Ch 14. Acids and Bases

Acids and Bases

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Acids and Bases

- Acids
  - Sour taste
  - Dissolve many metals
  - Turn litmus paper red.
  - Egs. Acetic acid (vinegar), citric acid (lemons)

- Bases
  - Bitter taste, slippery feel
  - Turn litmus paper blue.
  - Egs. Drano, ammonia, caffeine

Arrhenius Definition of Acids and Bases

Based on H^+ or OH^- Production in Water

- Acid: Substance that produces H^+ ions (protons) in aqueous solutions
  - $HCl\ (aq) \rightarrow H^+\ (aq) + Cl^-\ (aq)$
  - $HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$

- Base: Substance that produces OH^- (hydroxide) ions aqueous solutions
  - $NaOH\ (aq) \rightarrow Na^+(aq) + OH^-(aq)$

- Defn is limited– restricted to reactions in water only

Bronsted-Lowry Definition of Acids and Bases

Based on Proton Transfer:

- Acids: Proton donors

  $HCl \rightarrow H^+ + Cl^-$

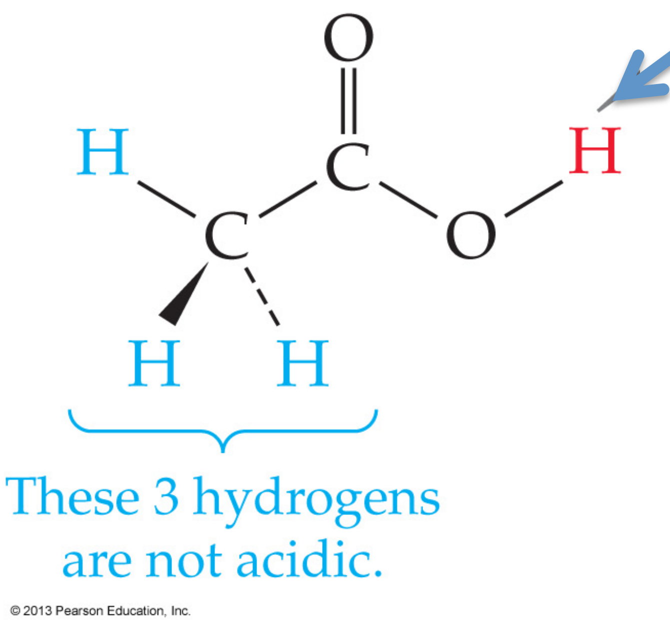
- Bases: Proton acceptors

  $NH_3 + H^+ \rightarrow NH_4^+$

Bronsted-Lowry Acid

- An acid must have an acidic proton that can be donated.

acidic proton

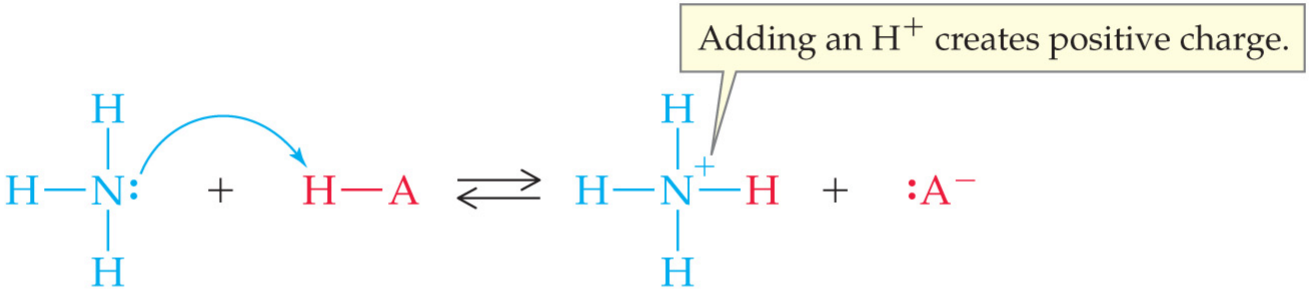


Bronsted-Lowry Base

- A base must have a lone pair to accept a proton.
- A base can be neutral or negatively charged.

Hydroxide ion
(a negatively charged base, B:⁻)

Water



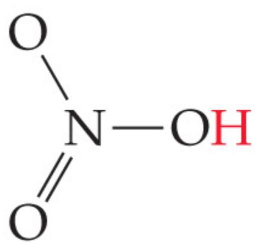
Ammonia
(a neutral base, B:)

Ammonium ion

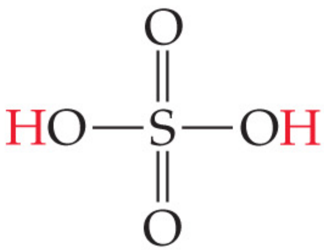
Number of Acidic Protons on a Molecule

- An acid molecule can be monoprotic, diprotic, or triprotic (based on number of acidic protons).

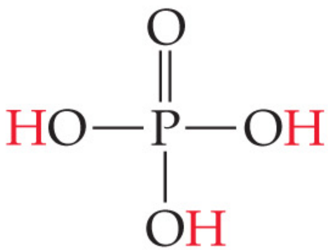
Hydrochloric acid
(monoprotic)



Nitric acid
(monoprotic)



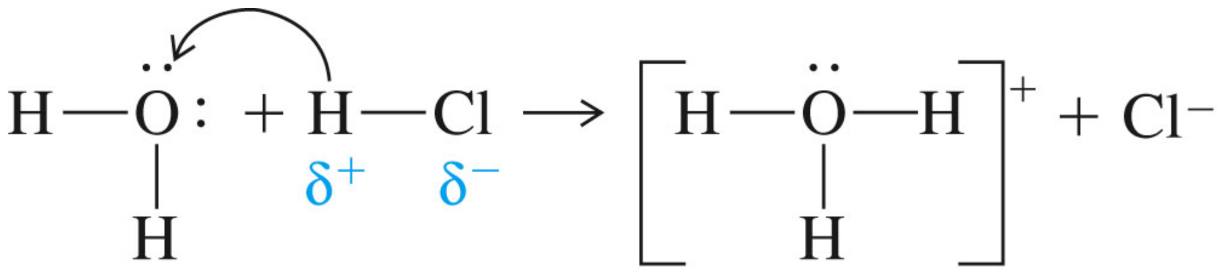
Sulfuric acid
(diprotic)



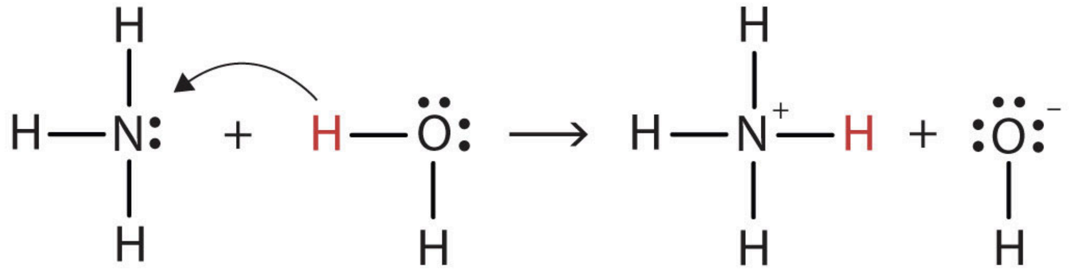
Phosphoric acid
(triprotic)

Water Acts as Both Acid and Base (Amphoteric)

- Water can act as a base, accepting a proton from an acid → Forms hydronium ion H_3O^+.



- Water can act as an acid, donating a proton to a base → Forms hydroxide ion OH^-

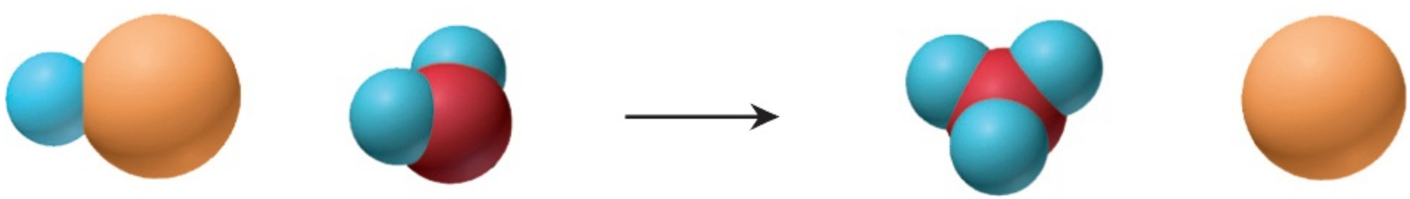


The Proton in Water

- A H^+ ion in water is not isolated! A H^+ ion in water attracts the negative pole of water molecule so strongly that it forms a covalent bond to water.

$$HCl\,(aq) + H_2O(l) \rightarrow \underline{H_3O^+}\,(aq) + Cl^-(aq)$$

hydronium ion (hydrated proton)



*Also written as: $HCl\,(aq) \rightarrow H^+\,(aq) + Cl^-(aq)$

Bronsted-Lowry Acid and Base Work Together

- An acid and base must always work together to transfer a proton!

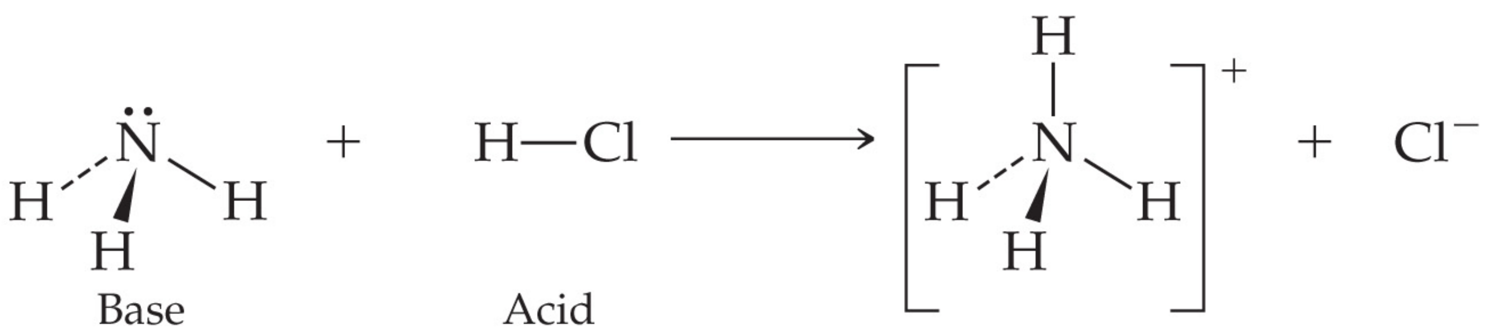
In aqueous solutions of acids or bases, water acts as the base or acid "partner":

- $HCl + \underline{H_2O} \rightarrow Cl^- + H_3O^+$ $(HCl \rightarrow Cl^- + H^+)$
  acid base

- $NH_3 + \underline{H_2O} \rightarrow NH_4^+ + OH^-$ $(NH_3 \rightarrow NH_4^+ + OH^-)$
  base acid

Acid-Base Reactions Don't Have to Involve Water

- A Bronsted-Lowry acid-base reaction does not always occur in water.



Ex Probs

List of Common Acids and Bases

Strong Acids HCl hydrochloric acid HBr hydrobromic acid HI hydroiodic acid HNO_3 nitric acid H_2SO_4 sulfuric acid $HClO_4$ perchloric acid	Strong Bases: Metal Hydroxides of Group 1A cations LiOH, NaOH, KOH, etc. *and* Heavier Group 2A cations: $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$
Common Weak Acids $HC_2H_3O_2$ or CH_3COOH acetic acid H_2CO_3 carbonic acid	Common Weak Base NH_3 ammonia

Conjugate Acid-Base Pairs

Conjugate Acid-Base Pairs

Conjugate Acid-Base Pair

Conjugate acid-base pair: two substances whose formulas differ by only a hydrogen ion

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Base | Acid | Conjugate acid | Conjugate base

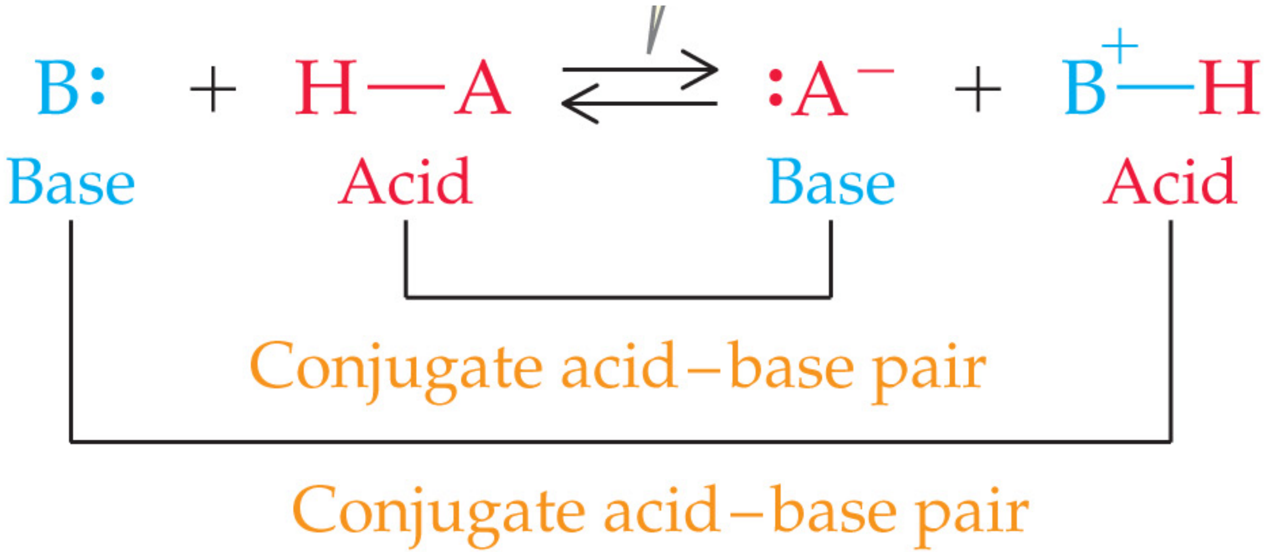
- NH_3 and NH_4^+ are a conjugate acid-base pair.
- H_2O and OH^- are a conjugate acid-base pair.

Conjugate Acid-Base Pair

$$\underset{\text{Acid}}{H_2SO_4(aq)} + \underset{\text{Base}}{H_2O(l)} \longrightarrow \underset{\text{Conjugate base}}{HSO_4^-(aq)} + \underset{\text{Conjugate acid}}{H_3O^+(aq)}$$

- H_2SO_4 and HSO_4^- are a conjugate acid-base pair.
- H_2O and H_3O^+ are a conjugate acid-base pair.

Conjugate Acid-Base Pair



Example Problems

Which of the following represent conjugate acid-base pairs?

a. HF, F^-
b. NH_4^+, NH_3
c. HCl, H_2O
d. $HClO_4$, ClO_4^-
e. HCl, ClO^-
f. $H_2PO_4^-$, HPO_4^{2-}
g. HNO_3, NO_3^-

Ex Probs

Reactions of Acids and Bases

Reactions of Acids and Bases

Acid-Base Reaction: Neutralization

Neutralization: a reaction in which an acid and base react quantitatively with each other

Acid-Base Reaction: Acid + Strong Base

When an acid and a strong base (metal hydroxide) react, they form <u>water</u> and a <u>salt</u> (ionic compound).

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

Net ionic eqn: $H^+(aq) + OH^-(aq) \rightarrow H_2O\ (l)$

Acid-Base Reaction: Gas-Forming

When an acid and a bicarbonate (HCO_3^-) or carbonate (CO_3^{2-}) react, water and gas are formed.

- $HCl(aq) + KHCO_3\mathit{(aq)} \rightarrow KCl(aq) + [H_2CO_3\mathit{(aq)}] \rightarrow$
  $KCl + H_2O\mathit{(l)} + CO_2(g)$

  $H^+(aq) + HCO_3^-\mathit{(aq)} \rightarrow [H_2CO_3\mathit{(aq)}] \rightarrow H_2O\mathit{(l)} + CO_2(g)$

- $2HCl(aq) + K_2CO_3\mathit{(aq)} \rightarrow 2KCl(aq) + [H_2CO_3\mathit{(aq)}] \rightarrow$
  $2KCl(aq) + H_2O\mathit{(l)} + CO_2(g)$

  $2H^+(aq) + CO_3^{2-}\mathit{(aq)} \rightarrow [H_2CO_3\mathit{(aq)}] \rightarrow H_2O\mathit{(l)} + CO_2(g)$

Reaction of Acids with Metals (Redox Rxn)

- Acids dissolve many metals: Acids oxidize metals, causing metals to go into solution.

- $2HCl(aq) + Mg(s) \rightarrow H_2(g) + MgCl_2(aq)$

  acid metal H_2 gas salt

- $H_2SO_4(aq) + Zn(s) \rightarrow H_2(g) + ZnSO_4(aq)$

  acid metal H_2 gas salt

Titration

- Titration: Determination of the concentration of a solution with an unknown concentration (analyte) by combining it with a standard solution of known concentration (titrant).

Acid-Base Titration

- Equivalence point: the point where enough titrant has been added to react exactly with the analyte present

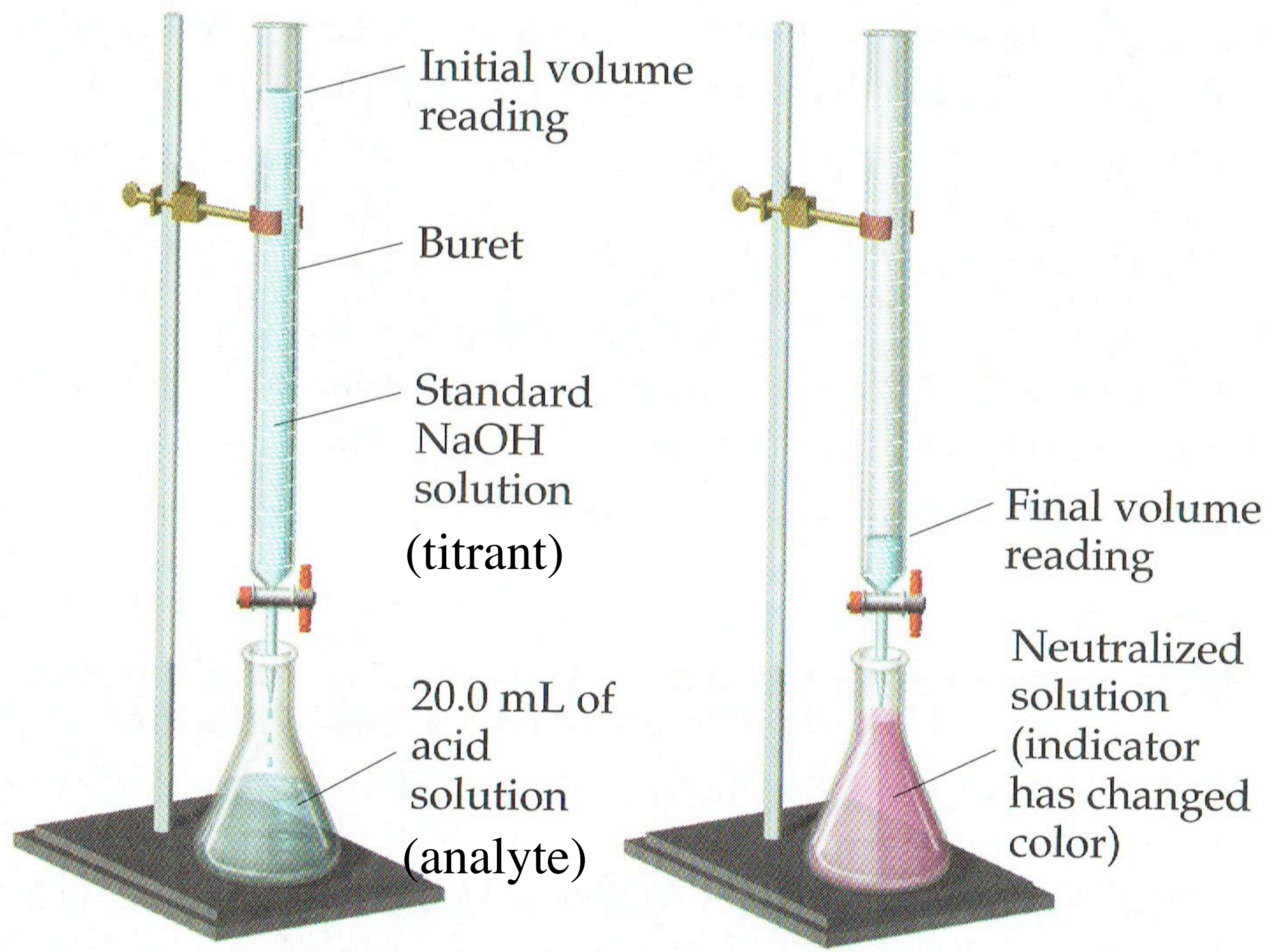
$$\boxed{\text{mole } H^+ = \text{mole } OH^-}$$

- Indicator for acid-base titration: Compound whose color is different in acid than in base

- Endpoint: the point at which the indicator's signal is triggered (The endpoint may or may not come exactly at equivalence point).

Acid-Base Titration

$CH_3COOH(aq) + NaOH(aq) \rightarrow H_2O(l) + Na^+(aq) + CH_3COO^-(aq)$



Acid-Base Titration

- Acid-base titration problems are stoichiometry problems (solution stoichiometry).

Ex Probs

Acid Strength

Acid Strength

- Strong acid: Completely ionized (dissociated).

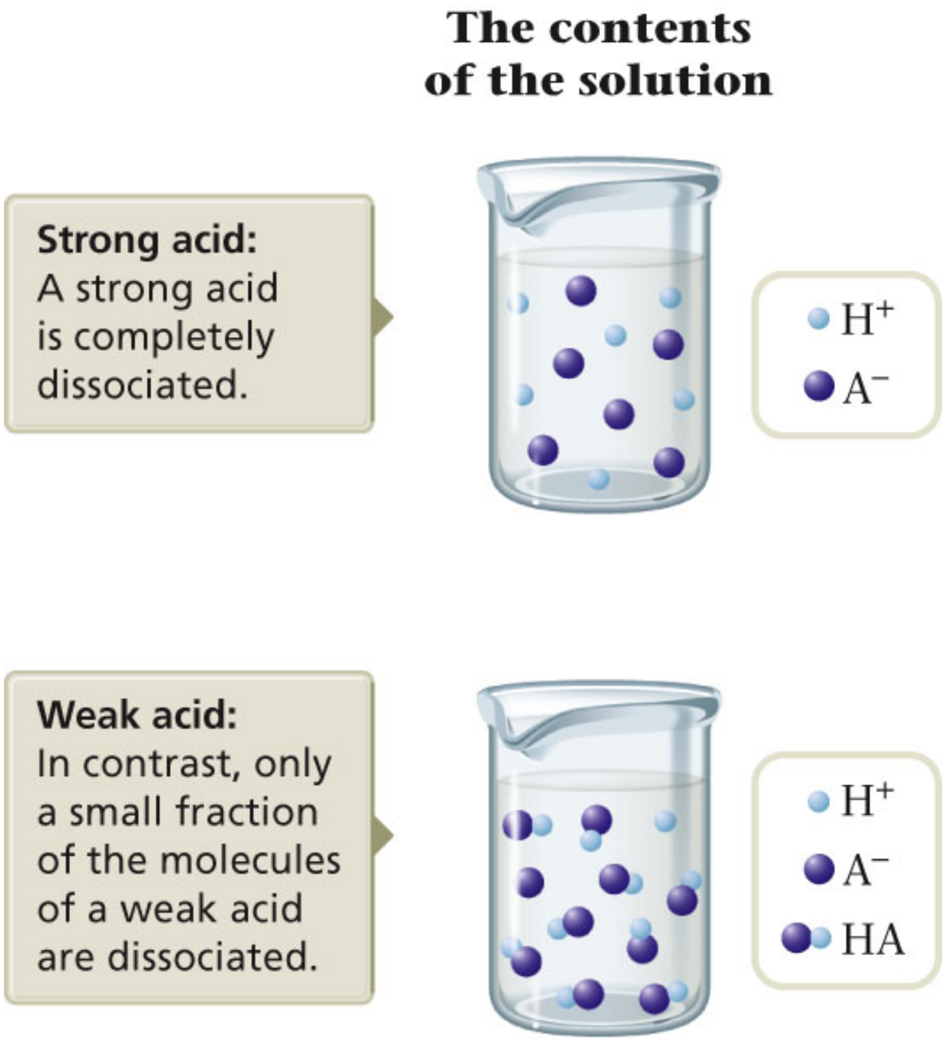
$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

- Weak acid: Most of the acid molecules remain intact.

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

equilibrium arrow

Acid Strength



Strong Acid Solutions

- Strong acids ionize completely:

$$HCl \rightarrow H^+ + Cl^-$$

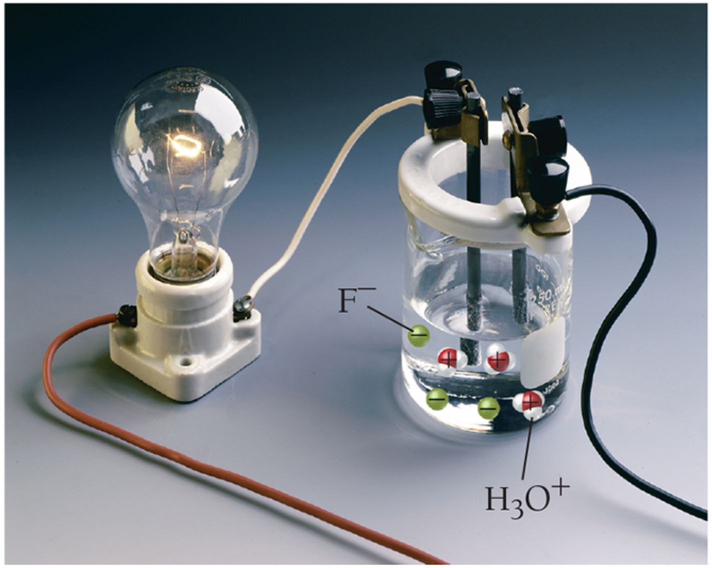
So: $\underline{0.01\ M}\ HCl \rightarrow \underline{0.01\ M}\ H^+ + \underline{0.01\ M}\ Cl^-$

$\underline{0.5\ M}\ HCl \rightarrow \underline{0.5\ M}\ H^+ + \underline{0.5\ M}\ Cl^-$

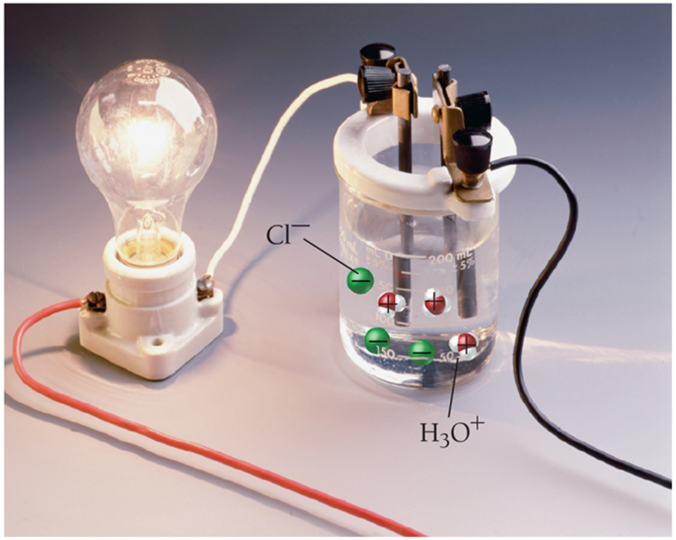
- Weak acids do not ionize completely:

Acids as Electrolytes

Electrolyte: a solution of free ions, conducts electricity



Weak acid
Weak electrolyte



Strong acid
Strong electrolyte

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Common Weak Acids $HC_2H_3O_2$ or CH_3COOH acetic acid H_2CO_3 carbonic acid	Common Weak Base NH_3 ammonia

Some Weak Acids and Weak Bases

TABLE 14.4 Weak Acids

hydrofluoric acid (HF)	sulfurous acid (H_2SO_3) *(diprotic)*
acetic acid ($HC_2H_3O_2$)	carbonic acid (H_2CO_3) *(diprotic)*
formic acid ($HCHO_2$)	phosphoric acid (H_3PO_4) *(triprotic)*

TABLE 14.6 Some Weak Bases

Base	Ionization Reaction
ammonia (NH_3)	$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
pyridine (C_5H_5N)	$C_5H_5N(aq) + H_2O(l) \rightleftharpoons C_5H_5NH^+(aq) + OH^-(aq)$
methylamine (CH_3NH_2)	$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$
ethylamine ($C_2H_5NH_2$)	$C_2H_5NH_2(aq) + H_2O(l) \rightleftharpoons C_2H_5NH_3^+(aq) + OH^-(aq)$
bicarbonate ion (HCO_3^-)*	$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

Ex Probs

Ion-Product Constant for Water

Ion-Product Constant for Water

Water as an Acid and Base

- Water is amphoteric: can react as either an acid or base

- Water undergoes auto-ionization:

  $H_2O(l) + H_2O(l) \rightarrow H_3O^+ + OH^-$

- For pure water, 25°C: $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M

- Ion-product constant for water (K_w)

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

  *True for pure water and <u>all aqueous solutions</u>!

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

We can relate $[H^+]$ and $[OH^-]$ through K_w:

$$\boxed{K_w = [H^+][OH^-] = 1.0 \times 10^{-14}}$$

for water and all aqueous solutions

Definitions: Neutral, Acidic, Basic

- Neutral solution: $[H^+] = [OH^-]$
- Acidic solution: $[H^+] > [OH^-]$
- Basic solution: $[H^+] < [OH^-]$

In each case, however,

$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$

for water and all aqueous solutions

Ex Probs

pH and pOH

The pH Scale

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

- A compact way to represent solution acidity.

eg. $pH = -\log(1.0 \times 10^{-7}\ M) = 7$

Logarithmic Function

$y = \log x \quad \rightarrow \quad 10^{y} = x$

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

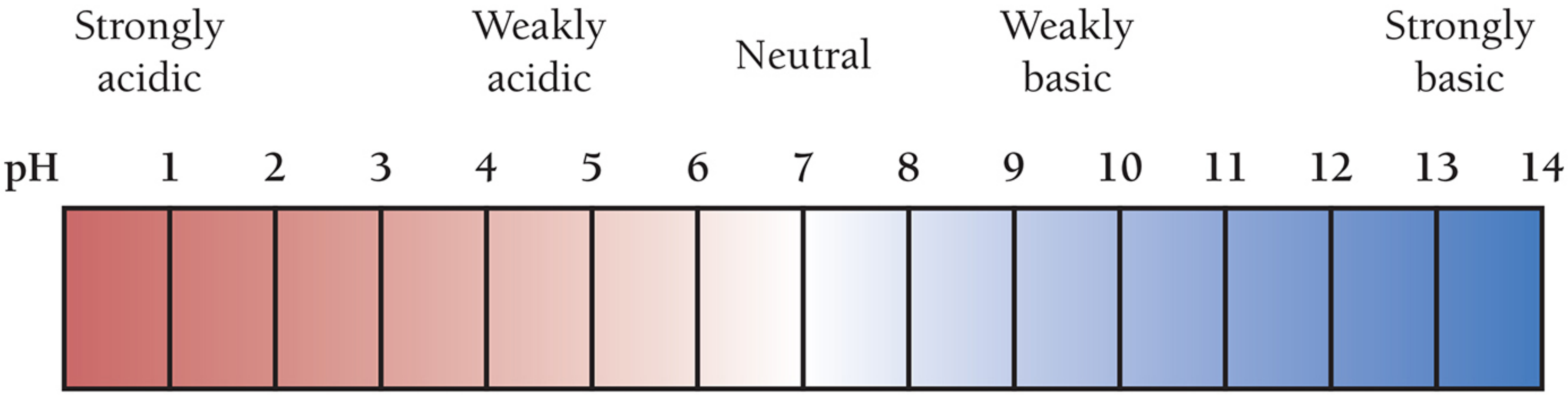
$-pH = \log [H^{+}] \quad \rightarrow \quad 10^{-pH} = [H^{+}]$

- As $[H^{+}]$ increases, pH decreases.

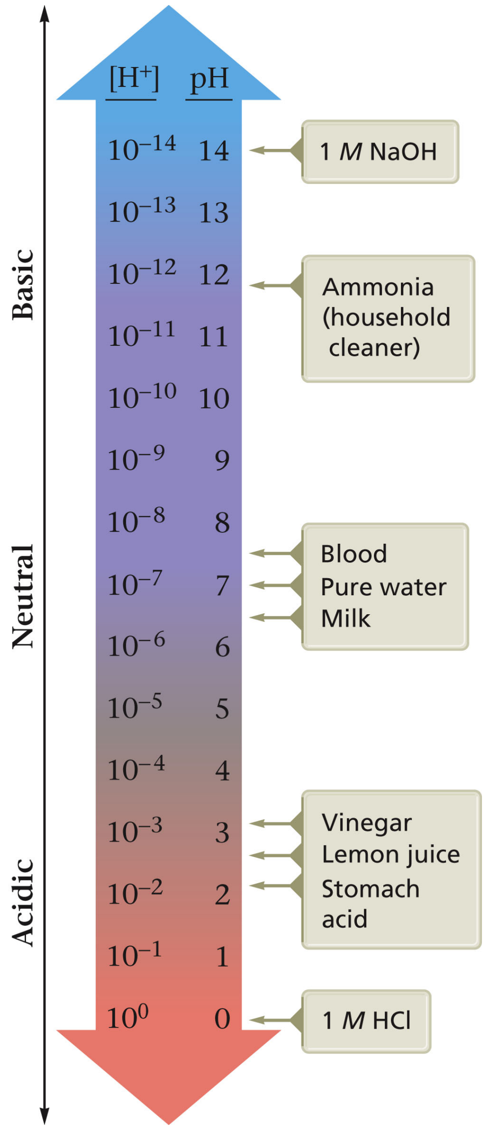
pH Range

pH 7	$[H^+] = [OH^-]$	Neutral
pH < 7	$[H^+] > [OH^-]$	Acidic
pH > 7	$[H^+] < [OH^-]$	Basic

- Lower the pH, more acidic the solution.
- Higher the pH, more basic the solution.



pH Range



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

Every time pH drops by 1, there is 10x increase in $[H^+]$.

Calculating pH and $[H^+]$

- <u>To get pH from $[H^+]$:</u>

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

- <u>To get $[H^+]$ from pH:</u>

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

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

  Inverse log $(-pH)$ = inverse log $(\log [H^+])$

  Inverse log $(-pH) = [H^+]$

  OR $10^{-pH} = [H^+]$

Ex Probs

pOH

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

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

Calculating pOH and $[OH^-]$

- <u>To get pOH from $[OH^-]$:</u>

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

- <u>To get $[OH^-]$ from pOH:</u>

  Inverse log $(-pOH) = [OH^-]$

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

Ex Probs

Relationship Between pH and pOH

$$1.0 \times 10^{-14} = [H^+][OH^-]$$

$$\log 1.0 \times 10^{-14} = \log[H^+] + \log[OH^-]$$

$$-14.00 = -\log[H^+] - \log[OH^-]$$

$$14.00 = \log[H^+] + \log[OH^-]$$

$$\boxed{14.00 = pH + pOH}$$

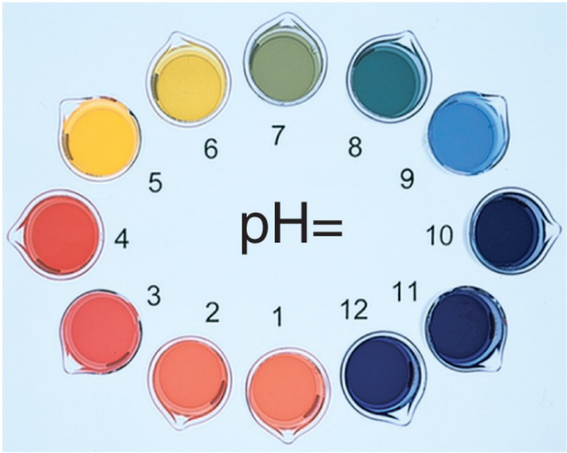
Ex Probs

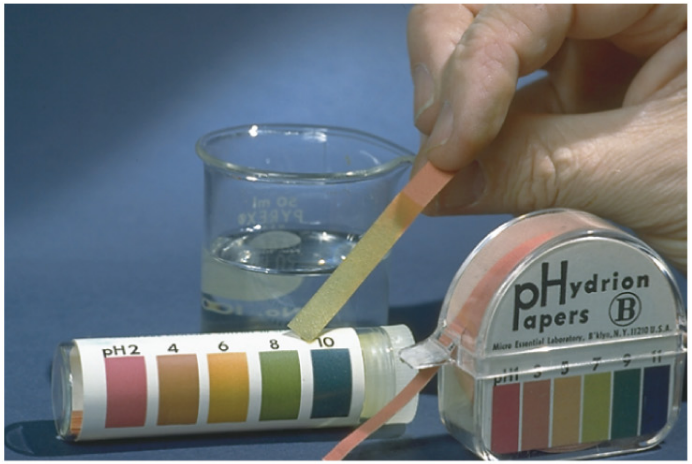
Equations for pH Calculation Problems

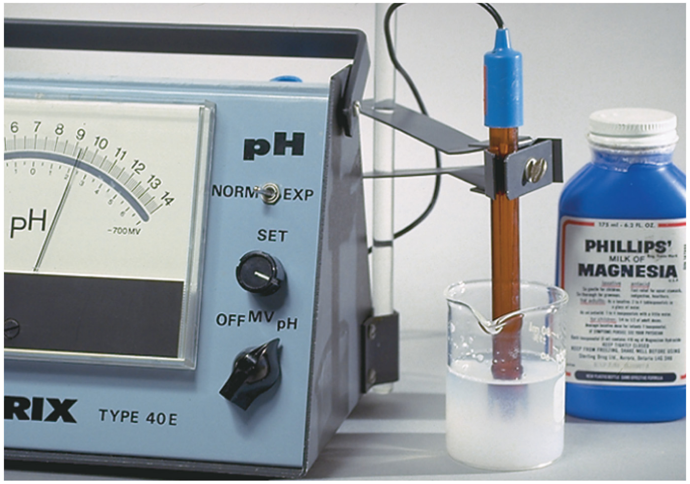
1. $pH = -\log[H^+]$ (analogous for pOH)
2. Inv log $[-pH] = [H^+]$ (analogous for $[OH^-]$)
3. $1.0 \times 10^{-14} = [H^+][OH^-]$
4. $14.00 = pH + pOH$

Ex probs

Determining pH in Laboratory







Buffers

Buffers

Buffers

- Buffer: a solution that contains both an acid and a base, thus resists pH change
- Buffers contain significant amounts of both a weak acid and its conjugate base.
- The weak acid neutralizes added base.
  The conjugate base neutralizes added acid.
- Eg. of buffer: CH_3COOH and CH_3COO^-
  (CH_3COO^- comes from $NaCH_3COO$)

Water vs. Buffer