Factors impacting acid strength

Basic concept: easier to lose H, stronger acid

Charge:

- Higher charge = stronger bond to H^+ = weaker acid
- Ex: $HAsO_4^{2-}$ weaker than $H_2AsO_4^-$ weaker than H_3AsO_4

Bond length:

- Shorter bonds are stronger so H^+ harder to remove
- Ex: HF weaker than HCl weaker than HBr

Electronegativity:

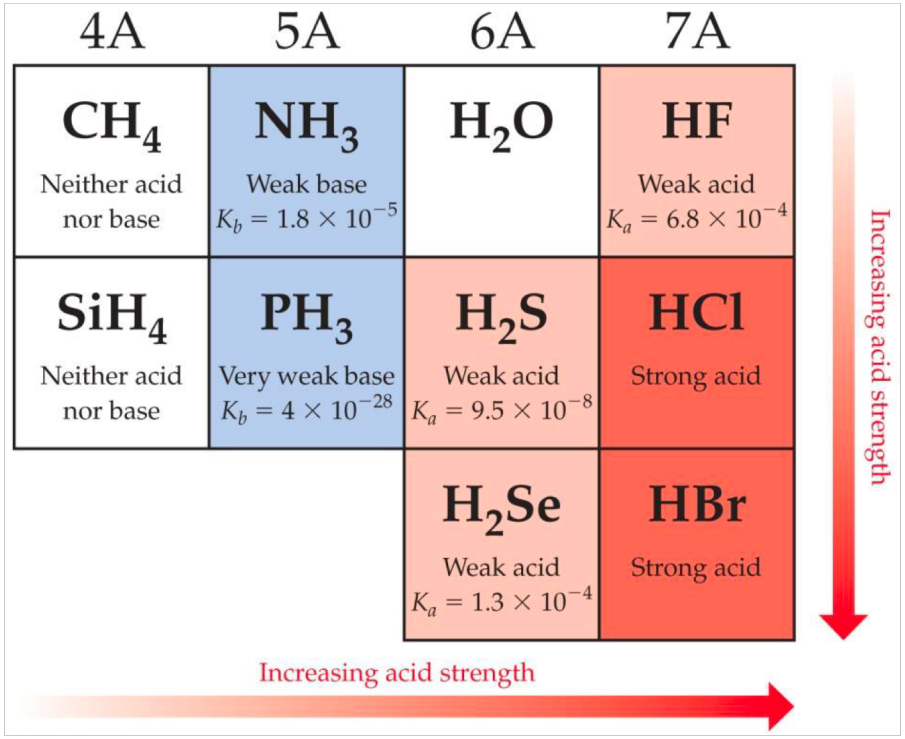
- Less electronegative, share e^- more equally, stronger bond
- Ex: P less electronegative than S, H_3PO_4 weaker than H_2SO_4

For Oxoacids

- If all else is equal, fewer double bonds = weaker acid
- Ex: H_2SO_3 weaker than H_2SO_4
- More double bonds = more resonance = more stability with H^+ removed

Stability of conjugate base: more stable anion = more acidic

Binary Acid Trends in the Periodic Table



For binary acids:

Acidity increases left to right across a row

- Electrongativity increases left to right

Acidity increases top to bottom down a group

- Atomic size increases down a group

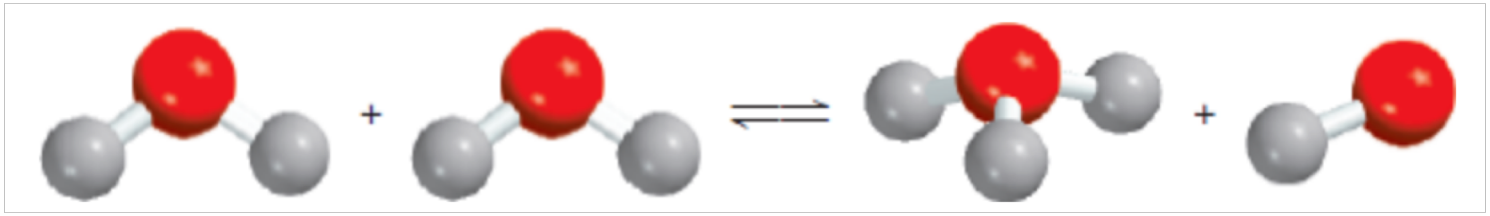
Acid-Base Properties of Salts

Salts are ionic compounds – contain cation (+) & anion (-)

- Can be formed from acid/base neutralization
- Dissociate (separate into ions) when dissolve in water
- Can sometimes impact pH through hydrolysis

Hydrolysis:

- Hydro = water
- lysis = to cut
- Reactions that break water into H^+ (H_3O^+) & OH^-



Extent to which a salt causes hydrolysis determines extent of its acid-base properties (if any)

Acid-Base Properties of Salts

Hydrolysis can be caused by both cations & anions, but <u>does not happen in every situation</u>

Salts formed from conjugates of a strong acid + a strong base do not cause hydrolysis

For salt solutions: no hydrolysis = neutral (no OH^- or H_3O^+)

$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$ – no hydrolysis

NaOH
strong base

HCl
strong acid

$LiNO_3(aq) \longrightarrow Li^+(aq) + NO_3^-(aq)$ – no hydrolysis

LiOH
strong base

HNO_3
strong acid

Acid-Base Properties of Salts

Soluble salts derived from a strong acid + weak base or a strong base + weak acid will cause hydrolysis

Ex 1: NaOH: strong base

CH_3COOH: weak acid

Weak base

$$CH_3COONa(s) \xrightarrow{H_2O} Na^+(aq) + CH_3COO^-(aq)$$

$$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

Ex 2: NH_3: weak base

HCl: strong acid

Weak acid

$$NH_4Cl(s) \xrightarrow{H_2O} NH_4^+(aq) + Cl^-(aq)$$

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

Determining if a Salt Solution is Acidic, Basic, or Neutral

Look at the ions that the salt dissociates into:

- If cation would come from a strong base – not acidic
- If anion would come from a strong acid – not basic
- If both cation & anion have a strong parent – neutral
- If cation would come from a weak base – acidic
- If anion would come from a weak acid – basic
- If both cation & anion have a weak parent
  - Acidity/basicity depends on relative strength
  - Math can be complicated
  - Qualitatively:
    - $K_a < K_b$ basic solution
    - $K_a > K_b$ acidic solution
    - $K_a \approx K_b$ pretty close to neutral

Determining if a Salt Solution is Acidic, Basic, or Neutral

1.) Predict whether aqueous solutions of the following compounds are acidic, basic, or neutral.

a.) NH_4NO_3

b.) $FeCl_3$

c.) Na_2CO_3

d.) $KClO_4$

e.) ZnF_2

2.) For each of the following pairs, indicate which salt would form the more acidic (i.e. less basic) 0.010M solution.

a.) $NaNO_3$ or $Fe(NO_3)_3$

b.) CH_3NH_3Cl or $BaCl_2$

c.) KNO_2 or KNO_3

d.) $(NH_4)_2SO_4$ or $(NH_4)_2SO_3$

3.) Calculate the $[OH^-]$ and pH of a 0.10M NaCN solution. K_a for HCN is 4.9×10^{-10}.

A: $[OH^-] = 1.4 \times 10^{-3}$M
pH = 11.16

4.) Calculate the pH of a 0.42M NH_4Cl solution. K_b for NH_3 is 1.8×10^{-5}

A: pH = 4.82

What if an ion can act as either an acid or a base?

Occurs with salts from polyprotic acids such as H_2CO_3, H_3PO_4, H_2SO_4.

Need to compare K_a and K_b values

- Example: $NaHCO_3$

Acting as acid

- $HCO_3^-(aq) + 2H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$

$K_a = 4.8 \times 10^{-11}$

Acting as base

- $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

$K_b = K_w/K_a = 2.4 \times 10^{-8}$

- $K_b > K_a$ so solution will be basic!

Lewis Acids & Bases

Acid:

An electron pair acceptor

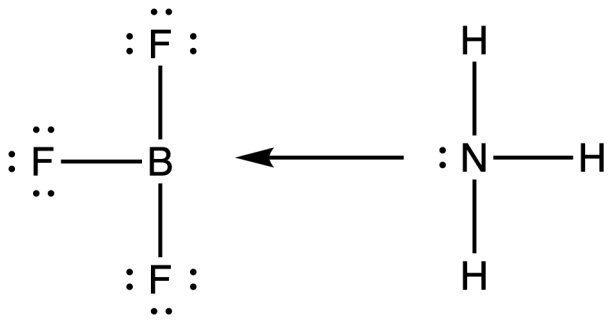
- Do not need to have a removable proton
- Must have an empty orbital to put electrons in

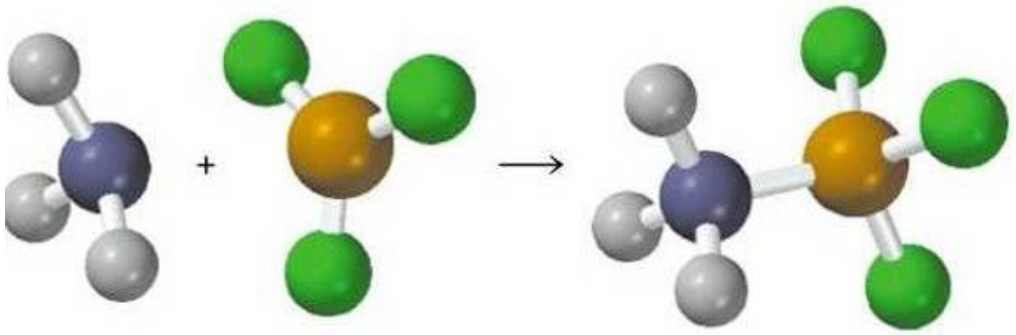
Base:

An electron pair donor

- Must have a pair of non-bonding electrons

Ex: $BF_3(g) + NH_3(g) \rightarrow F_3B\text{-}NH_3(g)$





Hydration of Metal Ions

Salts with small, highly charged metal cations (e.g. Al^{3+}, Fe^{3+}, Cr^{3+}, Be^{2+}) and the conjugate base of a strong acid can also have acidic properties

Example:
$AlCl_3$ dissolved in water produces:

$$Al(H_2O)_6^{3+}(aq) + H_2O(l) \rightleftharpoons Al(OH)(H_2O)_5^{2+}(aq) + H_3O^+(aq) + 3Cl^-(aq)$$

$K_a = 1.3 \times 10^{-5}$ → Acid!

- e^- in water molecules are pulled toward metal ion
- O-H bonds in attached waters become more polarized
- H^+ dissociates from water molecule forming H_3O^+

Hydration of Metal Ions

Higher charge & smaller size make cations more acidic

$$Na^+ < Ca^{2+} < Zn^{2+} < Al^{3+}$$

Hydration of Metal Ions

Which member of each pair would produce a more acidic solution & why?

1.) CuCl or $Cu(NO_3)_2$

2.) $CrCl_3$ or $NiCl_3$