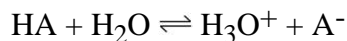


## Acid-Base Equilibria

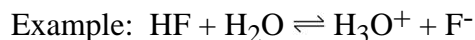
## BLUFFER'S GUIDE

1.  $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^-$   $K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$   
 $\text{pH} = -\log[\text{H}^+]$   $\text{pH} + \text{pOH} = 14$   $[\text{H}^+] = 10^{-\text{pH}}$   
 Convert between pH, pOH,  $[\text{H}^+]$ , &  $[\text{OH}^-]$

2. Acid Ionization Constant ( $K_a$ ):



$$K_a = [\text{A}^-][\text{H}_3\text{O}^+]/[\text{HA}]$$



$$K_a = [\text{F}^-][\text{H}_3\text{O}^+]/[\text{HF}]$$

3. Typical question: Given  $K_a$  and the starting concentrations of acid, find concentrations (or pH) of  $[\text{H}^+]$  at equilibrium.

Example:  $K_a$  for acetic acid =  $1.8 \times 10^{-5}$ .

Find the pH of 0.100M acetic acid.

4. Polyprotic Acids:  $\text{H}_3\text{PO}_4$ ,  $\text{H}_2\text{SO}_4$ ,  $\text{H}_2\text{C}_2\text{O}_4$ , etc.

The 1<sup>st</sup> dissociation is strong for  $\text{H}_2\text{SO}_4$ .

When using Hess's Law with a polyprotic acid:

$$K_{\text{Overall}} = K_{a1} \times K_{a2}$$

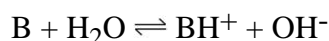
Calculating pH, use  $K_{a1}$

5. Bronsted-Lowry Definitions.

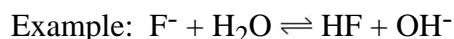
Acids =  $\text{H}^+$  donors; Bases =  $\text{H}^+$  acceptors

Conjugate acid-base pairs.

6. Base Ionization Constant ( $K_b$ ):

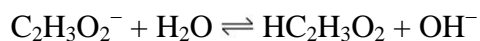


$$K_b = [\text{BH}^+][\text{OH}^-]/[\text{B}]$$

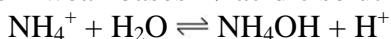


$$K_b = [\text{HF}][\text{OH}^-]/[\text{F}^-]$$

7. Salt solns can have pH's  $\neq 7$  (hydrolysis)  
 ions from weak acids  $\rightarrow$  basic solutions



ions from weak bases  $\rightarrow$  acidic solutions



8.  $K_a \times K_b = K_w = 10^{-14}$

only applies for **conjugate** acids & bases!

Example:  $K_a \text{ HC}_2\text{H}_3\text{O}_2 = 1.8 \times 10^{-5}$

$K_b \text{ C}_2\text{H}_3\text{O}_2^- = 10^{-14} / 1.8 \times 10^{-5}$

9. Percent ionization =

$$[\text{H}^+]_{\text{equilibrium}} / [\text{HA}]_{\text{initial}} \times 100$$

10. Acid Strength-know the 6 strong acids: HCl, HBr, HI,  $\text{HNO}_3$ ,  $\text{HClO}_4$ , and  $\text{H}_2\text{SO}_4$  (removal of the first  $\text{H}^+$  only)

(a) binary acids - acid strength increases with increasing size and electronegativity of the "other element". (NOTE: Size predominates over electronegativity in determining acid strength.)

Examples:  $\text{H}_2\text{Te} > \text{H}_2\text{O}$  &  $\text{HF} > \text{NH}_3$

(b) Oxoacids - Acid strength increases with increasing:

(1) electronegativity

(2) number of bonded oxygen atoms

(3) oxidation state of the "central atom".

Example:  $\text{HClO}_4$  or  $[\text{O}_3\text{Cl}(\text{OH})]$

is very **acidic**

$\text{NaOH}$  is very **basic**

Acid strength also increases with *decreasing* radii of the "central atom".

Example:

$\text{HOCl}$  (bond between Cl and OH is covalent--making  $\text{HOCl}$  **acidic**)

$\text{HOI}$  (bond between I and OH is ionic--making  $\text{HOI}$  **basic**)

11. Lewis Acids and Bases:

(This applies to coordinate covalent bonds.)

Lewis Acid--electron pair acceptor

Lewis Base--electron pair donor

"Have Pair... Will Share" – Lewis Base

In complex ion formation, metal ions are Lewis acids, and ligands are Lewis bases.

Example:  $\text{Cu}^{2+} + 4\text{NH}_3 \rightleftharpoons \text{Cu}(\text{NH}_3)_4^{2+}$

$\text{Cu}^{2+}$  acts as an acid;  $\text{NH}_3$  acts as a base.

12. Strong Bases: amide ion,  $\text{NH}_2^-$

hydride ion,  $\text{H}^-$ , methoxide ion,  $\text{CH}_3\text{O}^-$