ACIDS

Properties of Acids: sour, corrosive, reacts with metals to form hydrogen gas
react with bases to form salt (ionic compound) and water
pH < 7 Litmus red on pH paper strip

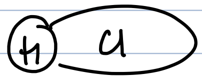
Strong Acids: dissociates 100% of the time
HCl, HBr, HI, HNO_3, $HClO_4$ and H_2SO_4

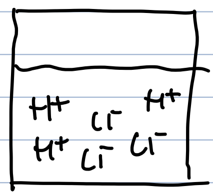
Weak Acids: dissociates less than 100% of the time
All other acids are weak acids except for 6

When Acids have a strong IMF between the hydrogen ion on the other metal, it will not dissociate 100% because it is more attracted to each other than the water.

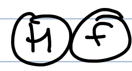
☆ strong IMF = weak acid ☆ The weak IMF makes it more dissociative

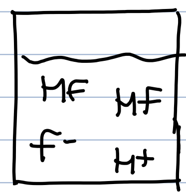
Ex. HCl (strong) =





vs. HF (weak)





→ some dissociate, but not all

↳ the nucleuses are so far apart that it makes it easy to pull apart and disosiate

↳ nucleus are so close close to each other which makes it hard to pull apart

the number of hydrogens do not determine the strength of the acid, the dissociation does

Because acids dissociate, they are considered electrolytes

Acids are sometimes covalent but act like ionic compounds

ACID / BASES NAMING

Binary (H + Halogen or very electronegative element):
- no oxygens Hydro ___ ic acid replace -ide with -ic

Oxyacid (H + oxygen containing Polyatomic ion)
- no hydro-prefix turn -ate into -ic acid and -ite to -ous acid

Exceptions: if sulfur is present, add -ur before ending
if phosphorous is present, add -or before ending

Bases: same as ionic rules exception: NH_3 = ammonia

BASES

Properties: bitter, caustic (burns), slippery, litmus blue, electrolyte pH > 7

Strong bases: Hydroxide and Group 1 or Group 2 metals are all strong and dissociate 100%.

Group 1 and Group 2 are the least +ve when hydroxide is −ve 1 charge. Have weak IMF which allow them to dissociate 100% in water

ARRHENIUS ACIDS AND BASES THEORY

Arrhenius theory of Acids: acid is a chemical compound that increases the concentration of hydrogen ions (H^+) in an aqueous solution

However, H^+ does not exist freely in water. The H^+ ion bonds with H_2O molecule

$H^+_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)}$ → hydronium ion   the +ve H ion is attracted to −ve O ion

Ex. $HCl_{(aq)} = H^+ + Cl^-$ so... $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

These reactions are the same thing. The H^+ dissociates from the HCl, making it a decomposition reaction

Theory of Bases: a compound that increases the concentration of hydroxide ion (OH^-) in an aqueous solution

Ex. $NaOH \rightarrow Na^+ + OH^-$ (basic)

but... some reactions don't go with Arrhenius theory:

Ex. $NH_3 + H_2O \rightarrow$ might seem acidic, but doing a pH strip makes it basic. For those we use Bronsted-Lowry Theory

BRONSTED-LOWRY THEORY FOR ACIDS/BASES

an extension of the Arrhenius Theory

Acid: a proton $(H)^+$ donor
↳ if hydrogen atom looses an electron, it is only proton

Base: a proton $(H)^+$ acceptor → definition is more general

* water can act like an acid <u>OR</u> base depending on situation

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$
↳ the NH_3 is a proton accepter, so this is basic and there is more OH ion at the end

conjugate base: after a species has donated a proton

conjugate acid = after a species has accepted a proton

The NH_4^+ is a conjugated acid. It is still a basic solution because the pH has more OH^- ions

This reaction increased the OH^- concentration, so its still basic

NH_3 (base) + H_2O (acid) → NH_4^+ (conjugate acid) + OH^- (conjugated base)

when base accepts H^+, it is conjugate acid
when acid donates H^+, it is conjugate base
The NH_3 is strong, so solution = basic

HCl (acid) + NH_3 (base) $\xrightarrow{H^+}$ ... → NH_4^+ (conjugate acid) + Cl^- (conjugated base)

Since its a chemical reaction, it can go both ways. so, when going backwards, the NH_4^+ would donate H^+ to Cl^- ion

If it has OH^-, it is basic. if it is H_3O^+ it is acidic, when we don't see either on the product side, the H^+ ions usually donate to H_2O so usually acidic

protons are transferred from one reactant (acid) to the other (base)

strong acids have weak conjugate bases. They are opposites because its hard for the other to overcome the strength of the other

Ex. $HCl + H_2O \rightarrow Cl^- + H_3O^+$ HCl is a strong acid and will donate easily, but Cl^- will not take the H^+ ion as easily

Acids and bases in a reaction differ by the loss/gain of a proton

reactants → products   the → is a change in protons

H^+ ions are usually the only one being transfered. they have 1 pt

reverse the acids and bases for the products side H^+ ions go A → B (alphabetical order)

TYPES OF ACIDS

monoprotic = (1-proton) an acid that can donate 1 pt per molecule

polyprotic = (>1 proton) an acid that can donate more than 1 proton/molecule
↳ Diprotic = can donate 2 protons
Triprotic = can donate 3 protons . . .

for polyprotic acids ionization occurs in 2 stages:

1) $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$
2) $HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$

They donate more than 1 protons to other water molecules never the same one

Acids/ base reactions are not restricted to aqueous solutions
some species can act as acids/ bases depending on what the other reactant is

1 acid 7 7 7 base 14
each # x 10 ← → ÷10 each # > molarity of H^+

when strong acid + strong base react
ionic compound + water are made

SOLUTIONS AND NEUTRALIZATION

Autoionization of water = $H_2O + H_2O \rightarrow H_3O^+ + OH^-$

amphoretic = any species that can react as either an acid or a base (gain/ lose a proton) depending on the other reactant

at 25°C → $[H_3O^+]$ and $[OH^-]$ both are 1.0×10^{-7} M
pure water is neutral at any temp which changes the molarity, but they change at the same rate, so both would still be equal

in neutral solution $[H_3O^+] = [OH^-]$ [ ] = molarity/ concentration

in acidic solution $[H_3O^+] > [OH^-]$ in basic $[H_3O^+] < [OH^-]$

acidic and basic solutions happen when adding other substances to the whole

each jump in pH scale is 10x difference   pH = logarithmic scale
↳ pH of 1 is 10 times more acid than pH of 2

$$\begin{tikzcd}
{[H^+]} \arrow[r, shift left, "\frac{10^{-14}}{[H^+]}"] \arrow[d, shift left, "10^{-pH}"] & {[OH^-]} \arrow[l, shift left, "\frac{10^{-4}}{[OH^-]}"] \arrow[d, shift left, "-\log[OH^-]"] \\
pH \arrow[u, shift left, "-\log[H^+]"] \arrow[r, shift left, "14-pH"] & pOH \arrow[l, shift left, "14-pOH"] \arrow[u, shift left, "10^{-pOH}"]
\end{tikzcd}$$

Kw : dissociation concentration for water $= 1 \times 10^{-14}$

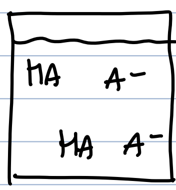
$[H^+]$ and $[OH^-]$ are inversely related

* new sig fig rule: # of sig figs = that many places past decimal point *

BUFFERS

buffer : made by combining a weak acid with conjugate base (or weak base with conjugate acid). A conjugate acid has 1 less H^+ ion than acid
help maintain pH of the system and are resistant to change

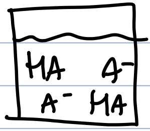
HA = weak acid
NaA : salt with same anion → turns to $Na^+ + A^-$
↳ salt of CB



Adding H^+ ions react with A^- ions and they become HA → weak acid so pH doesn't change

$\underset{A}{H^+} + \underset{CB}{A^-} \rightarrow \underset{\text{weak acid}}{HA}$ → still same acid

NaA don't combine again because it is strong conjugate base, so it will dissociate 100%.



adding OH^- will react with weak acid and create more conjugate base

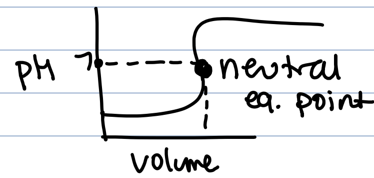
$$\underset{B}{OH^-} + \underset{weak}{HA} \rightarrow H_2O + \underset{CB}{A^-}$$

overcoming buffer capacity: pH will change after a certain level when there are no more ions for the H^+ or OH^- ions to react with

ACID - BASE TITRATION

Titration: a procedure for determining the amount of substance A (analyte) by adding a volume of a solution of known concentration B (titrant) until reaction is complete

used to find concentration of unknown acid/base



Equivalence Point: the point in a titration in which a stoichiometric amount of titrant has been added to the analyte
no more acid/base left

Endpoint: the point in a titration where the indicator changes color

Volumetric Analysis: method of analysis based on titration

for stoic. to find pH, look at the excess and find pH