

17 • Acid-Base Equilibria

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):



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



$$K_a = \frac{[\text{H}^+][\text{F}^-]}{[\text{HF}]}$$

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



Find the pH of 0.100M acetic acid.

4. 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: $K_{a(\text{total})} = K_{a1} \times K_{a2}$
 Calculating pH , use K_{a1}

5. Brønsted-Lowry Definitions:
 Acids = H^+ donors; Bases = H^+ acceptors
 Conjugate acid-base pairs.

6. Base Ionization Constant (K_b):



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



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

7. Salt solts can have $\text{pH} \neq 7$ (hydrolysis):
 ions from weak acids $\rightarrow$ basic solutions
 $\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$
 ions from weak bases $\rightarrow$ acidic solutions
 $\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{NH}_3 + \text{H}^+$

8. $K_a \times K_b = K_w = 10^{-14}$
 only applies for conjugate acids & bases!
 Example: $K_a \text{ HC}_2\text{H}_3\text{O}_2 = 1.8 \times 10^{-5}$
 $K_b \text{ C}_2\text{H}_3\text{O}_2^- = 10^{-14} / 1.8 \times 10^{-5}$

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9. Percent Ionization =
 $\frac{[\text{H}^+]_{\text{equilibrium}}}{[\text{HA}]_{\text{initial}}} \times 100$
10. Acid Strength: lower the K_a stronger acids: HCl , HBr , HI , HNO_3 , HClO_4 , and H_2SO_4 (removed of the first H^+ only)

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



Ex. Oxacids: Acid strength increases with increasing:

- (i) electronegativity
- (ii) number of bonded oxygen atoms
- (iii) oxidation state of the "central atom".



is very acidic

NaOH is very basic

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

Example:

HOCl (bond between Cl and OH is covalent—making HOCl acidic)

HCl (bond between Cl and OH is ionic—making HCl basic)

11. Lewis Acids and Bases:

(This applies to coordination covalent bonds.)

Lewis Acid—electron pair acceptor

Lewis Base—electron pair donor

"Have Pair, ... Be A Base" = Lewis Base

Ex. examples: ion formation, metal ion-ate Lewis acids, and ligands are Lewis bases.



Ca^{2+} acts as an acid; NH_3 acts as a base.

12. Strong Basic: anions like NH_2^-

hydride ions, H^- , methoxide ions, CH_3O^-

Based on a list by Alfred Reed, Inorganic 92

Acid Base Equilibria Bluffer S Guide Answer

Thomas Bond, Chris Hughes



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