Chem 1025 — **Introductory Chemistry, Zumdahl Decoste, 6th ed**
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Chapter 16: Acids and Bases

These Notes are to SUPPLIMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests! To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test!

The author is providing these notes as an addition to the students reading the text book and listening to the lecture. Although the author tries to keep errors to a minimum, the student is responsible for correcting any errors in these notes.

Arrhenius Concept of acids and bases, first introduced in the late 1800's

Acid Produce Hydrogen ions in aqueous solution

$HCl \rightarrow \mathbf{H^+} + Cl^-$ **Hydrochloric Acid** is a strong acid.

Sulfuric Acid H_2SO_4 was the 1st large quantity acid produced and used in the US

Base Produce hydroxide ions

$NaOH \rightarrow Na^+ + OH^-$ **Sodium Hydroxide** – a strong base

Bronsted-Lowry Model The Arrhenius Acid/Base theory works great in water solutions, but it does not work in non-aqueous solutions such as Benzene. The Bronstead-Lowry Model covers this

Acid Proton Donor **Base:** Proton Acceptor

$HCl + :NH_3 \rightarrow NH_4^+ + Cl^-$

HCl donates the proton and is the BL Acid. Ammonia accepts the proton and is the BL Base.

Conjugate Acid-Base Pair – Two substances related to each other as one donates and one accepts a single proton

HCl +	H_2O →	H_3O^+ +	Cl^-
Acid	**Base**	**Conjugate Acid**	**Conjugate Base**

Water is a **polar molecule.** The free electrons of the Oxygen pull away the H^+

[Instructor Draw pic of H-O-H] The H-O-H bond is at a 105^o

Water behaves as a base, it accepts a proton:

$H\text{-}OH + HCl \rightarrow H_3O^+ + Cl^-$

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  ..                      . .  +
H-O:  +  H+ – Cl-  →     H-O-H       + Cl-1
  |                        |
  H                        H
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Hydronium Ion [H_3O^+]

Acid Strength

$HA + H_2O \leftrightarrow H_3O^+ + A^-$ HA is an acid such as HCl, it ionizes in solution

$H_3O^+ + A^- \leftrightarrow HA + H_2O$ But the reaction is also reversable!

Completely ionized or Dissociated = strong acid HCl

Reverse Reaction = weak acid CH_3-COOH Acetic Acid

$CH_3\text{-}COOH + H_2O \leftrightarrow H_3O^+ + CH_3\text{-}COO^-$ This reaction goes about 1%, a weak acid
Acetic Acid Acetate Ion

$H_3O^+ + CH_3\text{-}COO^- \leftrightarrow CH_3\text{-}COOH + H_2O$ This is the reverse reaction and goes about 99%
Acetate Ion Acetic Acid

Strong Acid Sulfuric, Hydrochloric, Nitric, HCl

Weak Acid Acetic Acid – $HC_2H_3O_2$ = 1% dissociates, HF

Strong Base Sodium and Potassium Hydroxide

Diprotic Acid Can lose more than one H^+

Sulfuric Acid $H_2SO_4 \leftrightarrow H^+ + HSO_4^- \leftrightarrow H^+ + SO_4^-$

Oxyacid Hydrogen is attached to an Oxygen H_3PO_4 is really $O=P-(OH)_3$

H_2SO_4 is really $(O=)_2\ S\text{-}(OH)_2$

Organic acids – carboxyl groups – COOH Weak Acids

CH_3-COOH Acetic Acid - Vinegar

Water Acid / Base Amphoteric Substance – can behave as an acid or base

Water can ionize $H_2O + H_2O \leftrightarrow H_3O^+ + OH^-$

Accepts Donates a proton

K_w = Ion-Product Constant Concentration of $[H_3O^+]$ times $[OH^-]$ = Const = 1.0×10^{-14}

This is an Equilibrium Constant.

Pure water has 1.0×10^{-7} Moles / Liter of H^+ and 1.0×10^{-7} Moles / Liter of OH^-

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

Add 1 Mole of HCl to 1 Liter of water and you get 1 Mole/Liter of H^+ as HCl is a strong acid and does ionize completely in water solution.

For the following H^+ concentration, calculate the OH^- concentration: $[OH^-] = \mathbf{Kw} / [H^+]$

$[H^+] = 3.4 \times 10^{-4}$ M $[H^+] = 2.6 \times 10^{-8}$ M

$[H^+] = 6.2 \times 10^{-9}$ M $[H^+] = 8.1 \times 10^{-3}$ M

For the following OH^- concentration, calculate the H^+ concentration: $[H^+] = \mathbf{Kw} / [OH^-]$

2.9×10^{-11} M $[OH^-]$ 3.9×10^{-7} M $[OH^-]$

1.6×10^{-6} M $[OH^-]$ 1.2×10^{-12} M $[OH^-]$

pH - log [H^+] [H^+] = 1.0 x 10^{-5} M/L pH = 5.00

pH = 1-2 Strong Acid
pH = 7 Neutral = Pure Water with no dissolved CO_2
pH = 13-14 Strong Base

Calculate the pH

[H^+] = 1.0 x 10^{-9} moles/L pH = 9.00 **Discuss Significant Digits**

[OH^-] = 1.0 x 10^{-6} moles/L pH = 8.00

[H^+] = 3.6 x 10^{-9} moles/L pH = 8.44

[OH^-] = 9.2 x 10^{-2} moles/L pH = 12.96

pOH - log [OH^-] [OH^-] = 1.0 x 10^{-3} M/L pOH = 3.00

pH + pOH = 14

Rainwater has a pH of 4-5 due to the dissolved CO_2 which forms Carbonic Acid

$$\mathbf{CO_2 + H_2O \leftrightarrow H_2CO_3 \leftrightarrow H+ + HCO_3}$$

Buffered Solutions

A solution is buffered by the presence of a weak acid

$CH_3CH_2COOH \leftrightarrow H^+ + CH_3CH_2COO^-$

Show what happens with HCl and NaOH

$HCl \leftrightarrow H^+ + Cl^-$ Added to the above reaction forces the reverse reaction to occur. The added H+ reacts with the $CH_3CH_2COO^-$ and this absorbs the H^+ acid so there is no change in pH

$NaOH \leftrightarrow Na^+ + OH^-$ Added to the above reaction, the OH- reacts with the H+ to form water. Again there is no change in pH.

You'll learn a lot more about pH and Buffered Solutions in Chem 1046!