

Acid Base Equilibrium

General Acid/Base properties

Acids: taste sour (think vinegar)
React w/ active metals --> produce $H_{2(g)}$
Conduct electrical current
Turn litmus paper red
Reacts w/ base (neutralization)

Acid properties due to H^+ (H_3O^+) in solution

Bases: taste bitter (think soap)
Feels slippery
Conducts electrical current
Turns litmus blue
Reacts w/ acid (neutralization)

Basic properties due to OH^- in solution

Definitions:

Arrhenius

- Acids produce H^+ in solution (Chem 11)
- Base produce OH^-

Brønsted-Lowry

- Acid: proton (H^+) donor
- Base: proton acceptor

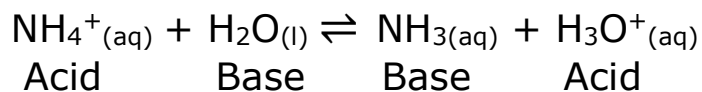
These can be ions (charged) OR molecules (non-charged)
Substances can also be **amphiprotic** → can act as an acid OR base

Amphiprotic: NH_3 , H_2O , HCO_3^- etc... (Both has H, and can accept H)

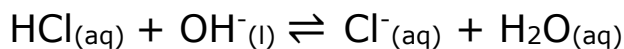
Conjugate acids & bases:

Conjugate acid/base pairs differ only by the number of hydrogens, the one with 1 MORE hydrogen is the acid, the one with 1 LESS hydrogen is the base. Conjugate pairs exist on either side of an acid-base equilibrium.

Ex.

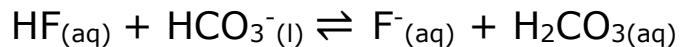


Ex. Identify the conjugates in the following

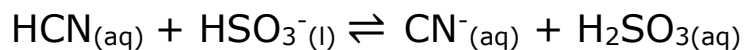


Water is amphiprotic!

Ex.



Ex.



Hebden pg 115-118 – READ, pg 118-120 do Q's 13-19

Relative Strength of acid/bases

Concentration: # of moles (acid/base) in one liter of solution
 $20 \text{ M} \rightarrow 1 \times 10^{-10} \text{ M}$

Generally:

> 5M = concentrated

< 5M = dilute

Strength: describes how readily a substance dissociates into ions.

Ex. $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$ 100%

$\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^-$ $\ll 100\%$

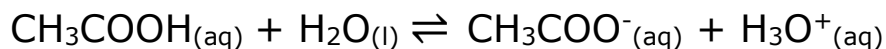
100% dissociation = strong, less than 100% = weak

Strong acids – dissociate into H_3O^+ 100%

Ex. HClO_4 , HI , HBr , HCl , HNO_3 , H_2SO_4

Weak acids – all other acids

Forms an equilibrium b/w un-dissociated acid and its ions.



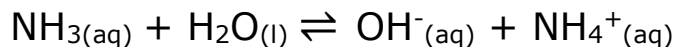
Usually \leftarrow towards reactants

Strong bases – there are 3 strong bases that you need to know about:

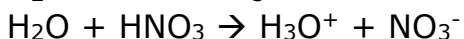
NH_2^- , O^{2-} , OH^- (Usually **OH^-** forms in aqueous solutions)



Weak bases – all remaining bases (on the table) are weak they form an equilibrium.

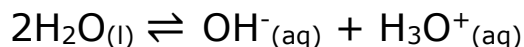


Levelling effect: all strong acids and strong bases are equal



in water the strongest acid is **H_3O^+**

K_w – the Water Constant



$$K_w = 1.0 \times 10^{-14} = [\text{OH}^-][\text{H}_3\text{O}^+] - \text{ @ } 25^\circ\text{C}$$

Ex. What is the $[\text{OH}^-]$ in a 0.005M solution of HCl?

Ex. What is the $[\text{H}_3\text{O}^+]$ in a 0.0000007 M solution of NaOH?

pH – a common way to represent concentration of an acid

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \text{ or } -\log [\text{H}^+]$$

As $[\text{H}^+]$ increases, pH decreases

Because pH is log base 10, means every pH level is 10x stronger

Ex. Is pH 2 stronger or weaker than pH 5, by how much?

Ex. What is the pH of a 5.2×10^{-5} M solution of HCl?

Reminder: SigFigs w/ pH are weird

$$\frac{\underline{1.00} \times 10^{-3}}{[\quad]} \rightarrow \underline{3.000} \text{ pH}$$

pH examples:

Ex. What's the pH of 4.20 moles of HCl in 3.15L?

Ex. What's the pH of 0.000375 moles of HNO₃ in 6235mL?

pOH – like pH but with OH

Ex. If I take 15mL of concentrated HCl (12M) and dilute it to 6L what's the pOH?

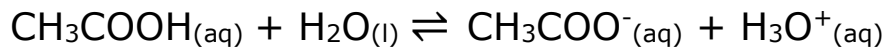
Ex. What's the pH and pOH of of a 4.23×10^{-8} M solution of HClO₄?

Hebden pg 134 – 143 try all questions

Quantitative comparison of acid strength (K_a)

Acid dissociation of weak acids usually lies to the reactant side.

Ex.



We can calculate how much by using **K_a**

$$K_a(\text{CH}_3\text{COOH}) = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} = 1.8 \times 10^{-5}$$

A K_a of 1x10⁻² is 1000 times larger than a K_a of 1x10⁻⁵

From the K_a we can calculate [H⁺] and pH values

Ex. What's the pH of a 0.500 M HNO₂ solution?

Reaction	HNO ₂ ⇌	H ⁺ +	NO ₂ ⁻
[Initial]			
Change			
[Equilibrium]			

We can ASSUME X is small, when compared to []

Ex. What's the pH of 0.0020 M ammonium?

Reaction	$\text{NH}_4^+ \rightleftharpoons$	$\text{H}^+ +$	NH_3
[Initial]			
Change			
[Equilibrium]			

Ex. If a solution of HF has a pH of 4.2 what is its concentration?

Base constant (K_b)

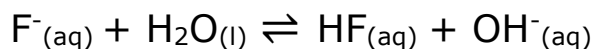
ALWAYS Have to calculate: cannot be taken from the table

Recall: $K_w = K_a \times K_b$

$$\text{So... } K_b = \frac{K_w}{K_a}$$

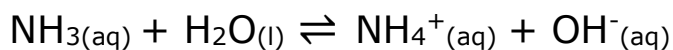
You MUST use the conjugate acid's K_a

Recall that the Brønsted-Lowry definition of a base means that it is a proton acceptor. So that the base will gain an H^+ from reactants to products.



$$\text{So } K_b = \frac{[HF][OH^-]}{[F^-]} = \frac{K_w}{K_a}$$

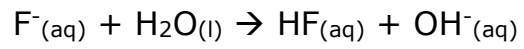
Ex. What's the K_b of NH_3 ?



Ex. What is the pH of a 0.200M F^- solution?

Reaction	F^-	$+ H_2O$	\rightleftharpoons	HF	$+$	OH^-
[Initial]						
Change	-					
[Equilibrium]						

Ex. What's the $[F^-]$ in water if water has a pH of 7.62?



* * * * *

Hebden pg 153/4 #'s 84-93