Chapter 14: Acids, Bases and Salts
Sections 14.1 – 14.5

CHM152 GCC

OpenSTAX: Chemistry

Why study acids & bases?

- Many household substances, including cleaning solutions and food/beverages that we consume, are acids or bases.
- In the environment, the pH of rain, water and soil can also have significant effects.
- Acid – base reactions occurring in our body are essential for life. They are also involved in many industrial processes.

Acid vs. Base Characteristics

<table>
<thead>
<tr>
<th>Characteristics</th>
<th>Acids</th>
<th>Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Taste</td>
<td>Sour</td>
<td>Bitter, chalky</td>
</tr>
<tr>
<td>Corrode Metals</td>
<td>Most</td>
<td>Soap, slippery</td>
</tr>
<tr>
<td>Turn Litmus</td>
<td>Paper red</td>
<td>Turn litmus paper blue</td>
</tr>
<tr>
<td>Have Low pHs</td>
<td></td>
<td>Have high pHs</td>
</tr>
<tr>
<td>E.g.</td>
<td>Vinegar, lemon juice, gastric juice, soft drinks</td>
<td>E.g. milk of magnesia, bleach, ammonia, drano</td>
</tr>
</tbody>
</table>

Acid-Base Models

CHM 150/151 taught Arrhenius theory
- Acids produce hydrogen ions in water
  - HA (aq) → H⁺ (aq) + A⁻ (aq)
- Bases produce hydroxide ions in water
  - MOH (aq) → M⁺ (aq) + OH⁻ (aq)

CHM 152 uses Brønsted-Lowry theory
- Bronsted Acid: Substance that can donate H⁺
  - Acid loses H⁺
- Bronsted Base: Substance that can accept H⁺
  - Base gains H⁺

Bronsted-Lowry Reactions involve the transfer of 1 H⁺

A conjugate base is formed when a Brønsted acid loses a proton.

\[
\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]

Conjugate pair: differ by 1 H⁺ ion

Conjugate pairs above: 1) HCl and Cl⁻ 2) H₂O and H₂O⁺

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  - Principles of General Chemistry (CC BY-NC-SA 3.0): http://2012books.lardbucket.org/pdfs/principles-of-general-chemistry-v1.0.pdf

Taste sour
Taste bitter, chalky
Corrode most metals
Feel soapy, slippery
Turn litmus paper red
Turn litmus paper blue
Have low pH's
Have high pH's
E.g. vinegar, lemon juice, gastric juice, soft drinks
E.g. milk of magnesia, bleach, ammonia, drano
A conjugate acid is formed when a Brønsted base accepts a proton. 

\[ \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq) \]

Note: Water is amphoteric – it can act as an acid or a base depending on what it's reacting with.

**Conjugate Pairs**

<table>
<thead>
<tr>
<th>Conjugate</th>
<th>Acceptor</th>
<th>Donor</th>
</tr>
</thead>
<tbody>
<tr>
<td>Acid</td>
<td>H$_3$O$^+$ (aq)</td>
<td>NH$_4^+$ (aq)</td>
</tr>
<tr>
<td>Base</td>
<td>NH$_3$ (aq)</td>
<td>OH$^-$ (aq)</td>
</tr>
</tbody>
</table>

**Identify Conjugate Acid/Base**

a) What is the conjugate acid of CO$_3^{2-}$?

b) What is the conjugate base of H$_2$PO$_4^-$?

**Bronsted-Lowry Reaction**

Write an equation for the dissociation of HCN in water. Identify the acid, the base, the conjugate acid, and the conjugate base.

Watch charges! Acid loses 1 H$^+$, base gains 1 H$^+$

**Hydronium Ions**

- HA (aq) ⇌ H$^+$ (aq) + A$^-$ (aq)
- H$^+$ is very reactive and will bond with O in H$_2$O to form H$_3$O$^+$
- H$^+$ and H$_3$O$^+$ are used interchangeably but H$_3$O$^+$ is a more accurate representation.

**Autoionization of Water**

In water a H$^+$ ion can be transferred between H$_2$O molecules.

\[ \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{O} + \text{H}^+ + \text{OH}^- \]

This equilibrium process is highly reactant favored thus it gives rise to a very small equilibrium constant.
Equilibrium Constant for Water Ionization, $K_w$

$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

$K_w = [H_3O^+][OH^-]$

Why is water omitted?

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

or $K_w = [H^+][OH^-]$

$K_w = 1 \times 10^{-14}$ at 25 ºC

Relationship between acidity, $[H^+]$, $[OH^-]$

$K_w = [H^+][OH^-] = 1 \times 10^{-14}$

Neutral: $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ M

Acidic: $[H^+] > [OH^-]$

Basic: $[H^+] < [OH^-]$

Example. $[H_3O^+] = 5.6 \times 10^{-4}$ M. What is $[OH^-]$?

Is this solution acidic or basic?

$pH$: A measure of Acidity

- $pH$ is the $-\log$ of the $H^+$ concentration (M)
- Scale was proposed to give us convenient numbers to work with

Relationship Between $pH$ and Acidity

$pH < 7$, $[H^+] > 1.0 \times 10^{-7}$ M Acidic

$pH = 7$, $[H^+] = 1.0 \times 10^{-7}$ M Neutral

$pH > 7$, $[H^+] < 1.0 \times 10^{-7}$ M Basic

$pH$ Scale

Memorize these equations!

$pH = -\log [H_3O^+] \quad [H_3O^+] = 10^{-pH}$

$pOH = -\log [OH^-] \quad [OH^-] = 10^{-pOH}$

$K_w = [H_3O^+][OH^-] = 1 \times 10^{-14}$

$pH + pOH = 14$

$pH$ Sig Figs

2 sig figs

$[H_3O^+] = 5.6 \times 10^{-4}$ M 2 decimal places

$pH = -\log (5.6 \times 10^{-4}) = 3.25$

3 related to exponent; tells us acidity

Find $[H_3O^+]$, $pH$, & $pOH$ if $[OH^-] = 9.8 \times 10^{-9}$ M 2 sf

$pOH = -\log [OH^-] = -\log 9.8 \times 10^{-9} = 8.01$ (2 dec pl)

$pH = 14 - pOH = 14 - 8.01 = 5.99$ (2 dec pl)

$[H_3O^+] = 10^{-pH} = 10^{-5.99} = 1.0 \times 10^{-6}$ M (2 sf)
Complete the following table

<table>
<thead>
<tr>
<th>pH</th>
<th>$[H^+] M$</th>
<th>$[OH^-] M$</th>
<th>pOH</th>
<th>A, B, N?</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0x10^{-9}</td>
<td>4.815</td>
<td>2.30</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

### Strong Acids

A **strong acid** ionizes completely in water to form $H_2O^+$ ions ($H^+$). SA's are strong electrolytes.

$HNO_3(aq) + H_2O(l) \rightarrow H_3O^+(aq) + NO_3^-(aq)$

- A one way arrow is used since this reaction is complete, all of the $HNO_3$ molecules break apart to form $H_3O^+$ and $NO_3^-$ ions.
- These acids have very weak conjugate bases
- For SA, eq lies very far to the right side!

**KNOW 7 Strong acids! Memorize these!**

HCl, HBr, HI, HNO$_3$, H$_2$SO$_4$, HClO$_4$, HClO$_3$

### Strong Bases

A **strong base** ionizes completely in water to form $OH^-$ ions. SB's are strong electrolytes!

SB dissociates in water; it doesn’t react with it:

$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$ (100 % ions)

- This reaction also goes to completion

$\Rightarrow$ SB’s completely dissociate into ions

**Strong Bases: These Group 1A & 2A hydroxides**

LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)$_2$,

Ba(OH)$_2$, Sr(OH)$_2$ **Memorize the 8 SB’s!**

### Weak Acids

- A **weak acid** ionizes only to a small extent in $H_2O$.
- WA’s are weak electrolytes!

E.g. $HCN(aq) + H_2O(l) \Leftrightarrow H_3O^+(aq) + CN^-(aq)$

- A two-way arrow is used since this reaction is not complete but instead consists of an equilibrium mixture of $HCN$ molecules, $H_3O^+$ and $CN^-$ ions.
- Equilibrium lies to the left since most of the acid molecules have not ionized.
- Common WA’s: HF, HNO$_2$, HCN, H$_3$PO$_4$, H$_2$CO$_3$,

- carboxylic acids like CH$_3$COOH and HCOOH

### Weak Base

A **weak base** ionizes only to a small extent in $H_2O$. WB’s are weak electrolytes!

$NH_3(aq) + H_2O(l) \Leftrightarrow NH_4^+(aq) + OH^-(aq)$

- Equilibrium lies to the left since most of the ammonia molecules have not accepted protons to form $NH_4^+$ and $OH^-$ ions.
- Common examples of WB’s are ammonia and amines (e.g. CH$_3$NH$_2$)

### SA, WA, SB, WB solutions

Draw the following in water: HNO$_3$, HF, Ca(OH)$_2$, and NH$_3$. Assume you have 3 formula units or molecules for each substance.
**Conjugate Acid-Base Pair Trends**

1. A stronger acid loses its proton more readily than a weaker acid and a stronger base gains a proton more readily than a weaker base.
2. The stronger the acid, the weaker its conjugate base. Likewise, the stronger the base, the weaker its conjugate acid.
3. Proton transfer reactions proceed from the stronger acid to the stronger base. Thus, the equilibria lies on the side of the weaker acid-base pair.

**Strengths of Acids/Bases**

e.g. HS\(^{-}\)(aq) + HF(aq) ⇌ H\(_2\)S(aq) + F\(^{-}\)(aq)

- From Table on slide
- Which acid is stronger?
- Which base is stronger?
- Which side does equilibria lie on?

**pH of Strong Acids/Bases**

Calculate pH of 0.103 M HNO\(_3\) solution.
HNO\(_3\)(aq) + H\(_2\)O(l) → H\(_3\)O\(^{+}\)(aq) + NO\(_3\)\(^{-}\)(aq)

Calculate pH of 0.020 M Sr(OH)\(_2\) soln.
Sr(OH)\(_2\)(aq) → Sr\(^{2+}\)(aq) + 2OH\(^{-}\)(aq)

**Weak Acids; Acid-Dissociation Constants (K\(_a\))**

The ionization of a weak acid, HA, in water:

\[ \text{HA}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{A}^{-}(aq) \]

\[ K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^{-}]}{[\text{HA}]} \]

\( K_a \) is called the acid dissociation constant.
The larger the value of \( K_a \) the stronger the acid.

**Weak Acids; Acid-Ionization Constants (K\(_a\))**

<table>
<thead>
<tr>
<th>Acids</th>
<th>( K_a ) values at 25 (^\circ) C</th>
</tr>
</thead>
<tbody>
<tr>
<td>HF</td>
<td>3.5 x 10(^{-4})</td>
</tr>
<tr>
<td>HNO(_2)</td>
<td>4.6 x 10(^{-4})</td>
</tr>
<tr>
<td>HNCO</td>
<td>2.0 x 10(^{-4})</td>
</tr>
<tr>
<td>HCO(_2)H</td>
<td>1.8 x 10(^{-4})</td>
</tr>
<tr>
<td>CH(_3)CO(_2)H</td>
<td>6.3 x 10(^{-5})</td>
</tr>
<tr>
<td>HClO</td>
<td>3.5 x 10(^{-8})</td>
</tr>
<tr>
<td>HBrO</td>
<td>2.8 x 10(^{-3})</td>
</tr>
<tr>
<td>HCN</td>
<td>4.9 x 10(^{-10})</td>
</tr>
<tr>
<td>C(_6)H(_5)COOH</td>
<td>6.3 x 10(^{-5})</td>
</tr>
</tbody>
</table>

NOTE: The reported constants vary – make sure to use the value provided in the question.

Additional \( K_a \) values found in appendix in the OpenSTAX book.
Equilibria of Weak Acids

HA(aq) + H₂O(l) ⇌ H₃O⁺(aq) + A⁻(aq)

• \( pK_a = -\log K_a \): stronger acid = lower \( pK_a \)

\[
\text{percent ionization} = \frac{[\text{H}_3\text{O}^+]}{[\text{HA}]} \times 100\%
\]

• For a stronger acid, indicate if the following will be higher or lower:
  - \([\text{H}_3\text{O}^+]\)
  - pH
  - % ionization
  - \( K_a \)
  - \( pK_a \)

\[\text{pH} = -\log [\text{H}_3\text{O}^+]\]

Weak Acid Problems

• Set up ICE tables & \( K_a \)
• With small \( K \) values, assume \( x \) will be small to simplify math.

WA problem types:
• Given \([\text{HA}]\) and \( K_a \), find all equilibrium concentrations and pH
• Given \([\text{HA}]\) and pH, find all equilibrium concentrations and \( K_a \)
• Given \([\text{HA}]\) and % dissociation, find eq concentrations, pH and \( K_a \)

Find pH of Weak Acid

Find the equilibrium concentrations of all substances and the pH of a 0.50 M solution of HClO (hypochlorous acid). \( K_a = 3.5 \times 10^{-8} \)

• Set up ICE Table

<table>
<thead>
<tr>
<th></th>
<th>HCIO + H₂O ⇌ H₃O⁺ + ClO⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
</tr>
<tr>
<td>E</td>
<td></td>
</tr>
</tbody>
</table>

\[x = [\text{H}_3\text{O}^+] = [\text{ClO}^-] = [\text{HClO}] = 0.50 - x = \]

\[\text{pH} = -\log [\text{H}_3\text{O}^+] = \]

Find \( K_a \) for weak acid

Find the \( K_a \) of a 1.25 M solution of nitrous acid, HNO₂. The pH of this solution is measured to be 1.62.

Write hydrolysis reaction & \( K_a \) expression!

\[\text{pH} = 10^{-\text{pH}} = \]

(keep 2 sf since pH has 2 dec pl)

\[\text{NO}_2^- = [\text{HNO}_2] = \]

*Plug in x (H₃O⁺) since we know value from pH!\n
\[K_a = \]

Percent Ionization

% dissociation is the same calculation as checking assumption of a small \( x \).

• % dissociation = \(((\text{H}_3\text{O}^+)_\text{eq} / [\text{HA}]) \times 100\%

• Calculate the percent dissociation from the previous problem.

\[\% \text{dissociation} = \]

Will the percent dissociated increase or decrease as an acid is diluted?
**Effect of Dilution on % Ionization**

\[ \text{HF(aq)} + \text{H}_2\text{O(l)} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq) \]

As we add water, we can dissociate more. More dilute = more dissociation.

---

**Find K_a for weak acid**

A 0.200 M solution of a weak acid, HA, is 9.4 percent dissociated. Using this information, calculate [H_3O^+], [A^-], [HA]_{aq}, and K_a.

Answer: [H_3O^+] = 0.019 M, [A^-] = 0.019 M, [HA] = 0.181 M, and K_a = 2.0 \times 10^{-3}

See Worksheet key

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**Polyprotic Acids**

\[ \text{H}_2\text{C}_2\text{O}_4(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_2\text{O}_4^{2-}(aq) + \text{H}_3\text{O}^+(aq) \]

* K_{a1} = [\text{C}_2\text{O}_4^{2-}][\text{H}_3\text{O}^+] / [\text{H}_2\text{C}_2\text{O}_4] = 5.9 \times 10^{-2}

\[ \text{C}_2\text{O}_4^{2-}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_2\text{O}_4^{3-}(aq) + \text{H}_3\text{O}^+(aq) \]

* K_{a2} = [\text{C}_2\text{O}_4^{3-}][\text{H}_3\text{O}^+] / [\text{C}_2\text{O}_4^{2-}] = 6.4 \times 10^{-5}

In general, K_{a1} > K_{a2} > K_{a3} ... Why?

* pH of a polyprotic acid solution primarily arises from [H^+] formed in the 1st step.

---

**Weak Bases; Base-Ionization Constants (K_b)**

The ionization of a weak base in water:

\[ \text{B(aq)} + \text{H}_2\text{O}(l) \rightleftharpoons \text{HB}^-(aq) + \text{OH}^-(aq) \]

\[ K_b = \frac{[\text{HB}^-][\text{OH}^-]}{[\text{B}]} \]

K_b is called the **base dissociation constant**.

The larger the value of K_b, the stronger the base.

---

**Weak Bases; K_b Constants**

<table>
<thead>
<tr>
<th>Base</th>
<th>Formula</th>
<th>K_b</th>
<th>pK_b</th>
</tr>
</thead>
<tbody>
<tr>
<td>phosphate ion</td>
<td>PO_4^{3-}</td>
<td>2.1 \times 10^{-2}</td>
<td>1.68</td>
</tr>
<tr>
<td>dimethylamine</td>
<td>(CH_3)NH</td>
<td>5.4 \times 10^{-4}</td>
<td>3.27</td>
</tr>
<tr>
<td>methylamine</td>
<td>CH_3NH_2</td>
<td>4.6 \times 10^{-4}</td>
<td>3.34</td>
</tr>
<tr>
<td>trimethylamine</td>
<td>(CH_3)N</td>
<td>6.3 \times 10^{-5}</td>
<td>4.20</td>
</tr>
<tr>
<td>ammonia</td>
<td>NH_3</td>
<td>1.8 \times 10^{-5}</td>
<td>4.75</td>
</tr>
<tr>
<td>pyridine</td>
<td>C_5H_5N</td>
<td>1.7 \times 10^{-9}</td>
<td>8.77</td>
</tr>
<tr>
<td>aniline</td>
<td>C_6H_5NH_2</td>
<td>7.4 \times 10^{-10}</td>
<td>9.13</td>
</tr>
</tbody>
</table>

Similar to WA problems but base hydrolysis reaction forms OH^- (this is x!) and use K_b

* NH_3(aq) + H_2O(l) ⇌ NH_4^+(aq) + OH^-(aq)

    \[ K_b = [\text{NH}_4^+][\text{OH}^-] / [\text{NH}_3] \]

* K_b Values for Weak Bases in appendix

* Why are amines (derivatives of NH_3) basic?

⇒ Lone e- pair on N attracts H from water. N pulls H^+ off of water.
Weak Bases

Calculate the pH of a 0.50 M dimethylamine \((\text{CH}_3)_2\text{NH}\) solution. \(K_b = 5.4 \times 10^{-4}\)

Write reaction, set up ICE, and \(K_b\).

\[x = \text{Check approximation!}\]

\[\text{Plug in } x!\]

\[\begin{align*}
[(\text{CH}_3)_2\text{NH}_2^+] &= [\text{OH}^-] = \\
[(\text{CH}_3)_2\text{NH}] &= 
\end{align*}\]

Weak Bases

2 ways to find pH

1) Convert \([\text{OH}^-]\) to \([\text{H}_3\text{O}^+]\):

\[x = [\text{OH}^-] = \\
[H_3\text{O}^+] = K_w / [\text{OH}^-] = \]

\[\text{pH} = -\log [\text{H}_3\text{O}^+] = \]

2) Convert \([\text{OH}^-]\) to pOH

\[\text{pOH} = -\log [\text{OH}^-] = \]

\[\text{pH} = 14 - \text{pOH} = \text{this is easiest way}\]

Relationship between \(K_a\) and \(K_b\)

The ionization constant of a weak acid (\(K_a\)) and the ionization constant of a weak base (\(K_b\)) are related:

\[\begin{align*}
\text{CH}_3\text{COOH}(aq) &\rightleftharpoons \text{H}^+(aq) + \text{CH}_3\text{COO}^-(aq) \\
\text{CH}_3\text{COO}^-(aq) + \text{H}_2\text{O}(l) &\rightleftharpoons \text{CH}_3\text{COOH}(aq) + \text{OH}^-(aq) \\
\text{H}_2\text{O}(l) &\rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq)
\end{align*}\]

\[K_a = \frac{[\text{H}^+] [\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} \quad K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} \quad K_a \cdot K_b = K_w\]

Calculate \(K_b\)

What is the \(K_b\) for acetate if \(K_a\) is \(1.8 \times 10^{-5}\) for acetic acid?

For conjugate pairs: \(K_a \cdot K_b = 10^{-14}\)

Properties of Salts

- Acid + Base → Salt + Water
  E.g. \(\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)\)
- Salts are ionic products of acid-base reaction.
- Cation in salt: comes from the base
- Anion in salt: comes from the acid
- Salts can be neutral, acidic, or basic depending on the strength of the acid or base from which they’re made.

General Ion Categories

1) Acidic ions: Most cations (except 1A & 2A)
  E.g. \(\text{NH}_4^+, \text{Al}^{3+}, \text{Cu}^{2+}\)

2) Neutral ions: cations from strong bases and anions from strong acids
  Cations: Group 1A & 2A (except Be)
  Anions: \(\text{NO}_3^-, \text{Cl}^-, \text{Br}^-, \text{I}^-, \text{ClO}_4^-, \text{ClO}_3^-\) (except \(\text{HSO}_4^-\))

3) Basic ions: Most anions derived from weak acids.
  E.g. \(\text{F}^-, \text{CN}^-, \text{CH}_3\text{COO}^-, \text{NO}_2^-\)
**Some Metal Ions form Acidic Salts**

Small, highly charged metal ions can react with water to produce acidic solution. Easier to break O-H bonds and form hydronium’s.

Types of Salts

**Neutral Salts** – salt formed in SA-SB rxn
- Cation from SB and anion from SA are neutral ions
- Neutral ions don’t hydrolyze!
- E.g. NaCl
  \[ \text{Na}^+ + \text{H}_2\text{O} \rightarrow \text{NR} \; ; \; \text{Cl}^- + \text{H}_2\text{O} \rightarrow \text{NR} \]

**Acidic Salts** – typically contain acidic cation and neutral anion (from SA)
- Cation is
  1) conjugate acid of WB (e.g. NH$_4^+$)
  2) small, highly charged metal ion (e.g. Al$^{3+}$)
- E.g. NH$_4$Cl
  \[ \text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_3(aq) + \text{H}_3\text{O}^+(aq) \]

**Basic Salts** – typically contain basic anion (from weak acid) and neutral cation (from SB)
- Anion is conjugate base of WA (e.g. F$^-$)
- E.g. NaF
  \[ \text{F}^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HF}(aq) + \text{OH}^-(aq) \]

Salts containing Acidic cations and basic anions
- Compare K$_a$ and K$_b$ values to tell if its acidic or basic. (If K$_a$ is larger, salt is acidic and visa versa.)
- E.g. NH$_4$NO$_2$

Example Salt Problems

- Is KF an acidic, basic or neutral salt?
  - Write the hydrolysis reaction and calculate the pH of a 0.10 M KF solution. K$_a$ (HF) = 3.5 x 10$^{-4}$

- Is NH$_4$Cl acidic, basic, or neutral?
  - Write the hydrolysis reaction and calculate the pH of a 0.10 M NH$_4$Cl solution.
  K$_b$ (NH$_3$) = 1.8 x 10$^{-5}$

See Worksheet key

Factors that Affect Acid Strength

**Acid strength** is a measure of how completely it ionizes.

\[ \text{HX} \rightarrow \text{H}^+ + \text{X}^- \]

1) Acid strength ↑ as strength of H-X bond ↓
- easier to break weak H-X bonds

2) Acid strength ↑ as polarity of H-X bond ↑
- Polarity↑ when X is more electronegative (easier to ionize)
Why is HF a weak acid?
• Determined by the strength of H–X bond.
• H–F has strongest bonds (F is small so it’s closer to H)
• Acid Strength ↑ as Bond Strength ↓

### Acid Strength & Bond Polarity
As we move to the right across a row, the acid strength is related to the polarity of the H–X bond.
C has the lowest electronegativity, so C–H is least polar and least acidic. F is most electronegative, so F–H is most polar, and most acidic (easiest to ionize).

![Acid Strength & Bond Polarity](image)

### Oxoacids – Common Examples
Oxoacids: An oxoacid contains hydrogen, oxygen, and a central nonmetal atom (acids made from polyatomic ions).

#### Oxoacid Strength in Same Group
1) Oxoacids having different central atoms that are from the same group.

- Acid strength ↑ as electronegativity ↑
- Cl is more electronegative so O–H bond is more polar.
- Thus HOCl > HOBr > HOI

<table>
<thead>
<tr>
<th>HOX</th>
<th>Electronegativity of X</th>
<th>$K_a$</th>
<th>$pK_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>HOCl</td>
<td>3.0</td>
<td>4.0 x 10$^8$</td>
<td>7.40</td>
</tr>
<tr>
<td>HOBr</td>
<td>2.8</td>
<td>2.8 x 10$^9$</td>
<td>8.55</td>
</tr>
<tr>
<td>HOI</td>
<td>2.5</td>
<td>3.2 x 10$^{11}$</td>
<td>10.5</td>
</tr>
</tbody>
</table>

### Lewis Acids and Bases
- Broadest acid base definition since Lewis acids don’t have to contain H$^+$.
- Looks at electron transfer instead of H$^+$ transfer

A **Lewis base** is an electron pair donor.

A **Lewis acid** is an electron pair acceptor.

E.g. $H^+ + :NH_3 \rightarrow NH_4^+$

⇒ Ammonia donates the electron pair to H$^+$ to make the bond; H$^+$ accepts the electron pair

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**Strength Hydrohalic acids:**

HF << HCl < HBr < HI

- Why is HF a weak acid?
  - Determined by the strength of H–X bond.
  - H–F has strongest bonds (F is small so it’s closer to H)
  - Acid Strength ↑ as Bond Strength ↓

<table>
<thead>
<tr>
<th>Relative Acid Strength</th>
<th>HF</th>
<th>&lt;</th>
<th>HCl</th>
<th>&lt;</th>
<th>HBr</th>
<th>&lt;</th>
<th>HI</th>
</tr>
</thead>
<tbody>
<tr>
<td>H–X Bond Energy (kJ/mol)</td>
<td>570</td>
<td>432</td>
<td>366</td>
<td>298</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$K_a$</td>
<td>6.3 x 10$^4$</td>
<td>1.3 x 10$^9$</td>
<td>7.9 x 10$^8$</td>
<td>2.0 x 10$^9$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$pK_a$</td>
<td>3.20</td>
<td>-6.1</td>
<td>-8.9</td>
<td>-9.3</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>