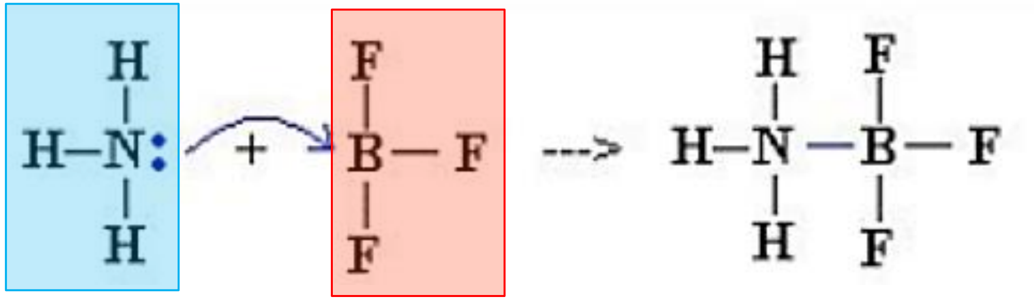
N36 – Acid Base

Target: I can review Acid Base definitions and perform pH calculations for strong acids/bases

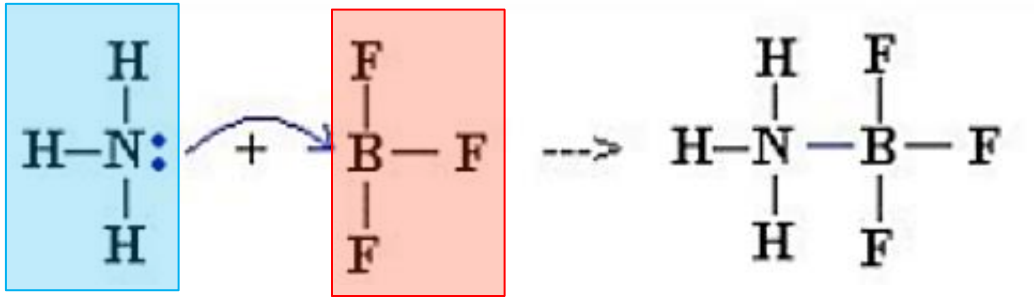
N36 – Acid Base

Quick Review

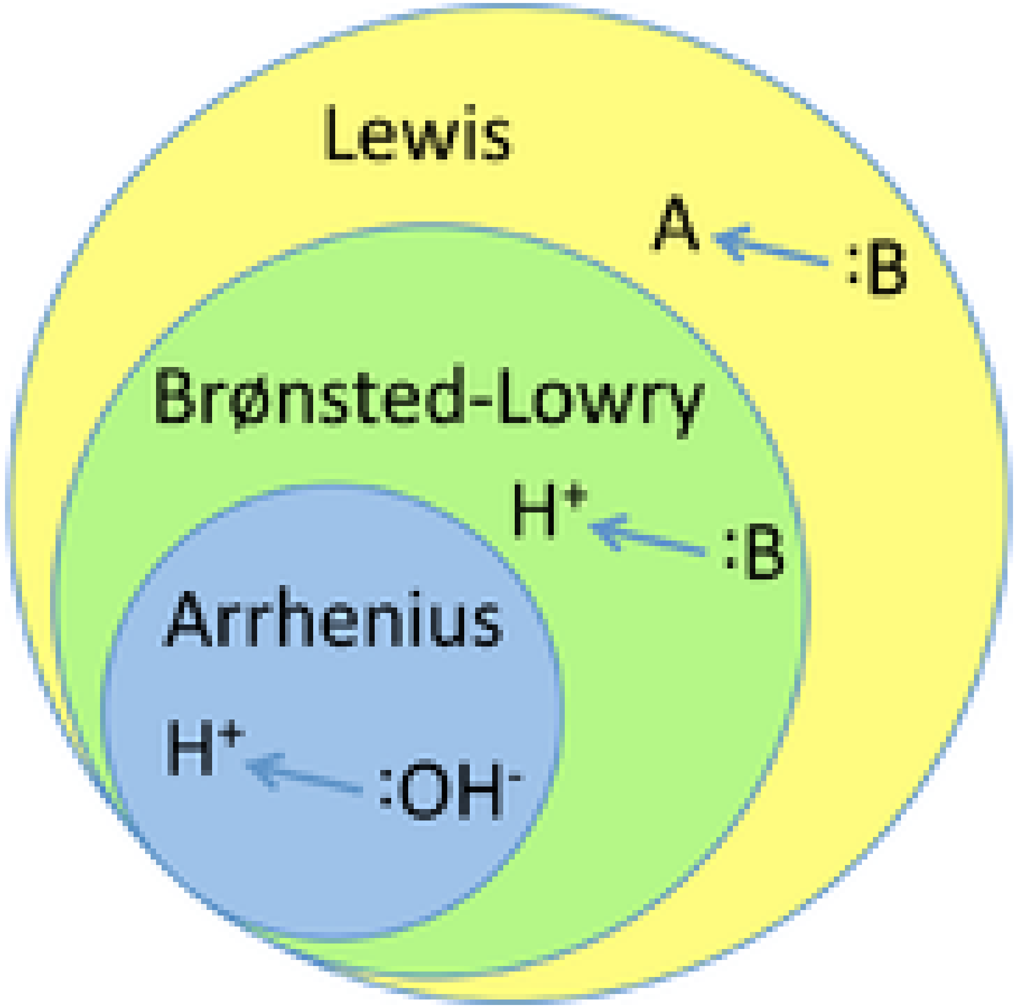
Acid/Base Definitions

Arrhenius Model	$HNO_3 \rightarrow H^+ + NO_3^-$ $KOH \rightarrow K^+ + OH^-$
Bronsted-Lowry Model	$HCl + NH_3 \leftrightarrow Cl^- + NH_4^+$
Lewis Acid Model	

Acid/Base Definitions

Arrhenius Model •Acids produce H^+ •Bases produce OH^-	$HNO_3 \rightarrow H^+ + NO_3^-$ $KOH \rightarrow K^+ + OH^-$
Bronsted-Lowry Model •Acids are proton donors •Bases are proton acceptors	$HCl + NH_3 \leftrightarrow Cl^- + NH_4^+$
Lewis Acid Model •Acids are electron pair acceptors •Bases are electron pair donors	

Acid/Base Definitions



- **Arrhenius is MOST specific**
- **Brønsted-Lowry is less specific**
- **Lewis is LEAST specific**

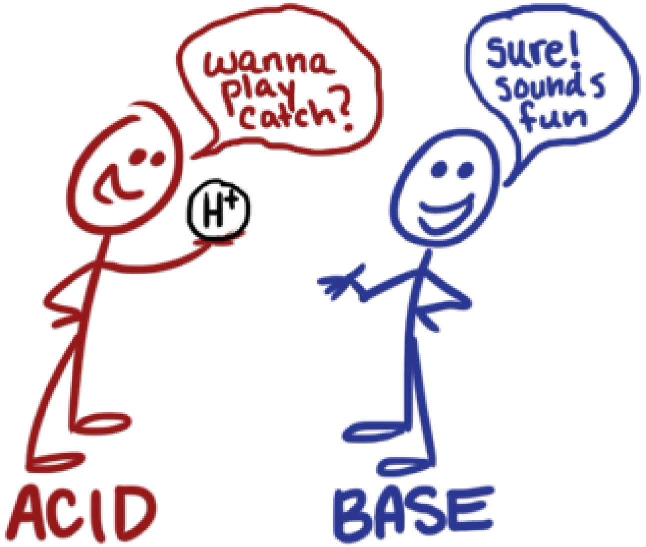
Problems with Arrhenius Theory

Does not explain why:

- Some molecular substances, (NH_3) dissolve in water to form basic solutions, even though they do not contain $OH^–$ ions.
- How some ionic comp, (Na_2CO_3 or Na_2O) dissolve in water to form basic sol’ns, even though they don’t contain OH^-
- Why some molecular substances, (CO_2) dissolve in water to form acidic solutions, even though they do not contain H^+ ions.
- Acid–base reactions that take place outside aqueous solution.

Brønsted–Lowry Acid–Base Theory

- It defines acids and bases based on what happens in a rxn.
- Any reaction involving H^+ (proton) that transfers from one molecule to another is an acid–base reaction, regardless of whether it occurs in aqueous solution or if there is OH^- present.
- All reactions that fit the Arrhenius definition also fit the Brønsted–Lowry definition.



Brønsted–Lowry Theory

In a Brønsted–Lowry acid–base reaction, the acid molecule donates an H^+ to the base molecule.

$$\mathbf{H{-}A} + \mathbf{:B} \leftrightarrow \mathbf{:A^{-}} + \mathbf{H{-}B^{+}}$$

$$\mathbf{HCl} + \mathbf{NH_3} \leftrightarrow \mathbf{Cl^{-}} + \mathbf{NH_4^{+}}$$

- The **acid** is an H^+ donor.
- The **base** is an H^+ acceptor.
  - Base structure must contain an atom with an unshared pair of electrons.

Brønsted–Lowry Acids

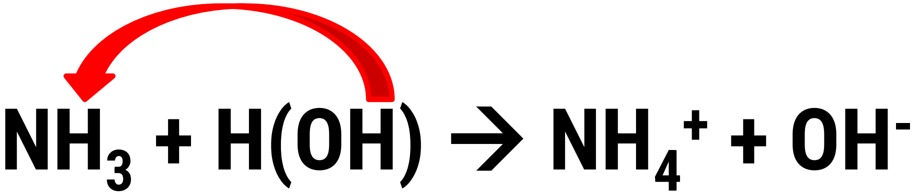
- H^+ donors.
  - Any material that has H can potentially be a Brønsted–Lowry acid.
  - Because of the molecular structure, often one H in the molecule is easier to transfer than others.

- When HCl dissolves in water, the HCl is the acid because HCl transfers an H^+ to H_2O, forming H_3O^+ ions.
  - Water acts as base, accepting H^+.

$$\mathbf{\underset{\color{red}{Acid}}{HCl(\mathit{aq})} + \underset{\color{blue}{Base}}{H_2O(\mathit{l})} \rightarrow Cl^-(\mathit{aq}) + H_3O^+(\mathit{aq})}$$

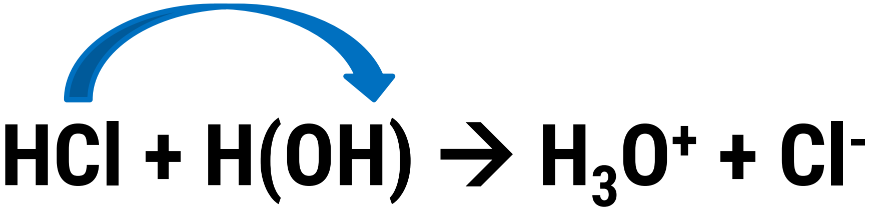
Amphoteric Substances

Amphoteric substances can act as either an acid or a base because they have both a transferable H and an atom with lone pair electrons.



$NH_3 + H(OH) \rightarrow NH_4^+ + OH^-$

*Water is donating a proton…**ACID!***



$HCl + H(OH) \rightarrow H_3O^+ + Cl^-$

*Water is accepting a proton…**BASE!***

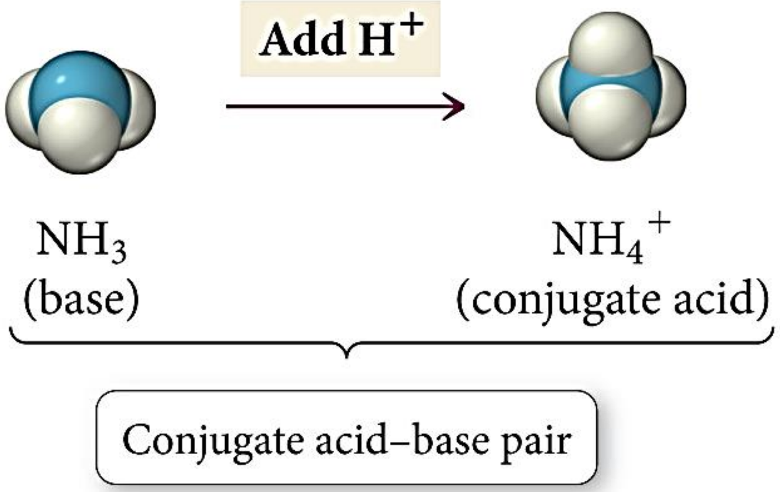
Conjugate Acid–Base Pairs

Acids turn into **“Conjugate Bases”** once they have lost their proton/hydrogen

Bases turn into **“Conjugate Acids”** once they have gained a proton/hydrogen

Conjugate Acid–Base Pairs

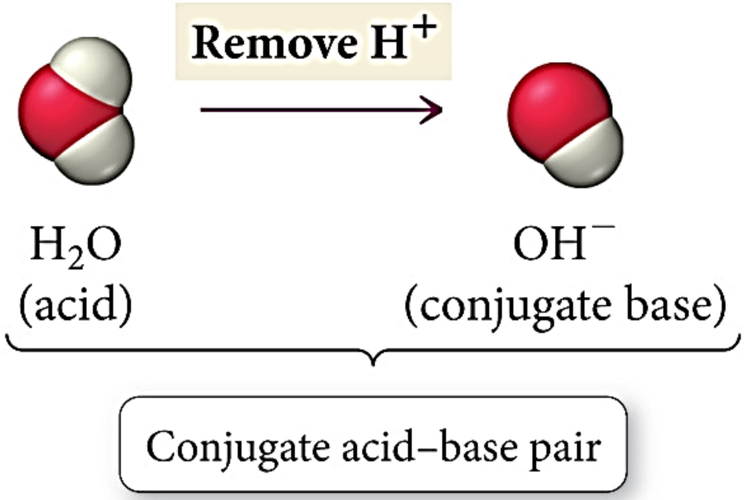
- Each reactant and the product it becomes is called a **conjugate pair**.



Add H^+

NH_3 (base) → NH_4^+ (conjugate acid)

Conjugate acid–base pair

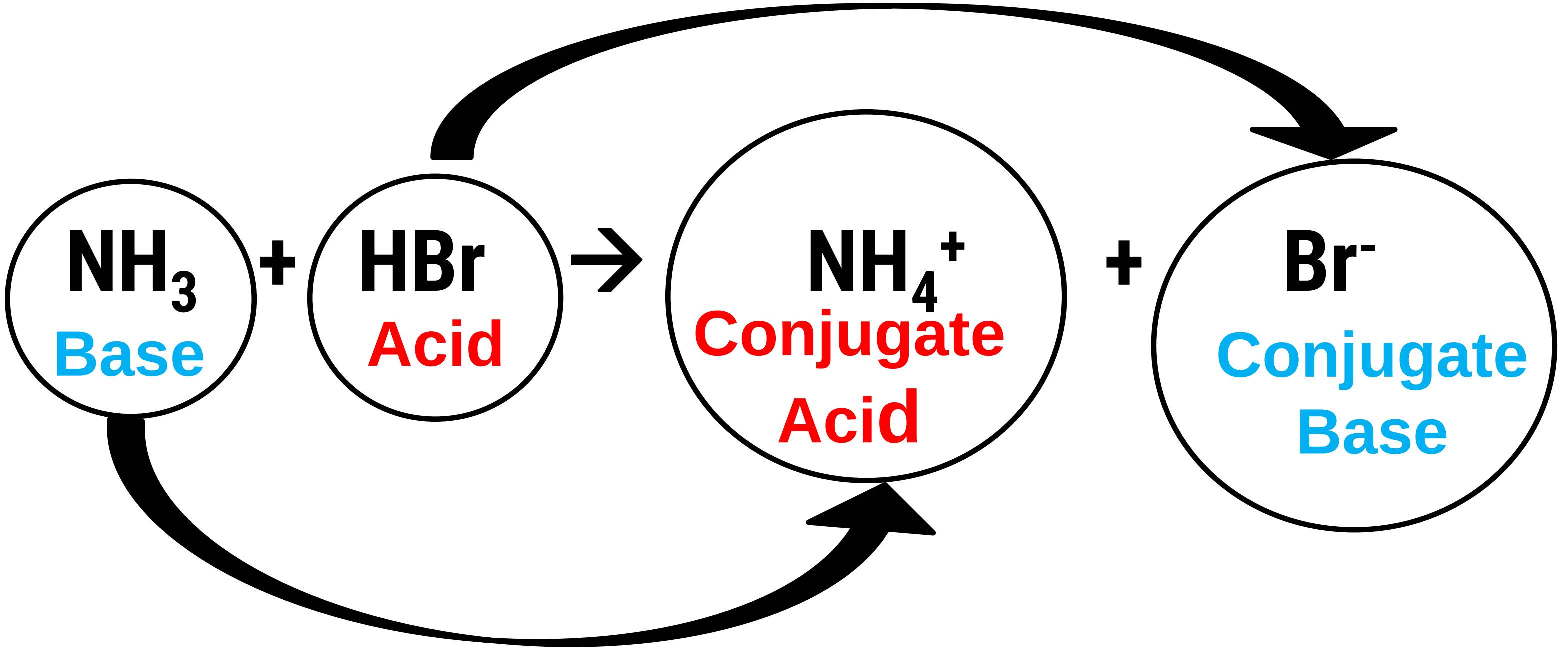


Remove H^+

H_2O (acid) → OH^- (conjugate base)

Conjugate acid–base pair

Conjugate Pairs



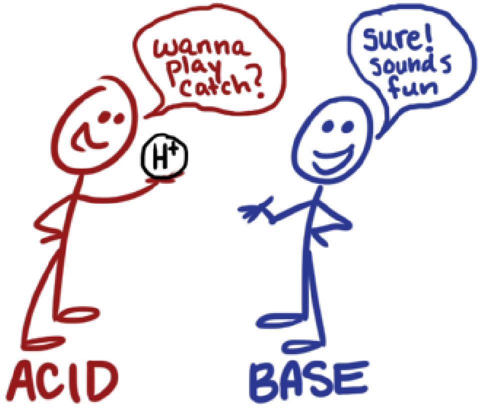
$NH_3 + HBr \rightarrow NH_4^+ + Br^-$

NH_3 Base

HBr Acid

NH_4^+ Conjugate Acid

Br^- Conjugate Base



Tips for Finding Each

- Find the Acid First – usually easiest!
- Find It's Conjugate Base – the part left after donating its H+!
- Repeat with Base and Conjugate Acid

Acid Dissociation

Back to equilibrium!!!! Woohoo!

$$HA \rightarrow H^{+} + A^{-}$$

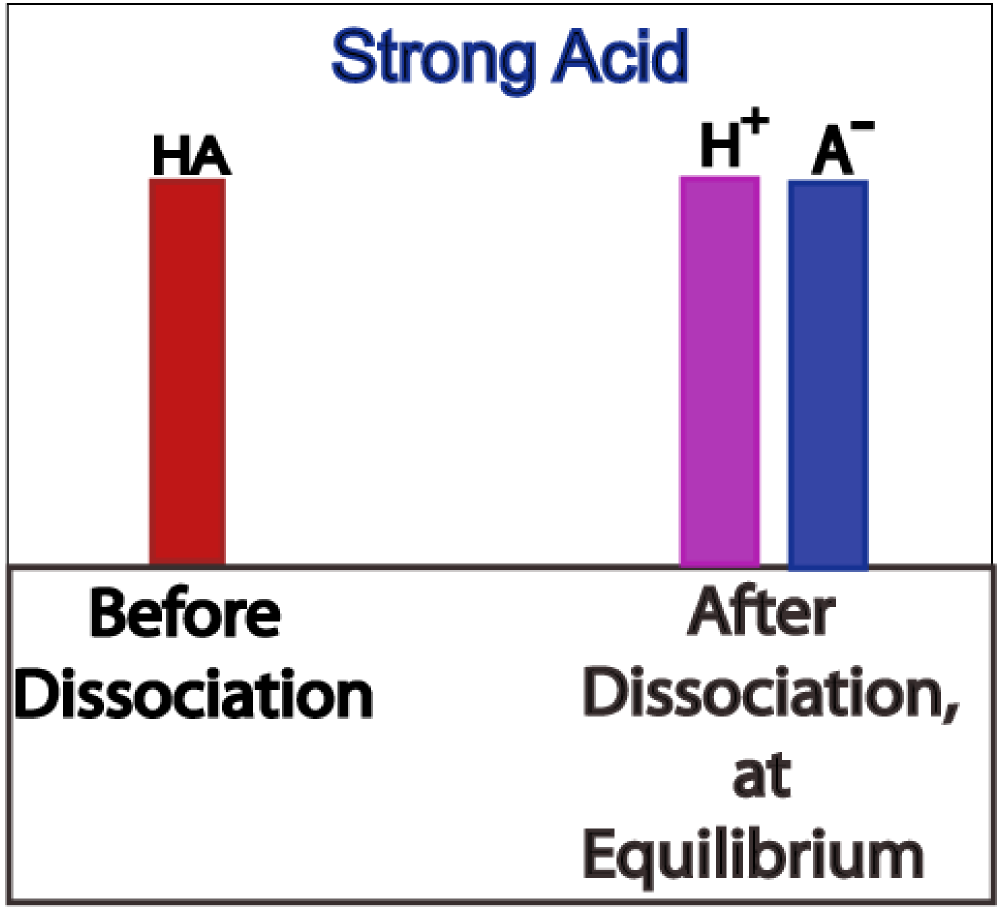
Acid **Proton** **Conjugate base**

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Can also H^+ be written in its hydrated form, H_3O^+ (hydronium ion)

Dissociation of Strong Acids

Strong acids are assumed to dissociate completely in solution.



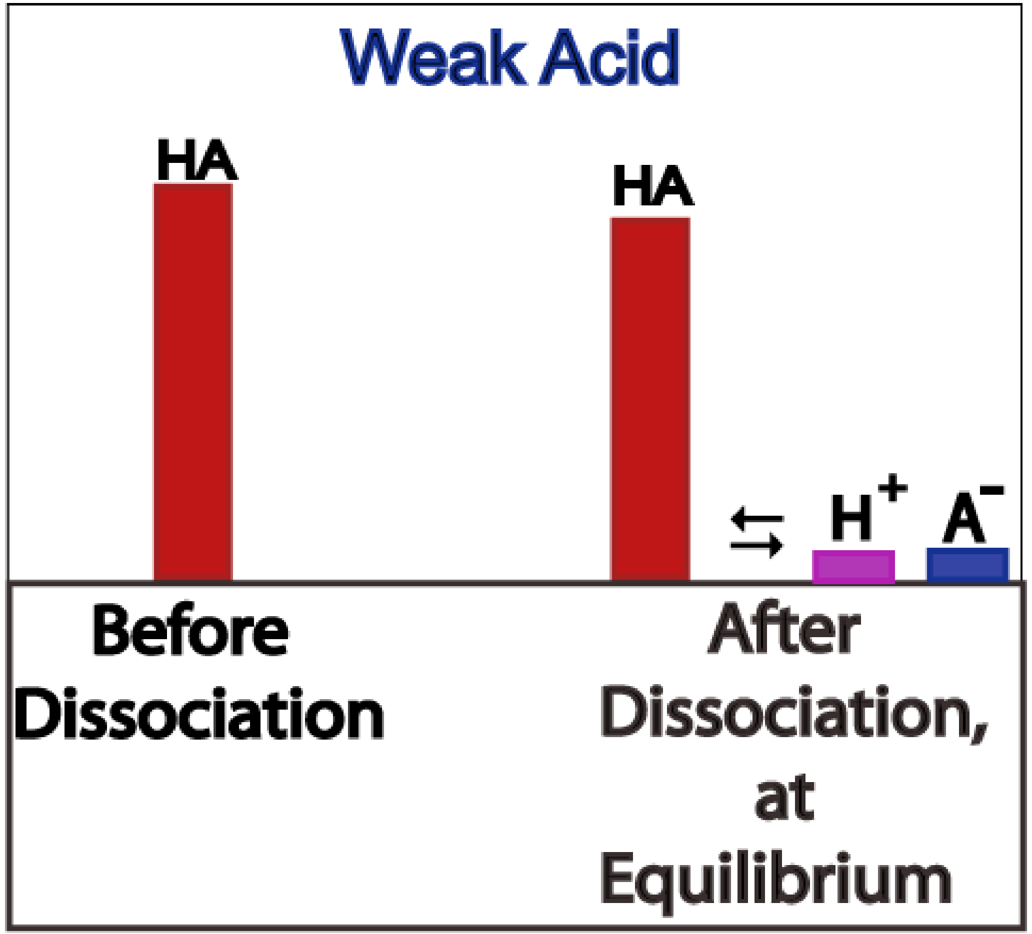
Reactant favored or product favored?	**Product Favored**
Large K_a or small K_a?	**Large K_a**

Dissociation Constants: Strong Acids

Acid	Formula	Conjugate Base	K_a
Perchloric	$HClO_4$	ClO_4^-	Very large
Hydriodic	HI	I^-	Very large
Hydrobromic	HBr	Br^-	Very large
Hydrochloric	HCl	Cl^-	Very large
Nitric	HNO_3	NO_3^-	Very large
Sulfuric	H_2SO_4	HSO_4^-	Very large
Hydronium ion	H_3O^+	H_2O	1.0

Dissociation of Weak Acids

Weak acids are assumed to dissociate only slightly (less than 5%) in solution.



Reactant favored or product favored?	**Reactant Favored**
Large K_a or small K_a?	**Small K_a**

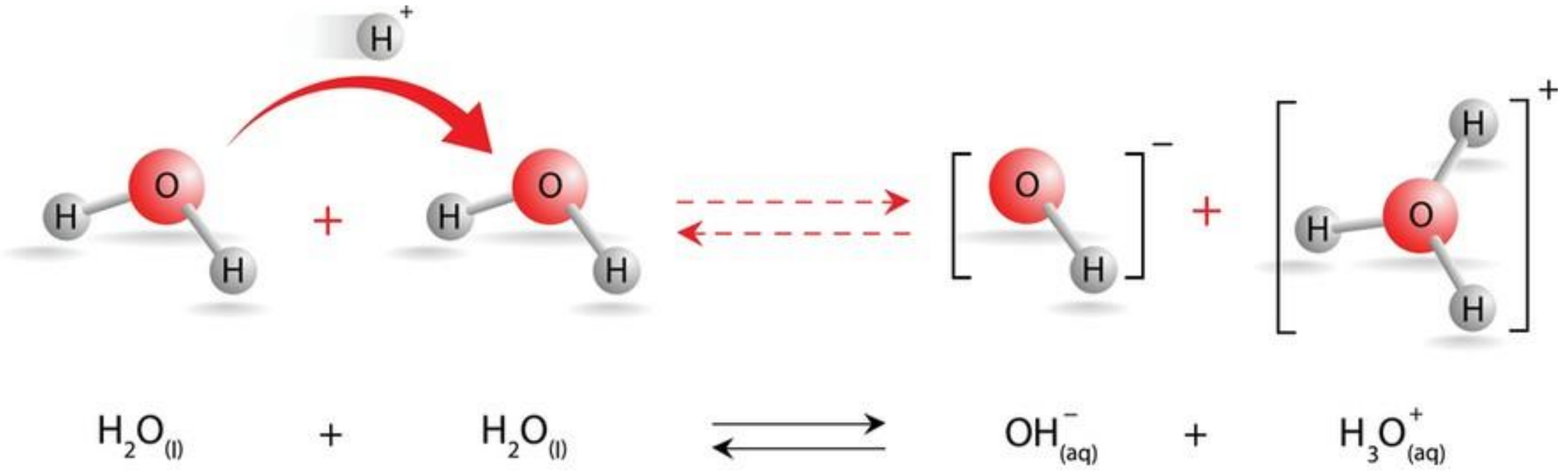
Dissociation Constants: Weak Acids

Acid	Formula	Conjugate Base	K_a
Iodic	HIO_3	IO_3^-	1.7×10^{-1}
Oxalic	$H_2C_2O_4$	$HC_2O_4^-$	5.9×10^{-2}
Sulfurous	H_2SO_3	HSO_3^-	1.5×10^{-2}
Phosphoric	H_3PO_4	$H_2PO_4^-$	7.5×10^{-3}
Citric	$H_3C_6H_5O_7$	$H_2C_6H_5O_7^-$	7.1×10^{-4}
Nitrous	HNO_2	NO_2^-	4.6×10^{-4}
Hydrofluoric	HF	F^-	3.5×10^{-4}
Formic	$HCOOH$	$HCOO^-$	1.8×10^{-4}
Benzoic	C_6H_5COOH	$C_6H_5COO^-$	6.5×10^{-5}
Acetic	CH_3COOH	CH_3COO^-	1.8×10^{-5}
Carbonic	H_2CO_3	HCO_3^-	4.3×10^{-7}
Hypochlorous	$HClO$	ClO^-	3.0×10^{-8}
Hydrocyanic	HCN	CN^-	4.9×10^{-10}

Random fact

Weak acids will dissociate more when they are dilute! The lower the concentration the higher % dissociation they will have.

Self-Ionization of Water



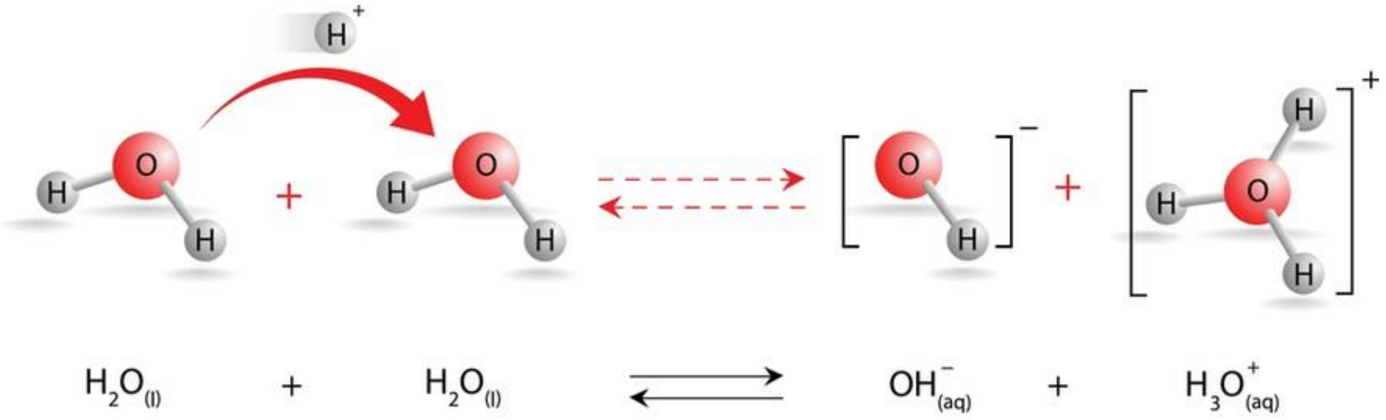
$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + H_3O^+_{(aq)}$$

***K_w* is a constant at 25 °C:**

$K_w = [H_3O^+][OH^-]$ **At 25°, $[H_3O^+] = [OH^-] = 1 \times 10^{-7}$**

$K_w = (1 \times 10^{-7})(1 \times 10^{-7}) = 1 \times 10^{-14}$

Self-Ionization of Water



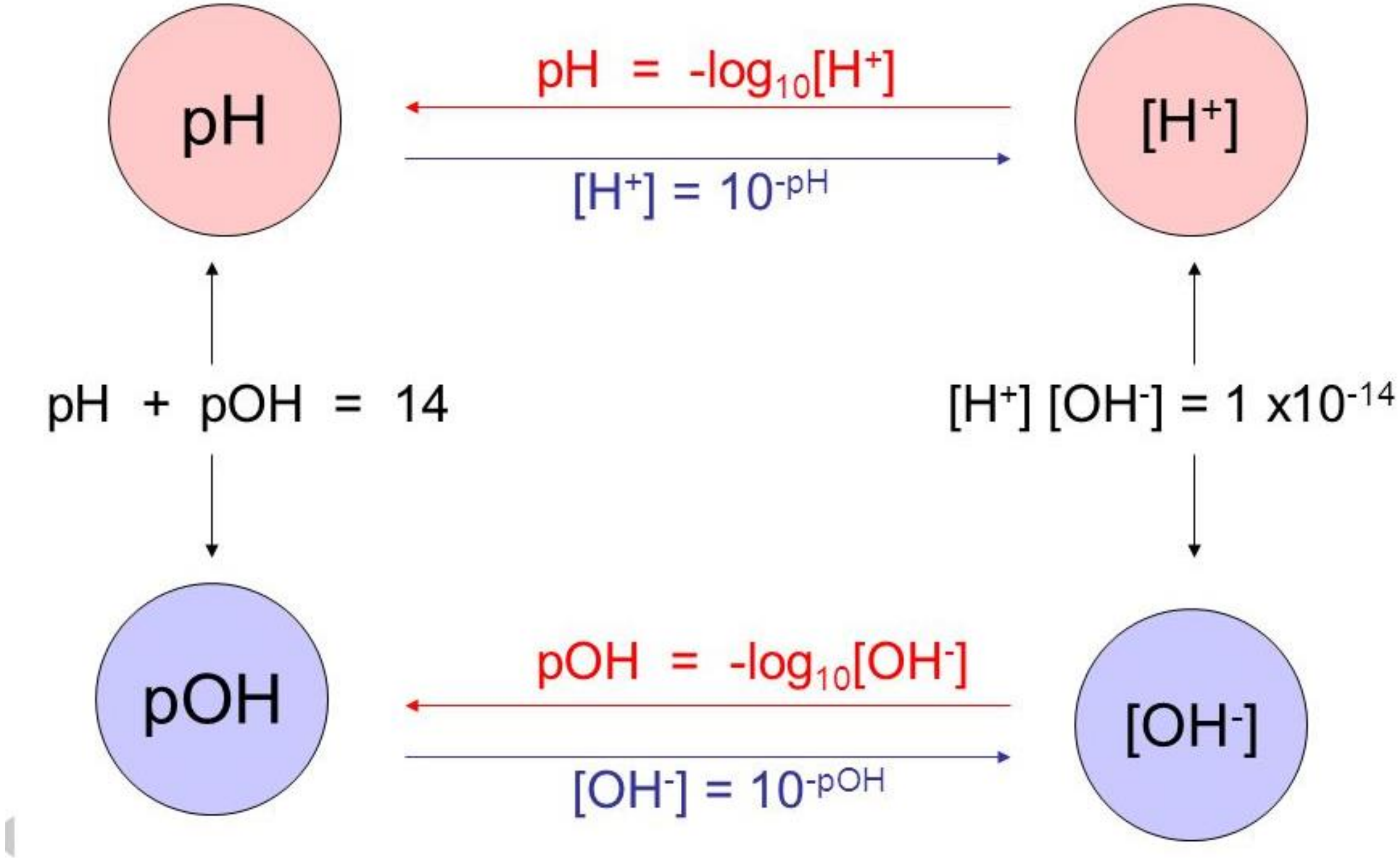
$$H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + H_3O^+_{(aq)}$$

- K_w is only 1 x 10^{-14} when at 25 °C
- Neutral is only pH 7 when at 25 °C
- K_w increases as temp increases – more dissociates!
- pH of neutral gets lower as temp increases

pH and pOH Calculations

$pH = -\log [H^+]$	$pOH = -\log [OH^-]$
$[H^+] = 10^{-pH}$	$[OH^-] = 10^{-pOH}$
$pH + pOH = 14$	
$[H^+][OH^-] = 1 \times 10^{-14}$	
With these calculations you can plug in, rearrange, substitute and find everything no matter what you are given in the problem!	

pH and pOH Calculations



pH

$[H^+]$

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

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

$pH + pOH = 14$

$[H^+][OH^-] = 1 \times 10^{-14}$

pOH

$[OH^-]$

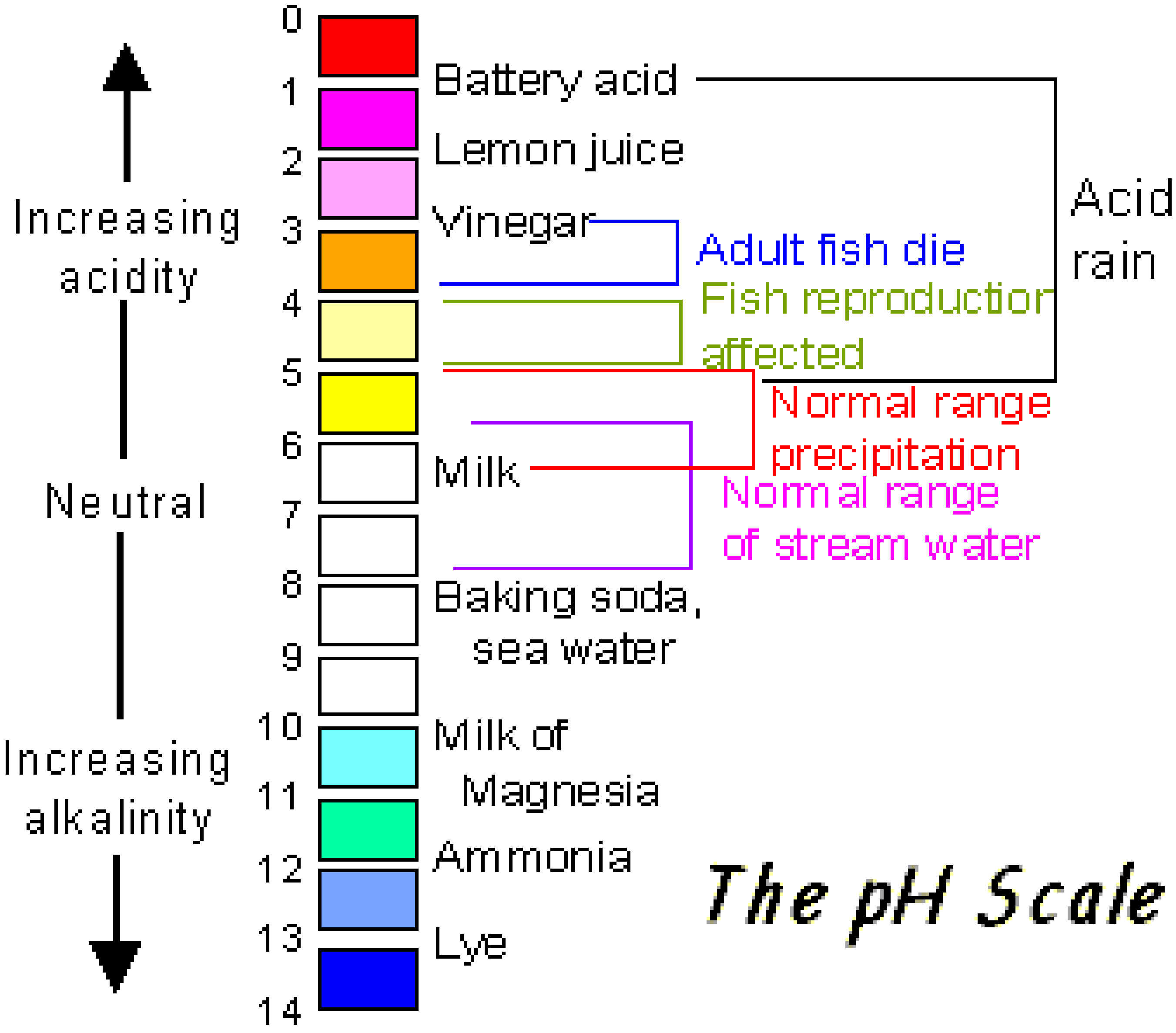
$pOH = -\log_{10}[OH^-]$

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

pH and pOH Calculations

$$\begin{tikzcd}[column sep=huge, row sep=huge]
H^{+} \arrow[r, shift left=1ex, "{[OH^{-}] = \frac{1 \text{ x } 10^{-14}}{[H^{+}]}}"] \arrow[d, shift left=1ex, "{pH = -\log[H^{+}]}"] & OH^{-} \arrow[l, shift left=1ex, "{[H^{+}] = \frac{1 \text{ x } 10^{-14}}{[OH^{-}]}}"] \arrow[d, shift left=1ex, "{pOH = -\log[OH^{-}]}"] \\
pH \arrow[u, shift left=1ex, "{[H^{+}] = 10^{-pH}}"] \arrow[r, shift left=1ex, "{pOH = 14 - pH}"] & pOH \arrow[l, shift left=1ex, "{pH = 14 - pOH}"] \arrow[u, shift left=1ex, "{[OH^{-}] = 10^{-pOH}}"]
\end{tikzcd}$$

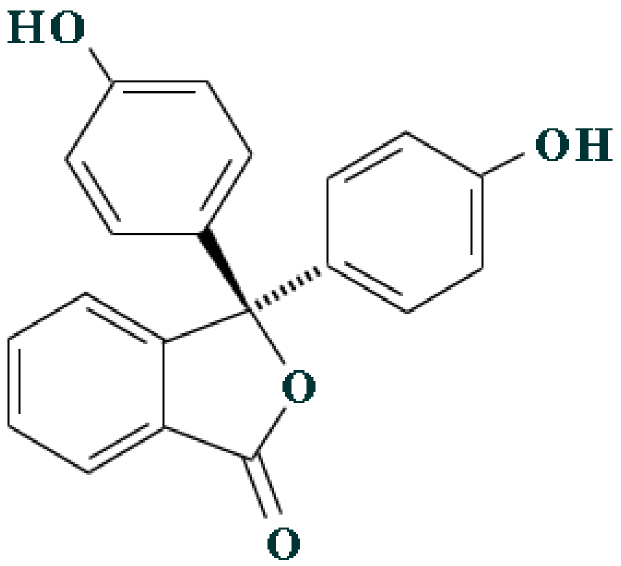
pH Scale

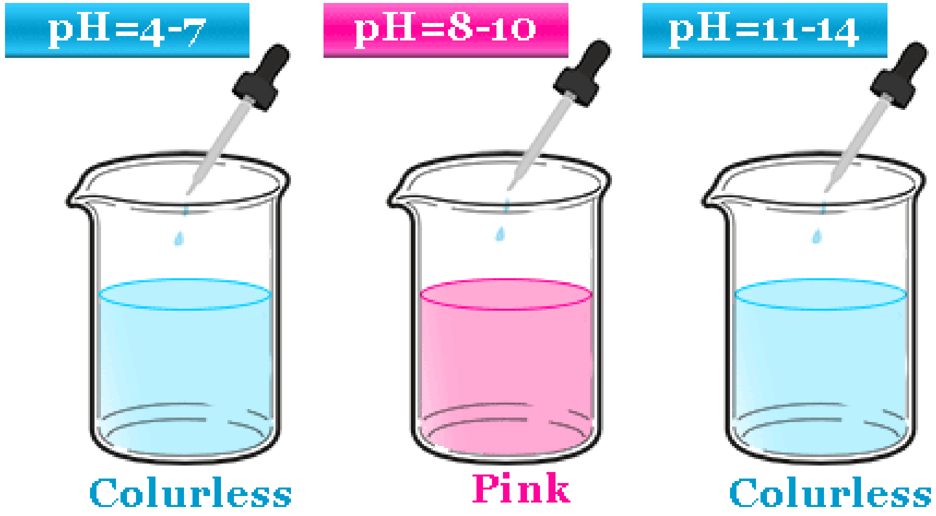


Courtesy of Environment Canada (http://www.ns.ec.gc.ca/)

pH Indicators

Some substances turn colors when the pH changes into a certain range, because their structure is changing. This is a very handy tool for us!





Example:

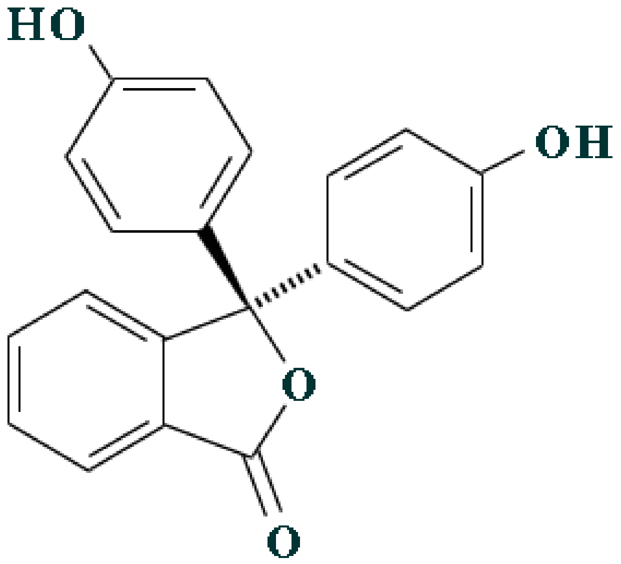
Phenolphthalein turns from clear to pink when the pH reaches
~ 8 - 10.

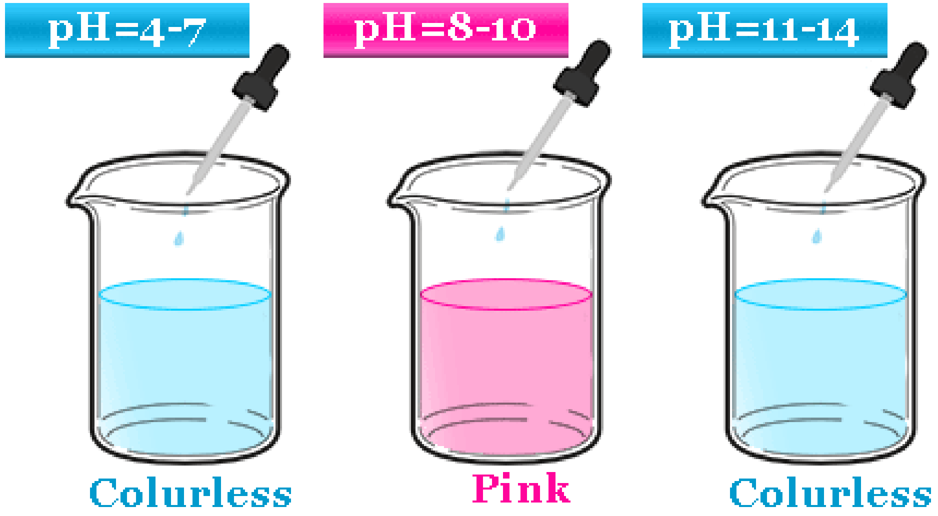
pH Indicators

You can predict the effective color change region of an indicator by looking at it’s pKa.

Color Change Region = +/- 1 pH unit from the indicator’s pKa.

The pKa of phenolphthalein is 9.3, therefore it will change colors between a pH of ~ 8 - 10





YouTube Link to Presentation

https://youtu.be/uS0P-Qp9tB8