Acid and Base Chemistry

What comes to mind when you think acid?
General Characteristics of All Acids

• Result in the formation of Hydronium ($H_3O^+$)
• Most chemical formulas start with a hydrogen (H)
  • $HCl$, $H_2SO_4$, $H_2CO_3$
• Typically liquids or gases
• Taste sour (i.e. Lemons)
  o NEVER TASTE TO IDENTIFY ACIDS
• Can cause severe chemical burns
  • The acids wick water away
  • Exothermic reaction
  • Results in blisters, swelling, & infections
  o NEVER TOUCH TO IDENTIFY ACIDS

Don’t mess with acids, or else you’ll get burned
Common examples of Acids

- Carbonic Acid ($\text{H}_2\text{CO}_3$)
  - Pop
- Ascorbic Acid ($\text{HC}_6\text{H}_7\text{O}_6$)
  - Vitamin C (Lack of leads to Scurvy)
  - Many vitamins come in the form of acids
- Acetic Acid ($\text{HC}_2\text{H}_3\text{O}_2$)
  - Vinegar
- Citric Acid ($\text{HC}_6\text{H}_7\text{O}_7$)
  - Citrus Fruits, Mountain Dew
- Sulfuric Acid ($\text{H}_2\text{SO}_4$)
  - Battery Acid
What is an acid?

• **Definition** - Acids are substances, that when dissolved in water, increase the concentration of Hydronium (H$_3$O$^+$)
• An acid must have a proton (H$^+$) to donate!
• Strong acids ionize *completely*
  \[ \text{HNO}_3(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{NO}_3^-_{(\text{aq})} \]
• Weak acids ionize *incompletely*
  \[ \text{HF}_{(\text{aq})} + \text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{F}^-_{(\text{aq})} \]
• Hydronium...
  ✓ Is a product
  ✓ Is a polyatomic ion
  ✓  \( \text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+ \)
Why Hydronium?

Behind the reaction

\[ \text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+ \]

Provided by the acid

For example, the “\(\text{H}^+\)” is a product of the ionized (disassociated) acid.

\[ \text{HCl} \rightarrow \text{H}^+ + \text{Cl}^- \]

“\(\text{H}^+\)” is often called a proton. Why?

1 Proton, 1 Electron

0 Neutrons

SO...
This proton (H\(^+\)) then joins together with water. Why?

\[
\text{H}_2\text{O} + \text{H}^+ \rightarrow \text{H}_3\text{O}^+
\]

The negative side of water (oxygen side) attracts the positive proton to form hydronium.
HCl + H₂O $\rightarrow$ H₃O⁺ + Cl⁻

Step 1: Add acid
Step 2: Acid Disassociates
Step 3: Hydronium Forms
Acid Strength

- Acids can be strong or weak
- The greater the concentration of $\text{H}_3\text{O}^+$, the stronger the acid
- Strong acids
  - The acid is very soluble in water
  - Ionizes (breaks apart / disassociates) completely in aqueous solutions
  - The more it breaks apart, the stronger the concentration
- Weak acids
  - Not very soluble in water
  - Don’t ionize completely
  - Acetic Acid ($\text{HC}_2\text{H}_3\text{O}_2$)
    - Maybe only 1 in 10 break up
Acids conduct electricity

• Strong acids conduct electricity very well
  • Strong electrolyte
  • Electrolyte – A substance that, when dissolved in a solvent, increases the solvent’s conductivity.

• Weak acids are weak electrolytes
Examples of Strong Acids

These acids disassociate completely

**The Big Three**

1. Hydrochloric Acid (HCl)
   - Stomachs, Household Cleaner
2. Nitric Acid (HNO₃)
   - Explosives, Nitroglycerin, Car Bombs, Oklahoma City Bombing
3. Sulfuric Acid (H₂SO₄)
   - Battery Acid
Examples of Weak Acids

These acids do **NOT** ionize completely

1. Acetic Acid (HC$_2$H$_3$O$_2$; CH$_3$COOH)
   - Vinegar
2. Hydrofluoric Acid (HF)
3. Hydrosulfic Acid (H$_2$S)
Concept Check

1. What is hydronium?
   • $\text{H}_3\text{O}^+$
   • It’s what’s formed when an acid is added to water.

2. How do we determine acid strength?
   • Electricity, pH

3. What are characteristics of acids?
   • Liquid / Gas
   • Tastes sour
   • Causes burns
   • Molecular compound starts with “H”
• Acids are ionic substances
• Acid names are based on ionic naming rules
Naming Ionic Compounds

Ionic Compound = Metal + Non-metal

Binary – Made of two parts

1. First Part
   ✓ Always the metal (the Cation)
   ✓ Keeps its atomic name

2. Second Part
   ✓ Always the non-metal (the Anion)
   ✓ “-ide” replaces the ending of the atomic name.

Note: If a polyatomic ion is involved, the polyatomic ion keeps its name
When *naming* ionic compounds, subscripts are not included in the name due to electroneutrality.

Aluminum Sulfide

\[
\text{Al}^{+3} + \text{S}^{-2} \rightarrow \text{Al}_2\text{S}_3
\]
**Modifications for acids?**

Rules: To name an acid from a chemical formula

1. Identify the ionic name.
2. Eliminate the cation name.
3. Modify the anion based on the following rules.

<table>
<thead>
<tr>
<th>Anion Endings</th>
<th>Acid Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>-ide</td>
<td>Hydro- (stem)-ic acid</td>
</tr>
<tr>
<td>-ate</td>
<td>(stem)-ic acid</td>
</tr>
<tr>
<td>-ite</td>
<td>(stem)-ous acid</td>
</tr>
</tbody>
</table>

**Example: HCl**

- Ionic Name: **Hydrogen Chloride**
- Acid Name: **Hydro-chlor-ic Acid**
- Hydrochloric Acid
Example #2

Example: $\text{H}_2\text{SO}_4$

- Ionic Name: **Hydrogen Sulfate**
- Acid Name: **Sulfur-ic Acid (Sulfic Acid)**
- Sulfuric Acid

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</tr>
<tr>
<td>-ite</td>
<td>(stem)-ous acid</td>
</tr>
</tbody>
</table>
Now, figure out the chemical formula from the name

Example: Hydrobromic Acid

1. Identify the stem
   • → “brom”
2. Hydro-(stem)-ic = “ide”
   • “Bromide”
3. Bromide = Bromine (Br)
   • Ion = Br^{-}
4. All acids have Hydrogen
   • Ion = H^{+}
5. Electroneutrality...
   • H^{+} + Br^{-} →
6. Hydrobromic Acid = HBr

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<td>-ite</td>
<td>(stem)-ous acid</td>
</tr>
</tbody>
</table>
Example #2

Example: Carbonic Acid

1. Identify the stem
   • \( \rightarrow \) “Carbon”

2. “ic” = “ate”
   • “Carbonate”

3. Carbonate = Polyatomic \( (CO_3^{-2}) \)
   • Polyatomic = \( CO_3^{-2} \)

4. All acids have Hydrogen
   • Ion = \( H^+ \)

5. Electroneutrality...
   • \( H^+ + CO_3^{-2} \rightarrow \)

6. Carbonic Acid = \( H_2CO_3 \)
Do these in your notes

<table>
<thead>
<tr>
<th>Chemical Formula</th>
<th>Ionic Name</th>
<th>Acid Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. $H_3PO_4$</td>
<td>Hydrogen Phosphate</td>
<td>Phosphoric Acid</td>
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<tr>
<td>2. $HClO_3$</td>
<td>Hydrogen Chlorate</td>
<td>Chloric Acid</td>
</tr>
<tr>
<td>3. $H_2CO_3$</td>
<td>Hydrogen Carbonate</td>
<td>Carbonic Acid</td>
</tr>
<tr>
<td>4. HCN</td>
<td>Hydrogen Cyanide</td>
<td>Hydrocyanic Acid</td>
</tr>
<tr>
<td>5. HF</td>
<td>Hydrogen Fluoride</td>
<td>Hydrofluoric Acid</td>
</tr>
<tr>
<td>6. HI</td>
<td>Hydrogen Iodide</td>
<td>Hydroiodic Acid</td>
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</table>
### Your turn

<table>
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<tr>
<th>Acid Name</th>
<th>Rule</th>
<th>Ions</th>
<th>Chemical Formula</th>
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<tbody>
<tr>
<td>1. Nitric acid</td>
<td>“ate”</td>
<td>H$^+$ + NO$_3^{-1}$</td>
<td>HNO$_3$</td>
</tr>
<tr>
<td>2. Sulfuric acid</td>
<td>“ate”</td>
<td>H$^+$ + SO$_4^{-2}$</td>
<td>H$_2$SO$_4$</td>
</tr>
<tr>
<td>3. Acetic acid</td>
<td>“ate”</td>
<td>H$^+$ + C$_2$H$_3$O$_2^{-1}$</td>
<td>HC$_2$H$_3$O$_2$</td>
</tr>
<tr>
<td>4. Hydroiodic acid</td>
<td>“ide”</td>
<td>H$^+$ + I$^{-1}$</td>
<td>HI</td>
</tr>
<tr>
<td>5. Hydrobromic acid</td>
<td>“ide”</td>
<td>H$^+$ + Br$^{-1}$</td>
<td>HBr</td>
</tr>
<tr>
<td>6. Hydrofluoric acid</td>
<td>“ide”</td>
<td>H$^+$ + F$^{-1}$</td>
<td>HF</td>
</tr>
</tbody>
</table>
Bases

• **Definition** – Bases are substances, that when dissolved in water, increase the concentration of hydroxide ions (OH⁻)

• \( \text{NaOH}_\text{(s)} + \text{H}_2\text{O}_\text{(l)} \rightarrow \text{Na}^+_{\text{(aq)}} + \text{OH}^-_{\text{(aq)}} \)

• \( \text{OH}^- = \text{Product} \)

• **Arrhenius Bases** – Need Water
General Characteristics of Bases

- Bases follow normal ionic naming rules
- Typically solids
- Taste bitter
- Feel slippery
- Usually end in “OH-”
- Bases are Dangerous!!!
**Alkalinity**

Means the base strength

- Is dependent on compound solubility (How well it breaks into ions)
- The greater the solubility, the stronger base
  - High concentration of OH$^-$
- Low solubility, weak base.
  - Low concentration of OH$^-$
Examples of Strong Bases

- Potassium Hydroxide - KOH
- Sodium Hydroxide - NaOH
- Calcium Hydroxide - Ca(OH)$_2$
- Sodium Phosphate - Na$_3$PO$_4$
**Weak Bases**

- Ammonia – \( \text{NH}_3 \)
- Potassium Carbonate – \( \text{K}_2\text{CO}_3 \)
- How do these produce \( \text{OH}^- \)?
  - They steal a proton from water
  \[
  \text{NH}_3(aq) + \text{H}_2\text{O}(l) \iff \text{NH}_4^+(aq) + \text{OH}^-(aq)
  \]
Naming Bases

• Bases follow normal ionic compound naming rules
  – NaOH (Sodium Hydroxide)
  – Mg(OH)$_2$ (Magnesium Hydroxide)
The Behavior of Water

- Water is *highly polar* and in *continuous motion*
  - Collisions occur
  - *Some* collisions are energetic enough to transfer H\(^+\) from one molecule to another.
  - \( \text{H}_2\text{O} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{OH}^- \) (*Self Ionization*)
    - In this situation, \([\text{H}_3\text{O}^+] = [\text{OH}^-]\)
    - \([x] = “\text{Concentration of } x”\)
  - “\(\text{H}_3\text{O}^+\)” = Acid, “\(\text{OH}^-\)” = Base
    - When these cancel one another, we have a...
  - “Neutral Solution”
    - \([\text{H}_3\text{O}^+] = [\text{OH}^-]\)
Self Ionization of Water

- In **pure** water, the \([H_3O^+]\) **always** equals the \([OH^-]\)
  - \(H_2O(l) \leftrightarrow H^+(aq) + OH^-(aq)\)
  
  Always forms “\(H_3O^+\)”

- Therefore, “\(H^+\)” is synonymous with “\(H_3O^+\)”

- The self ionization of water occurs at a very low, constant, measurable rate!

- For pure water, at 25°C...
  - \([H_3O^+] = 1.0 \times 10^{-7} \text{ M (0.0000001 mols/L)}\)
  - \([OH^-] = 1.0 \times 10^{-7} \text{ M (0.0000001 mols/L)}\)
In solutions, $[\text{H}^+]$ & $[\text{OH}^-]$ have an Indirect Relationship

- When the $[\text{H}_3\text{O}^+]$ goes up, the $[\text{OH}^-]$ goes down.
- When the $[\text{H}_3\text{O}^+]$ goes down, the $[\text{OH}^-]$ goes up.
- This relationship can be summarized by...

$$[\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.0 \times 10^{-14}$$
Mathematical Relationships

• *This is true for all aqueous solutions!*

• For example, pure water...
  
  • \([H_3O^+] = 1.0 \times 10^{-7} \text{ M}\)
  
  • \([OH^-] = 1.0 \times 10^{-7} \text{ M}\)

\[
(1.0 \times 10^{-7} \text{ M}) \times (1.0 \times 10^{-7} \text{ M}) = 1.0 \times 10^{-14}
\]

• 1.0×10⁻¹⁴ is a constant for all aqueous solutions!

• \(K_W = 1.0 \times 10^{-14}\)

  • \(K_W\) is called the “Ion-product Constant for Water”

• Therefore, for acids & bases, the product of the \([H_3O^+]\)
  
  \(\text{times the } [OH^-]\) is always equal to 1.00×10⁻¹⁴ M
Visualizing the Relationship

\[ [\text{H}^+] \ & \ [\text{OH}^-] \ \text{in 3 Solutions...} \]
The Big Idea

- All solutions contain
  - Some acid \(\text{H}_3\text{O}^+\)
  - Some base \(\text{OH}^-\), and
  - The product of their concentrations is always \(1.0 \times 10^{-14} (K_W)\)
Practice Problem #1

What is the \([\text{OH}^-]\) in a \(3.00 \times 10^{-5}\) M solution of HCl?

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

\(3.00\times10^{-5}\) M HCl

• HCl ionizes completely
• Therefore, \([\text{H}^+] = 3.00\times10^{-5}\) M
• If \([\text{H}_3\text{O}^+] \times [\text{OH}^-] = 1.0\times10^{-14}\)
• Then...

\[
[\text{OH}^-] = \frac{1.0\times10^{-14}}{3.00\times10^{-5}}
\]

\[
[\text{OH}^-] = 3.33\times10^{-10}
\]

\([\text{H}_3\text{O}^+]) > [\text{OH}^-]

Therefore, Acidic Solution
**Practice Problem #2**

What is the $\left[ H_3O^+ \right]$ in a $6.00 \times 10^{-4}$ M solution of the strong base NaOH?

NaOH(s) + H$_2$O $\rightarrow$ Na$^+$$_{(aq)}$ + OH$^-$_{(aq)}

6.00×10$^{-4}$ M NaOH

• NaOH ionizes completely
• Therefore, $[OH^-] = 6.00 \times 10^{-4}$ M
• If $[H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}$
• Then...

$$[H_3O^+] = \frac{1.0 \times 10^{-14}}{6.00 \times 10^{-4}}$$

$$[H_3O^+] = 1.67 \times 10^{-11}$$

$[OH^-] > [H_3O^+]$

Therefore, Basic (Alkaline) Solution
Practice Problem #3

An aqueous solution is prepared by dissolving 86.5 g of Chloric Acid in water to make 3.25 L of solution. What is \([\text{OH}^-]\) in this solution?

\[
\frac{86.5 \text{ g HClO}_3}{1} \times \frac{1 \text{ mol HClO}_3}{84.458 \text{ g HClO}_3} = 1.02 \text{ mol HClO}_3
\]

\[
M = \frac{1.02 \text{ mols HClO}_3}{3.25 \text{ L}} = 0.314 \text{ M HClO}_3
\]

\[
[\text{H}_3\text{O}^+] = 0.314 \text{ M}
\]

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{0.314}
\]

\[
[\text{OH}^-] = 3.18 \times 10^{-14}
\]

\[
[\text{H}_3\text{O}^+] > [\text{OH}^-]
\]

Therefore, Acidic Solution
Stating concentrations can be confusing and inefficient!

For example

“In solution A, the concentration of Hydronium is $1.67 \times 10^{-11}$ M and the concentration of Hydroxide is $6.00 \times 10^{-4}$ M.”
The pH Scale!!!
**pH** is a statement of concentration

- **pH** – a measure of concentration of H⁺ (a.k.a. H₃O⁺)

\[
pH = -\log[H^+] \]

\[
[H_3O^+] = 1.67 \times 10^{-11} \text{ M}
\]

Calculator looks like... “-log(1.67E-11)”

\[
pH = 10.777 \text{ (3 sig figs, count only decimals)}
\]
The pH scale

- Numbers range from 1 to 14
- A pH of 7.00 = Neutral
  - \([H^+] = 1.0 \times 10^{-7}\)
  - \([OH^{-}] = 1.0 \times 10^{-7}\)
  - \([H^+] = [OH^{-}]\)
- pH numbers lower than 7.00 are acidic
  - \([H^+] > [OH^{-}]\)
- pH numbers higher than 7.00 are alkaline
  - \([OH^{-}] > [H^+]\)
<table>
<thead>
<tr>
<th>$[H^+]$</th>
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<th>pH</th>
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</thead>
<tbody>
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<td>$1.0\times10^{-1}$</td>
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<tr>
<td>$1.0\times10^{-14}$</td>
<td>0.00000000000001</td>
<td>14</td>
</tr>
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</table>

Each step in the pH scale is a factor of 10.

An acidic solution with a pH of 3 has a concentration of Hydronium that is 10 times greater than a solution with a pH of 4!

Coefficient must be “1” for exponent to match the pH.
Getting $[H^+]$ from pH

If the pH of a solution is 3.70, what is the concentration of Hydronium?

$$[H^+] = 10^{-pH}$$

$$[H^+] = 10^{-3.70}$$

$$[H^+] = 1.995262315 \times 10^{-4}$$

3.70 has 2 significant figures.

$$[H^+] = 2.0 \times 10^{-4}$$
pH Values for Different Concentrations of Hydronium (Logarithmic Scale)
\( pOH \)

We can do the same for the \([OH^-]\)

- \( pOH \) – measures concentration of \( OH^- \)
  \[ pOH = -\log [OH^-] \]
- \( pH + pOH = 14 \)
  \begin{align*}
  \text{pH} & = 14 - pOH \\
  pOH & = 14 - pH 
  \end{align*}
The pOH scale

- Numbers range from 1 to 14
- A pOH of 7.00 = Neutral
  - $[\text{H}^+] = 1.0 \times 10^{-7}$
  - $[\text{OH}^-] = 1.0 \times 10^{-7}$
  - $[\text{H}^+] = [\text{OH}^-]$
- pOH numbers > 7.00 are acidic
  - $[\text{H}^+] > [\text{OH}^-]$
- pOH numbers < 7.00 are alkaline
  - $[\text{OH}^-] > [\text{H}^+]$
Getting $[\text{OH}^-]$ from pOH

If the pOH of a solution is 8.157, what is the concentration of Hydroxide?

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

$$[\text{OH}^-] = 10^{-8.157}$$

$$[\text{OH}^-] = 6.966265141 \times 10^{-9}$$

8.157 has 3 significant figures.

$$[\text{OH}^-] = 6.97 \times 10^{-9}$$
### The pH & pOH relationship

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<thead>
<tr>
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<th>$[OH^-]$</th>
<th>pH</th>
<th>pOH</th>
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<td>13</td>
<td>1</td>
</tr>
<tr>
<td>$1.0 \times 10^{-14}$</td>
<td>1.0</td>
<td>14</td>
<td>0</td>
</tr>
</tbody>
</table>
Big Ideas to Remember

- Molarity (M) of acid = $[\text{H}_3\text{O}^+]$
- Molarity (M) of base = $[\text{OH}^-]$
- $K_w = [\text{H}_3\text{O}^+] [\text{OH}^-] = 1.00 \times 10^{-14} \text{ M}$
- $\text{pH} + \text{pOH} = 14$
Practice Problem #1

What is the pH of a 0.00010 M solution of HCl?

✓ $[H^+] = 1.0 \times 10^{-4}$
✓ $pH = -\log[H^+]$
✓ $-\log(1.0 \times 10^{-4})$
✓ $pH = 4.00$

2 Significant Figures
Practice Problem #2

What’s the pOH of 0.0365 M Ba(OH)$_2$, a strong base?

✓ $[\text{pOH}^-] = 3.65 \times 10^{-2}$
✓ $\text{pOH} = -\log[\text{OH}^-]$
✓ $-\log(3.65 \times 10^{-2})$
✓ $\text{pOH} = 1.438$

3 Significant Figures
Practice Problem #3

• If the pH of a solution is 4.250, what is the \([\text{OH}^-]\)?

• 2 different ways to solve, do both
  • Find pOH, then find \([\text{OH}^-]\) or…
  • Find \([\text{H}_3\text{O}^+]\), then find \([\text{OH}^-]\)

• \(\text{pOH} = 14 - 4.250 = 9.750\)
• \([\text{OH}^-] = 10^{-\text{pOH}} = 10^{-9.750} = 1.78 \times 10^{-10}\)
• Or….
• \([\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.250} = 5.62 \times 10^{-5}\)
• \([\text{OH}^-] = 1.00 \times 10^{-14} / 5.62 \times 10^{-5} = 1.78 \times 10^{-10}\)