1. H2O  H+ + OH Kw = [H+][OH] = 1014

 pH = -log[H+] pH+pOH = 14 [H+] =10pH

 Convert between pH, pOH, [H+], & [OH]

1. Acid Ionization Constant (Ka):

HA + H2O  H3O+ + A-

Ka = [A-][H3O+]/[HA]

Example: HF + H2O  H3O+ + F-

Ka = [F-][H3O+]/[HF]

1. Typical question: Given Ka and the starting concentrations of acid, find concentrations (or pH) of [H+] at equilibrium.

Example: Ka for acetic acid = 1.8 x 10-5.

Find the pH of 0.100M acetic acid.

1. Polyprotic Acids: H3PO4, H2SO4, H2C2O4, etc. The 1st dissociation is strong for H2SO4.

 When using Hess’s Law with a polyprotic acid: Koverall = Ka1 x Ka2

 Calculating pH, use Ka1

1. Bronsted-Lowry Definitions.

 Acids = H+ donors; Bases = H+ acceptors

 Conjugate acid-base pairs.

1. Base Ionization Constant (Kb):

B + H2O  BH+ + OH-

Kb = [BH+][OH-]/[B]

Example: F- + H2O  HF + OH-

Kb = [HF][OH-]/[F-]

1. Salt solns can have pH’s ≠ 7 (hydrolysis)

 ions from weak acids → basic solutions

C2H3O2 + H2O  HC2H3O2 + OH

 ions from weak bases → acidic solutions

NH4+ + H2O  NH4OH + H+

1. Ka x Kb = Kw = 10-14

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

Example: Ka HC2H3O2 = 1.8 x 10-5

 Kb C2H3O2- = 10-14 / 1.8 x 10-5

1. Percent ionization =

[H+]eqilibrium /[HA]initial x 100

1. Acid Strength-know the 6 strong acids: HCl, HBr, HI, HNO3, HClO4, and H2SO4 (removal of first 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: H2Te > H2O & HF > NH3

**(b)** Oxoacids - Acid strength increases with increasing:

 (1) electronegativity

 (2) number of bonded oxygen atoms

 (3) oxidation state of the “central atom”.

Example: HClO4 or [O3Cl(OH)]

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**)

HOI (bond between I and OH is ionic--making HOI **basic**)

1. 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: Cu2+ + 4NH3  Cu(NH3)42+

Cu2+ acts as an acid; NH3 acts as a base.

1. Strong Bases: amide ion, NH2

 hydride ion, H, methoxide ion, CH3O

*Based on a handout by William Bond, Snohomish HS*

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