, the more basic the solution.

- 1 pH unit corresponds to a **factor of 10 difference** in acidity.
- **Normal range** is 0 to 14
  pH 0 then $[H^+] = 1$ M, pH 14 then $[OH^-] = 1$ M; but pH can be negative (very acidic) or larger than 14 (very alkaline)

If $[H_3O^+] = 1 \times 10^{-6}$ *M*, then pH = $-\log 10^{-6}$ =

If $[H_3O^+] = 1.0 \times 10^{-12}$ *M*, pH = ________ acidic or basic?

If $[H_3O^+] = 3.6 \times 10^{-12}$ *M*, pH = ________ acidic or basic?

Text calculator: $\boxed{(-)}$ $\boxed{\text{LOG}}$ 3.6 $\boxed{\text{E}}$ $\boxed{(-)}$ 12 $\boxed{=}$

Numerical calculator: 3.6 $\boxed{\text{E}}$ 12 $\boxed{^+/\!-}$ $\boxed{\text{LOG}}$ $\boxed{^+/\!-}$

(in pH values, the sig figs are **after the decimal point**)

pH and H_3O^+ Calculations

Calculate the pH of the following solutions:

a) $[H_3O^+] = 5.9 \times 10^{-10}$ b) $[H_3O^+] = 2.5 \times 10^{-3}$ c) $[H_3O^+] = 6.32 \times 10^{-6}$

Conversions between pH and $[H_30^+]$

If pH = 8.0, $[H_3O^+] = 1 \times 10^{-8}$ M

$$\boxed{[H_3O^+] = 10^{-pH}}$$

If pH = 2.87, $[H_3O^+] = 10^{-2.87}$ = __________

text calculator: $\boxed{10^x}$ $\boxed{(-)}$ 2.87 $\boxed{=}$

numerical calculator: 2.87 $\boxed{^+/_-}$ $\boxed{10^x}$

If pH = 6.43, $[H_3O^+] = 10^{\boxed{}}$ = ________

(Use MODE or SCI if your calculator gives you 0.00000…)

Calculate the $[H_30^+]$ of the following solutions:

a) pH = 8.92 b) pH = 2.664

$[OH^-]$ and pOH Calculations

If $[OH^-] = 1 \times 10^{-4}\ M$, pOH = 4

$\boxed{pOH = -\log [OH^-]}$ $\boxed{[OH^-] = 10^{-pOH}}$

pOH and pH are related

$\boxed{pH + pOH = 14}$

Calculate the pH of the following solutions:

a) $[OH^-] = 1.0 \times 10^{-5}$ b) $[OH^-] = 6.7 \times 10^{-8}$ c) $[OH^-] = 6.32 \times 10^{-6}$

These equations will be given on the final exam:

$pH = -\log [H_3O^+]$

$[H_3O^+] = 10^{-pH}$

$pH + pOH = 14$

$pOH = -\log [OH^-]$

$[OH^-] = 10^{-pOH}$

We will not cover Sections 14. 10 "Buffers" or 14.11 "Acid Rain" (this material will not be on the final exam)

TABLE 14.7 The pH of Some Common Substances

Substance	pH
gastric (human stomach) acid	1.0–3.0
limes	1.8–2.0
lemons	2.2–2.4
soft drinks	2.0–4.0
plums	2.8–3.0
wine	2.8–3.8
apples	2.9–3.3
peaches	3.4–3.6
cherries	3.2–4.0
beer	4.0–5.0
rainwater (unpolluted)	5.6
human blood	7.3–7.4
egg whites	7.6–8.0
milk of magnesia	10.5
household ammonia	10.5–11.5
4% NaOH solution	14