

Salts and Molecular Structure

Hydrated metal ions

- Charged metals ions also produce an acidic solution.
- The metal itself does not act as a Brønsted-Lowry acid, but instead forms a hydrate that acts as a Brønsted-Lowry acid.
- Typically the higher the charge on the metal ion, stronger the acidity of the hydrated ion.
- $\text{Cu}^{2+} + 5 \text{H}_2\text{O} \rightarrow \text{Cu}(\text{H}_2\text{O})_6^{2+}$
- $\text{Cu}(\text{H}_2\text{O})_6^{2+} \rightarrow \text{CuOH}(\text{H}_2\text{O})_5^{2+} + \text{H}^+$

Acidity of hydrated metal ions

- M = metal
- For M^{n+} , a small, highly charged ion,
- $\text{M}(\text{H}_2\text{O})_x^{n+} + \text{H}_2\text{O} \rightleftharpoons \text{M}(\text{H}_2\text{O})_{x-1}\text{OH}^{(n-1)+} + \text{H}_3\text{O}^+$
- Calculate the pH of a 0.010 M AlCl_3 solution. The K_a value for $\text{Al}(\text{H}_2\text{O})_6^{3+}$ is 1.4×10^{-5} .

Salts of weak acids and bases

- For salts from a weak acid and weak base, the K_a and K_b of the ions must be compared.
- If $K_a > K_b$ the solution will be acidic
- If $K_a < K_b$ the solution will be basic
- If $K_a = K_b$ the solution will be neutral

Problem

- Predict whether an aqueous solutions of each of the following salts will be acidic, basic, or neutral.
- $\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$
- NH_4CN
- $\text{Al}_2(\text{SO}_4)_3$
- NH_4F
- $K_a \text{ Al}(\text{H}_2\text{O})_6^{3+} = 1.4 \times 10^{-5}$
- $K_a \text{ HF} = 7.2 \times 10^{-4}$
- $K_a \text{ HSO}_4^- = 1.2 \times 10^{-2}$
- $K_b \text{ C}_2\text{H}_3\text{O}_2^- = 5.6 \times 10^{-10}$
- $K_b \text{ CN}^- = 1.6 \times 10^{-5}$
- $K_b \text{ NH}_3 = 1.8 \times 10^{-5}$

Molecular Trends in Acid Strength

Trends in Binary Nonmetal Hydride Activity

- This applies to nonmetals only
- From left to right, electronegativity increases, polarity increases. H^+ removal is easier, and acidity increases.
- From top to bottom, bond length increases, bond strength decreases, H^+ removal is easier, and acidity increases.

Halogen Strong Acids

- All halogens are strong acids except fluorine.
- All halogens have a high electronegativity. The short bond length of fluorine makes it a weak acid.

Oxyacids

- With the same number of oxygen atoms around a central nonmetal (E), acid strength increases with the electronegativity of E.
- $\text{HClO} > \text{HBrO} > \text{HIO}$
- With different numbers of oxygen atoms around E, acid strength increases with the number of oxygen atoms.
- This is because oxygen atoms pull the electron density away from E, and H^+ is more easily removed.

Table 24.8 Several Series of Oxyacids and Their K_a Values

Oxyacid	Structure	K_a Value
HClO ₄		Large (~10 ⁷)
HClO ₃		~1
HClO ₂		1.2×10^{-2}
HClO		3.5×10^{-8}
H ₂ SO ₄		Large
H ₂ SO ₃		1.5×10^{-2}
HNO ₃		Large
HNO ₂		4.0×10^{-4}

Strong oxyanion acids

- ⊃ If there are at least 2 more oxygen atoms than hydrogen atoms, the acid will be strong.
- ⊃ H₂SO₄, HNO₃, HClO₃, HClO₄

Leveling Effect

- ⊃ It is difficult to differentiate the strengths of **strong acids** or bases because of water.
- ⊃ Water can act like an acid or a base.
- ⊃ By LeChâtelier's principle, the stronger the acid or base, the more water will shift to neutralize it.
- ⊃ Weak acids or bases are limited by their conjugate bases or acids, **strong acids and base are limited by water.**

Leveling effect

- ⊃ Water "levels" all strong acids (or bases) to an equal strength (HCl, HBr, and HI).
- ⊃ To determine the strength, K_a , of strong acids you have to dissolve it in an acid (weaker base than water), such as acetic acid, where different levels of protonation can occur.

Acidic Oxides (Acid Anhydrides)

- ⊃ Nonmetal oxides
 - O-X bond is strong and covalent.
 - SO₂, NO₂, CO₂
- ⊃ When H-O-X grouping is dissolved in water, the O-X bond will remain intact. It will be the polar and relatively weak H-O bond that will tend to break, releasing a proton.
- ⊃ SO₂ + H₂O → H₂SO₃
- ⊃ H₂SO₃ + H₂O ⇌ H⁺ + HSO₃⁻

Basic Oxides (Basic Anhydrides)

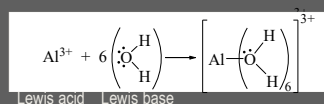
- ⊃ For metal oxides
 - O-X bond is ionic.
 - K₂O, CaO
- ⊃ If X has a very low electronegativity, the O-X bond will be ionic and subject to being broken in polar water, producing a basic solution.
- ⊃ K₂O + H₂O → KOH

Lewis Acids and Bases

- ⊃ This is the broadest definition.
- ⊃ An *acid* is an **electron pair acceptor**.
- ⊃ A *base* is an **electron pair donor**.
- ⊃ This is useful for: electron deficient compounds, polar double bonded compounds, metal ions, and protons (H⁺).

Lewis Acids and Bases

- ⊃ Lewis acid: *electron pair acceptor*
- ⊃ Lewis base: *electron pair donor*



Acid base models

Model	Definition of Acid	Definition of Base
Arrhenius	H ⁺ producer	OH ⁻ producer
Bronsted-Lowry	H ⁺ donor	H ⁺ acceptor
Lewis	Electron-pair acceptor	Electron-pair donor

Acid-base Equilibria

Chapter 15

Common Ion

- ⇒ A *common ion* is one that is both dissolved in an aqueous solution and introduced to the existing equilibrium system.
- ⇒ The shift in equilibrium as a result of the introduction of a common ion is called the *common ion effect*.
- ⇒ This is an application of Le Châtelier's Principle.

Acidic Solutions Containing Common Ions.

- ⇒ The equilibrium concentration of H_3O^+ in a 1.0 M HF solution is 2.7×10^{-2} M, and the percent dissociation of HF is 2.7%.
- ⇒ Calculate $[\text{H}_3\text{O}^+]$ and the percent dissociation of HF in a solution containing 1.0 M HF ($K_a = 7.2 \times 10^{-4}$) and 1.0 M NaF. *assume NaF dissociates at 100%