Acids and Bases V

In class question: Dissolve $(NH_4^+)SO_4$, in water. What is pH?

$$(NH_4)_2SO_4\,(aq) \rightarrow 2NH_4^+\,(aq) + SO_4^{2-}\,(aq)$$

- NH_4^+ is the conjugate acid of NH_3, a weak base. Thus, a weak acid.
- SO_4^{2-} is the conjugate base of HSO_4^-, a weak acid. Thus, a weak base.

$K_a > K_b$, so $pH < 7$

K_b for $NH_3 = 1.8 \times 10^{-5}$

$K_a \cdot K_b = K_w$

K_a for $NH_4^+ = \frac{K_w}{K_b} = 5.6 \times 10^{-10}$

K_a for $HSO_4^- = 1.01 \times 10^{-2}$

So, K_b for ~~HSO_4^-~~ $SO_4^{2-} = 9.9 \times 10^{-13}$

Dissolve Ammonium acetate in water. What is pH?

$$NH_4OAc\,(aq) \rightarrow NH_4^+\,(aq) + OAc^-\,(aq)$$

- NH_4^+ is the conjugate acid of NH_3, a WB. Thus, a WA.
- OAc^- is the conjugate acid of acetic acid, a ~~WB~~ WA. Thus a WB

$K_a = K_b$, so $pH = 7$

Dissolve $NaHCO_3$ in water. What is pH?

$$NaHCO_3\,(aq) \rightarrow Na^+\,(aq) + HCO_3^-\,(aq)$$

- Na^+ does not gain or accept H^+
- HCO_3^- is the conjugate base of H_2CO_3 a WA. But it's also an acid. (Polyprotic Acid)

~~H_2CO_3~~ $HCO_3^-\,(aq) + H_2O \rightleftharpoons H_2CO_3\,(aq) + OH^-$ $\quad K_b = 2.4 \times 10^{-8}$

$HCO_3^-\,(aq) + H_2O \rightleftharpoons CO_3^{2-}\,(aq) + H_3O^+$ $\quad K_a = 4.8 \times 10^{-11}$

$K_b > K_a$ so, base

Dissolve ferric chloride ($FeCl_3$) in water. What is pH?

Hydrated Cations

- Transition and post-transition metals form hydrated cations.
- H_2O attached to a metal ion is more acidic than free H_2O, making hydrated metal ions acidic.

(Ex) Al^{3+} or Fe^{3+} w/ six H_2O molecules bound

- O-H bonds are weakened by metal binding, leading to release of H^+. The metal ion stabilizes the resulting negative charge on the OH^-.

Lewis Acid / Base Chemistry

- Metal ions are Lewis Acids
  - Lewis acids are electron pair acceptors
- Water is a Lewis Base
  - Lewis bases are electron pair donors.

$$\underset{\text{Lewis Acid}}{A^+} + \underset{\text{Lewis Base}}{B^-} \longrightarrow \underset{\substack{\text{coordinate} \\ \text{covalent bond}}}{A-B}$$

(Ex) $Fe^{3+}(aq) + 6H_2O \rightarrow [Fe(H_2O)_6]^{+3}$

- Coordination Compound - Lewis acid-base complex of a metal cation (the Lewis Acid) associated with a group of neutral molecules or anions (the Lewis base aka the "ligand").

Factors that Affect Acid Strength

H—A

① Polarity: H—A bond must be polarized with δ^+ on H atom and δ^- on the A atom.

② Bond Strength: Weaker bonds can be broken more easily, making the acid stronger.

Optimal acid has highly polar but weak bonds.

Ex 1) CH_4

Polarity: C—H bond is not polar

Bond Strength: C—H bond is strong.

CH_4 is NOT an acid

2) HCl

Polarity: H—Cl bond is very polar

Bond Strength: H—Cl bond is pretty strong, but polarity dominates.

HCl is a good acid.

Binary Acids

- Binary acids consist of H and one other element
- Acid strength generally increase down a group
- Acid strength generally increases across a period

Oxyacids

- Oxyacids consist of H, O, and other nonmetal.
- Generally, as the electronegativity of the ~~metal~~ nonmetal increases, the acidity increases for acids with the same structure.

Families of Oxyacids

- If an element can form more than one oxyacid, the oxyacid with more O atoms is more acidic.

Carboxylic Acids

- Carboxylic acids are organic acids containing $-COOH$ group
- Factors contributing to their acidic behavior:
  - The other oxygen atom draws electron density from the O–H bond, increasing polarity
  - The conjugate base (carboxylic anion) has resonance forms to stabilize the anion.