

15 • Acid–Base Chemistry

BLUFFER'S GUIDE

- $\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^- \quad K_w = [\text{H}^+][\text{OH}^-] = 10^{-14}$
 $\text{pH} = -\log[\text{H}^+] \quad \text{pH} + \text{pOH} = 14 \quad [\text{H}^+] = 10^{-\text{pH}}$
 Convert between pH, pOH, $[\text{H}^+]$, & $[\text{OH}^-]$
- Acid Ionization Constant (K_a):
 $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$
 $K_a = [\text{A}^-][\text{H}_3\text{O}^+]/[\text{HA}]$
 Example: $\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-$
 $K_a = [\text{F}^-][\text{H}_3\text{O}^+]/[\text{HF}]$
- Typical question: Given K_a and the starting concentrations of acid, find concentrations (or pH) of $[\text{H}^+]$ at equilibrium.

$$[\text{H}^+] = [\text{A}^-] = \sqrt{K_a[\text{HA}]_0}$$

 Example: K_a for acetic acid = 1.8×10^{-5} .
 Find the pH of 0.100M acetic acid.
- Polyprotic Acids: H_3PO_4 , H_2SO_4 , $\text{H}_2\text{C}_2\text{O}_4$, etc. The 1st dissociation is strong for H_2SO_4 .
 When using Hess's Law with a polyprotic acid: $K_{\text{overall}} = K_{a1} \times K_{a2}$
 Calculating pH, $[\text{H}^+]$ dominated by K_{a1}
- Bronsted–Lowry Definitions.
 Acids = H^+ donors; Bases = H^+ acceptors
 Conjugate acid–base pairs.
- Base Ionization Constant (K_b):
 $\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{OH}^-$
 $K_b = [\text{BH}^+][\text{OH}^-]/[\text{B}]$
 Example: $\text{F}^- + \text{H}_2\text{O} \rightleftharpoons \text{HF} + \text{OH}^-$
 $K_b = [\text{HF}][\text{OH}^-]/[\text{F}^-]$

$$[\text{OH}^-] = [\text{BH}^+] = \sqrt{K_b[\text{B}]_0}$$
- Salt sol'ns can have pH's $\neq 7$ (hydrolysis)
 anions from weak acids \rightarrow basic solutions
 $\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$
 cations from weak bases \rightarrow acidic solutions
 $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4\text{OH} + \text{H}^+$
- $K_a \times K_b = K_w = 1.0 \times 10^{-14}$
 only applies for **conjugate** acids & bases!

 Example: $K_a(\text{HC}_2\text{H}_3\text{O}_2) = 1.8 \times 10^{-5}$
 $K_b(\text{C}_2\text{H}_3\text{O}_2^-) = 1.0 \times 10^{-14} / 1.8 \times 10^{-5}$
- Percent ionization =

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

$$\frac{[\text{OH}^-]_{\text{equilibrium}}}{[\text{B}]_{\text{initial}}} \times 100$$
- Acid Strength—know the 6 strong acids: HCl, HBr, HI, HNO_3 , HClO_4 , and H_2SO_4 (removal of the first H^+ only)
 - Binary acids:
 Acid strength increases with...
 - ... increasing electronegativity of the “other element”. (Period trend.)
 Ex: $\text{HF} > \text{NH}_3$
 - ... increasing size & *decreasing* bond strength (group trend, more important than period trend)
 Ex: $\text{H}_2\text{Te} > \text{H}_2\text{O}$
 - Oxoacids:
 Acid strength increases with increasing...
 - ... electronegativity or decreasing radius of the central atom for same # O atoms as the O–H bond becomes more polar.
 Ex. $\text{H}_2\text{SeO}_4 > \text{H}_3\text{AsO}_4$, $\text{HClO}_3 > \text{HBrO}_3$
 - ... number of bonded oxygen atoms (oxidation state of the central atom) as the O–H bond becomes more polar.
 Ex. $\text{HBrO}_4 > \text{HBrO}_3$
 Recall: X-O-H is acidic when X is a NM or weak metal (high EN) but basic with a strong metal (low EN)
- 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: $\text{Cu}^{2+} + 4\text{NH}_3 \rightleftharpoons \text{Cu}(\text{NH}_3)_4^{2+}$
 Cu^{2+} acts as an acid; NH_3 acts as a base.
- Strong Bases: hydroxides of alkali metals and Ca^{2+} , Sr^{2+} and Ba^{2+} ; amide ion, NH_2^- ; hydride ion, H^- , methoxide ion, CH_3O^-