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Checklist: Acids and Bases

☐ Acids and Bases	Hydrogen Ionization
☐ Amphoteric Species (Water)	Act as BOTH Acid and Base
☐ Conjugates [H_3O^+]	Acid-Base Pair trade Proton
☐ K_w Autoionization of Water	$H_2O \rightarrow H^+ + OH^-$
☐ pH Scale	$[H^+]$Ion Concentration
☐ Hydrolysis of Acid-Base	Donor/Acceptor
☐ Weak vs. Strong Acid-Base	Degree of Dissociation
☐ Neutralization Reactions	pH 7
☐ Titration	Quantitation

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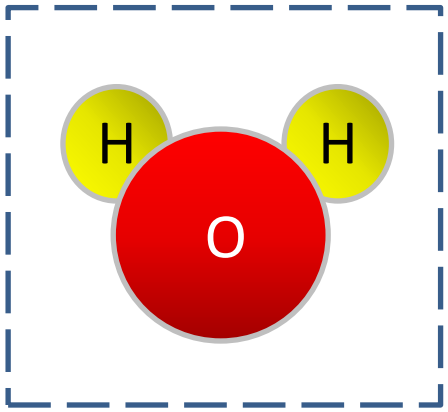
Acids and Bases

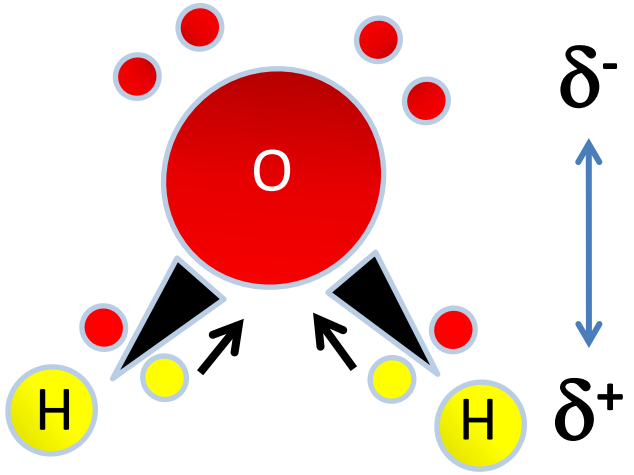


Water is the Ultimate Amphoteric Species

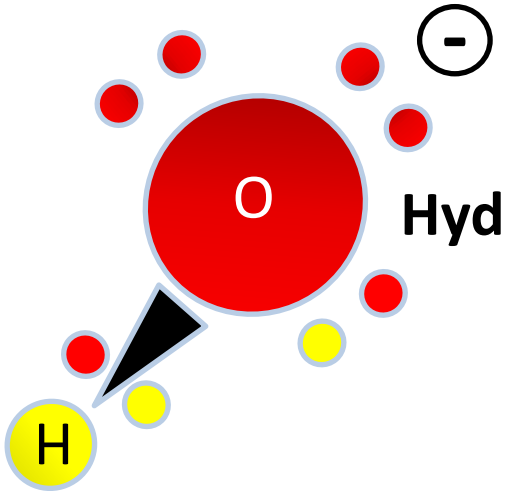
It can act as either an Acid or a Base

How?

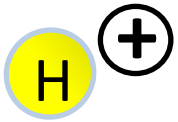
Oxygen exerts a Strong Electronegative pull on Hydrogen's Single Electron



The strength of this pull can dissociate the molecule into two Ionic Species



Hydroxyl Ion



Hydrogen Nucleus

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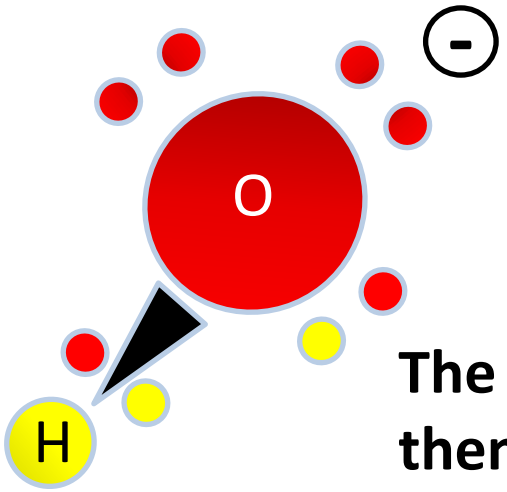
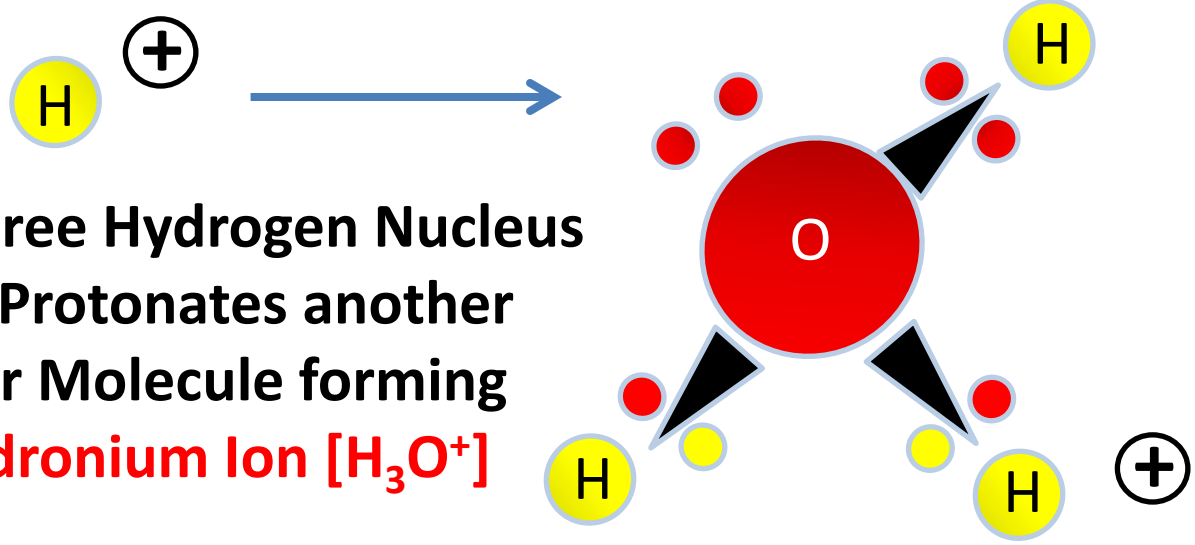
Autoionization and Conjugates



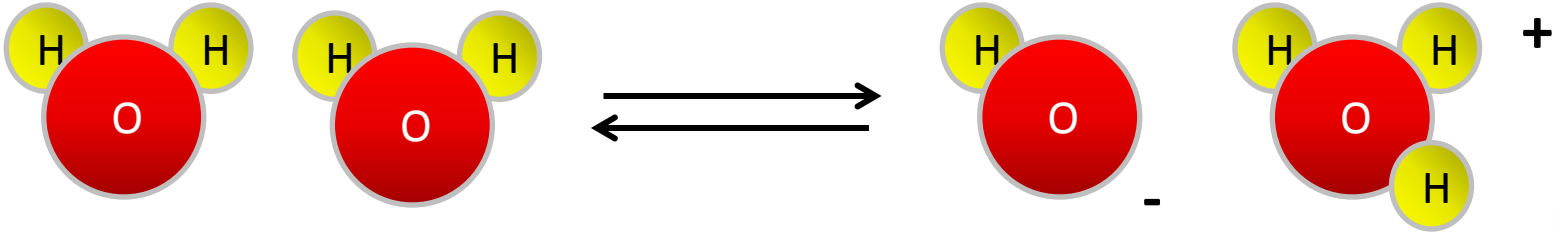
The Spontaneous Formation of Ionic Species from Water is called Autoionization (K_w)



The Free Hydrogen Nucleus then Protonates another Water Molecule forming A Hydronium Ion [H_3O^+]

The Positive Hydronium and Negative Hydroxyl Ions are Paired Species know as

Acid-Base Conjugates



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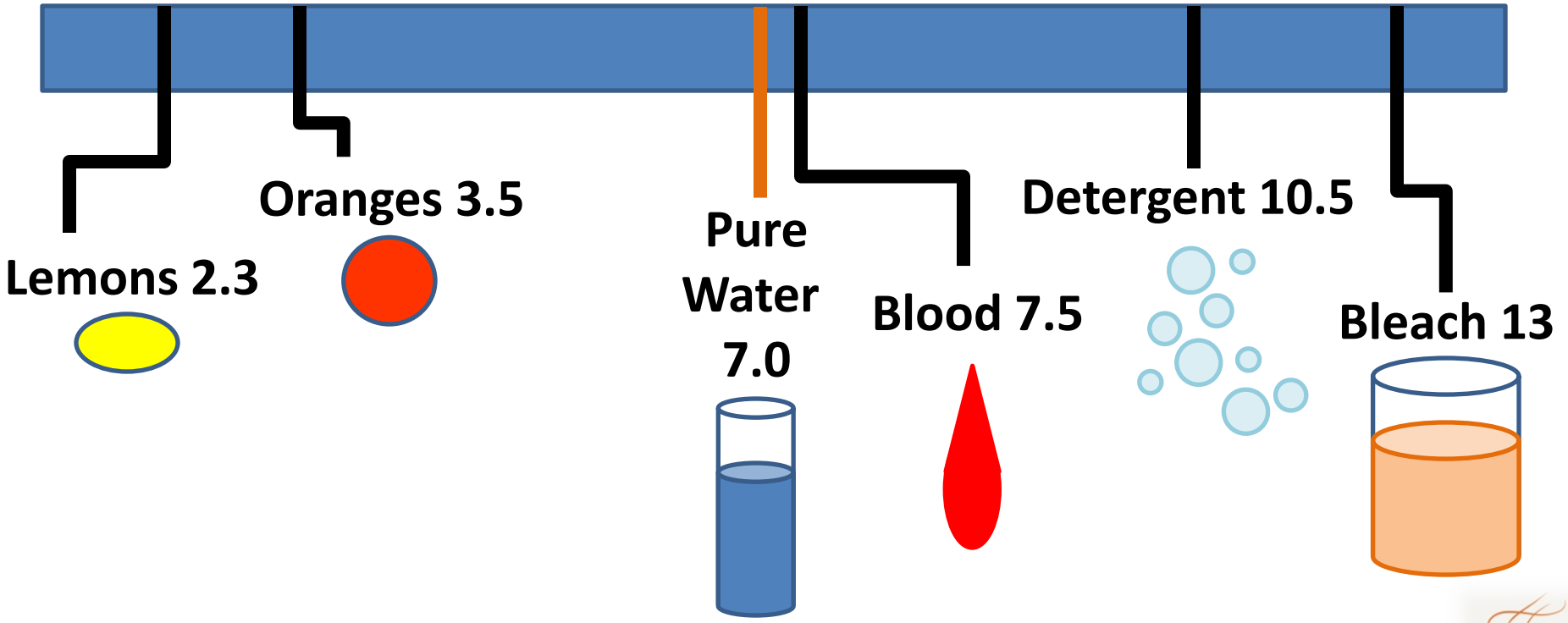
pH Scale The **Molar Concentration** of Hydronium and Hydroxyl Ions are very small: 1×10^{-7} M

For convenience a Log Scale (p) is used equaling the Negative Log of the Ion Concentration i.e. at **Neutral pH** $[H_3O^+] = [OH^+] = 1 \times 10^{-7}$ M

Therefore, $-\log [1 \times 10^{-7}] =$ **pH 7**

Acids = pH 1-7 Bases =pH 7-14



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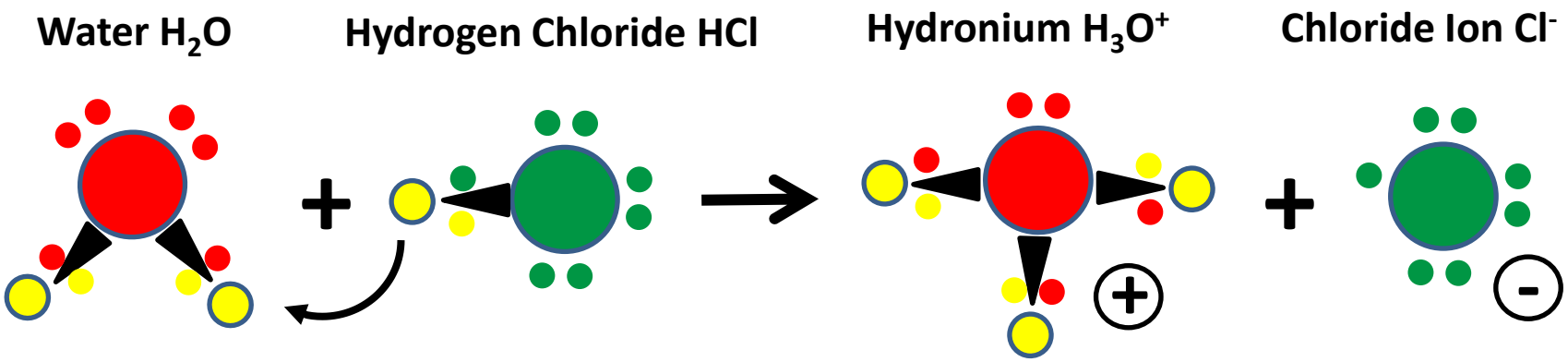
Hydrolysis

Acids and Base Conjugates may be defined by their Hydrogen Donor-Acceptor Relationship



Water Accepts the Hydrogen as a Base

HCl Donates the Hydrogen as a Acid

Acid-Base Definitions:

Arrhenius Acid/Base	Ion Donor/ Acceptor
Bronsted-Lowery Acid/Base	Hydrogen Donor/Acceptor
Lewis Acid/Base	Electron Pair Donor/Acceptor

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Weak vs. Strong Acid and Base

Strong Acids and Bases become **Completely Dissociated** Into their Ionic Components in Aqueous Solution.

Examples:

Hydrochloric Acid

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

Sodium Hydroxide

$H_2O + NaOH \longrightarrow Na^+ + H_3O^+ + OH^-$

Weak Acids and Bases only **Partially Dissociate** into their Ionic Components in Aqueous Solution.

Examples:

Acetic Acid

$H_2O + CH_3COOH \longrightarrow H_2O + CH_3COOH + H_3O^+ + CH_3COO^-$

Ammonia

$H_2O + NH_4 \longrightarrow H_2O + NH_4 + NH_3 + H_3O^+ + OH^-$

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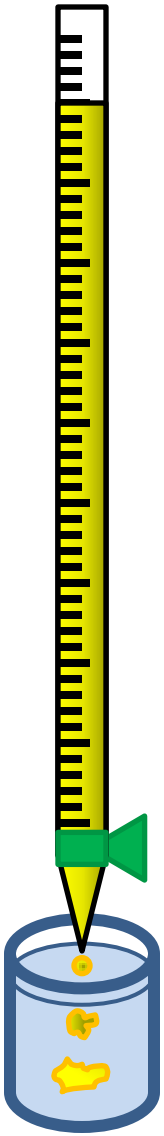
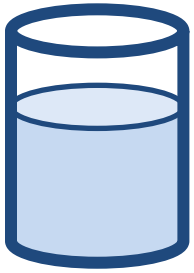
Titration and Neutralization Reactions

Burrette

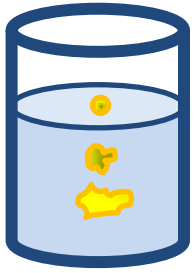


Titration is an Experimental Method for Determining the Neutralization Point of Acids and Bases

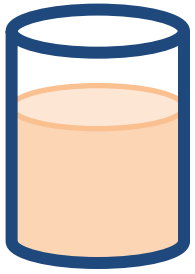
Using a Graduated Burrette, a Known concentration of a Strong Acid may be Added, Dropwise, to To an Unknown Base containing a pH Indicator Dye that is sensitive to color change at the Neutralization Point i.e. pH 7



Unknown Solution



Measured Addition



Neutralization Point

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pH Indicators

Indicator	Colour in acid (pH < 7)	Colour at pH = 7	Colour in base (pH > 7)
Red cabbage water	red, pink	purple	blue, green, yellow
Red onion water	red	violet	green
Turmeric water	yellow	yellow	red
Phenolphthalein	colourless	colourless	pink, red
Bromothymol blue	yellow	green	blue
Red litmus	red	red	blue
Blue litmus	red	blue	blue
Universal indicator	red, orange, yellow	green	Blue, violet, purple

http://www.mstworkbooks.co.za/natural-sciences/gr9/gr9-mm-05.html