

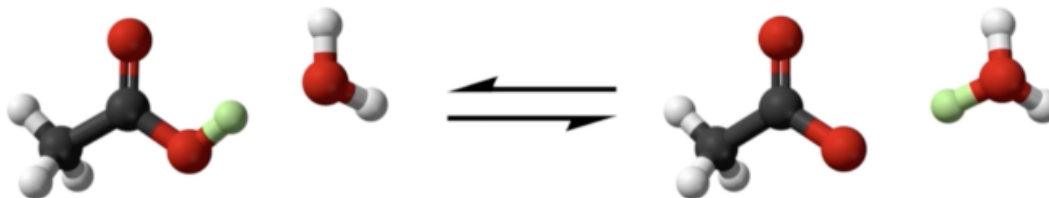


AP Chemistry Ultimate Review Packet

Unit 8: Acids and Bases

Bronsted-Lowry model for acids and bases deals with proton (H^+) transfer

- an *acid* is a proton donor
- a *base* is a proton acceptor
- an acid-base reaction is the donation of a proton from acid to base

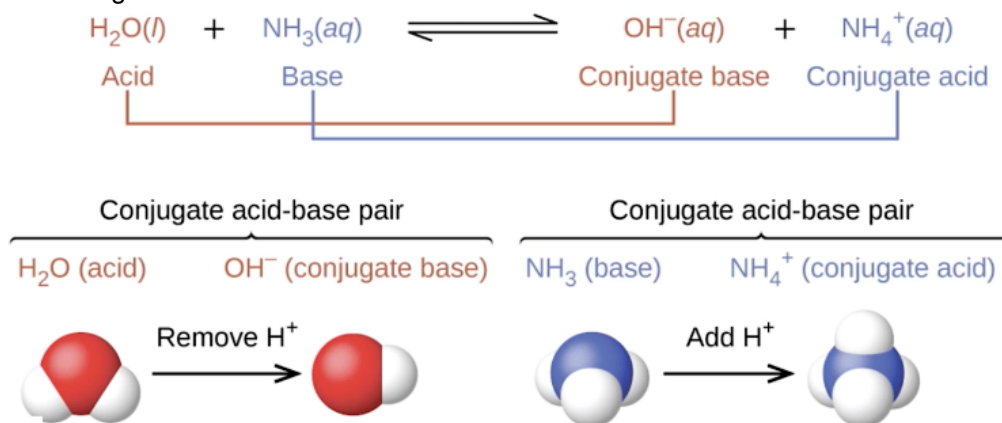


When an acid donates a proton it becomes its *conjugate base*

- must have a charge that is one less than the acid

When a base accepts a proton it becomes its *conjugate acid*

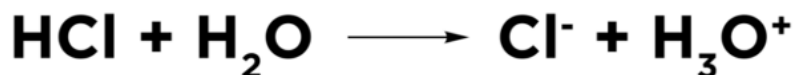
- must have a charge that is one more than the base



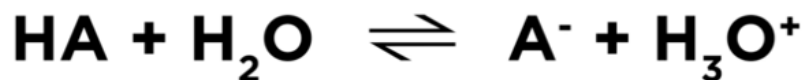
We should be able to identify acid/base/conjugate pairs as above

Lewis model for acids and bases deals with electron pair donation

- an *acid* is an electron pair acceptor
- a *base* is an electron pair donor
- allows for acidic behavior in compounds without hydrogen atoms



A **strong acid** (like HCl) will ionize entirely in solution to produce hydronium ions



A **weak acid** (represented as HA) will ionize partially, establishing an equilibrium between two forms

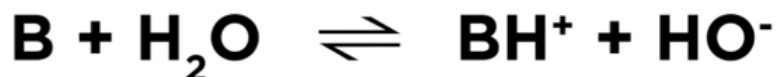
$$K_a = \frac{[A^-][H_3O^+]}{[HA]}$$

This equilibrium can be described by the **acid dissociation constant (K_a)**



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A **weak base** will ionize partially, establishing an equilibrium between two forms

$$K_b = \frac{[BH^+][HO^-]}{[B]}$$

This equilibrium can be described by the **base dissociation constant (K_b)**. We can calculate K_a/K_b values given equilibrium concentrations.

Some species are **amphoteric** (can act as both an acid and a base)

- $H_2O \rightarrow HO^-$ (when acting as an acid)
- $H_2O \rightarrow H_3O^+$ (when acting as a base)

We can predict **acid/base strength** using a few guidelines:

- The strength of an acid is proportional to the stability of its conjugate base
 - more stable conjugate base means a *stronger acid* (easy to lose a proton)
 - less stable conjugate base means a *weaker acid* (hard to lose a proton)
- acidity increases going down the periodic table
 - larger resulting anion can better accommodate the negative charge
 - stability: $I^- > Br^- > Cl^- > F^-$
 - acidity: $HI > HBr > HCl > HF$
- acidity increases moving to the right along a period
 - more electronegative elements can better accommodate negative charge
 - acidity: $HF > H_2O > NH_3 > CH_4$
- oxyacid strength increases with electronegativity of central atom
- conjugate bases with resonance stabilization correlate with stronger acids

Understanding pH and pOH

We must be able to do calculations using **pH and pOH**

pH = $-\log[H_3O^+]$ or an easier way of expressing hydronium concentration

pOH = $-\log[HO^-]$ or an easier way of expressing hydroxide concentration

also: $[H_3O^+] = 10^{-pH}$ and $[HO^-] = 10^{-pOH}$

These parameters are associated with the **autoionization of water**



- this means even pure water ionizes to a small extent at room temperature
- $K_w = [H_3O^+][HO^-]$ so $[H_3O^+] = \sqrt{1.0 \times 10^{-14}} = 1.0 \times 10^{-7}$
- pH of pure water at room temperature = $-\log(1.0 \times 10^{-7}) = 7$
 - pH = 7 means a *neutral* solution: $[H_3O^+] = [HO^-]$
 - pH < 7 means an *acidic* solution: $[H_3O^+] > [HO^-]$
 - pH > 7 means a *basic* solution: $[H_3O^+] < [HO^-]$

We can use this information and the equations above to calculate any of these values for a given solution.



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A **neutralization reaction** involves an acid and base reacting to form water and a salt.

Titration uses the stoichiometry of a neutralization reaction to get information about the concentration of an acid or base when it is unknown.

- the *analyte* is the substance of unknown concentration
- the *titrant* is the substance neutralizing the analyte
- volume of titrant added is closely monitored
- eventually the **equivalence point** is reached
 - a precise stoichiometric amount of titrant has been added
 - a substance called an indicator will glow a vibrant color
 - this means the analyte has been precisely neutralized
 - use volume of titrant to find moles and then concentration of analyte
- titrations with a strong acid and strong base have their equivalence point at pH = 7
- weak acid and strong base, EP at pH > 7; strong acid and weak base, EP at pH < 7

Buffers

A **buffer solution** is one that has the ability to resist changes in pH

- comprised of a weak acid or base and its conjugate in roughly equal amounts
- if acid is added to solution, it reacts with the base in the buffer
 - Le Chatelier's principle restores the reacted base
- if base is added to solution, it reacts with the acid in the buffer
 - Le Chatelier's principle restores the reacted acid

All buffers have a **buffer capacity**. Once all of the weak acid or base in the solution is reacted it can no longer neutralize any incoming strong acid or base and the pH will change rapidly

- more acid/base in buffer solution → maintains the same pH
- more acid/base in buffer solution → greater buffer capacity
- once the pH changes by around **1 pH unit** we can say that the buffer capacity of the buffer solution has been exhausted

$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

Relevant to buffer solutions is the **Henderson-Hasselbalch equation**

- allows us to calculate the pH of a solution given acid/base concentrations
 - must know the pKa of the acid ($\text{pK}_a = -\log K_a$)
- allows one to prepare buffer solutions of any pH
- if the acid/base and its conjugate are present in equal amounts, the logarithmic term will reduce to zero and thus disappear from the equation ($\log 1 = 0$)
 - this leaves **pH = pKa**
 - so when preparing a buffer solution we simply select a weak acid with a pKa equal to the desired pH of the solution and utilize it in equal amounts with its conjugate to prepare the solution