Acids and Bases: Chapter 14 & 15
HW:
• Read Ch 14:
• Fill in as much of the acid base table as you can, as you read
<table>
<thead>
<tr>
<th></th>
<th>Conductivity</th>
<th>Reactivity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>high</td>
<td>high</td>
</tr>
<tr>
<td>Acidic acid</td>
<td>low</td>
<td>medium</td>
</tr>
<tr>
<td>Distilled water</td>
<td>none</td>
<td>none</td>
</tr>
<tr>
<td>Ammonium hydroxide</td>
<td>med</td>
<td>none</td>
</tr>
<tr>
<td>Sodium hydroxide</td>
<td>high</td>
<td>none</td>
</tr>
<tr>
<td>Lemon juice</td>
<td>low</td>
<td>low</td>
</tr>
<tr>
<td>Dish soap</td>
<td>none</td>
<td>none</td>
</tr>
<tr>
<td>Tap water</td>
<td>none</td>
<td>none</td>
</tr>
<tr>
<td>Unknown a, b, c</td>
<td>None, High, low</td>
<td>None, none, medium</td>
</tr>
</tbody>
</table>
Start new section of notes:

Acids and Bases: Ch 14 & 15

Vocab:

\( \text{H}^+ \): hydrogen ion (has no electrons)

\( \text{H}_3\text{O}^+ \): hydronium ion

\( \text{OH}^- \): hydroxide (ion)

[ ] = concentration of
**Vocab:**

H\(^+\): hydrogen ion (has no electrons)

H\(_3\)O\(^+\): hydronium ion

OH\(^-\): hydroxide (ion)

[ ] = concentration of

- **acid:** if \([\text{H}_3\text{O}^+] > [\text{OH}^-]\)
- **base:** if \([\text{OH}^-] > [\text{H}_3\text{O}^+]\)

*Neutral*: \([\text{H}_3\text{O}^+] = [\text{OH}^-]\)  *pure water is neutral*

**Dissociation:** breaking apart into ions (ionization)

\[
\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-
\]
ACID-BASE NEUTRALIZATION

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$
Mixing Acids and Bases

What do you get when you mix an acid and a base?

- Recall that acids all have $H^+$
- Recall that bases all have $OH^-$
- What do you get if $H$ and $OH$ combine??
- $HOH = H_2O = WATER!!$

$$HCl + NaOH \rightarrow HOH + NaCl$$

When you mix an acid and a base equitably, you will ALWAYS get WATER and a SALT.
Relationship of \([\text{H}_3\text{O}^+]\) to \([\text{OH}^-]\)

- \([\text{H}_3\text{O}^+] > 10^{-7} \text{ M} > [\text{OH}^-]\) in acidic solution
- \([\text{H}_3\text{O}^+] = 10^{-7} \text{ M} = [\text{OH}^-]\) in neutral solution
- \([\text{H}_3\text{O}^+] < 10^{-7} \text{ M} < [\text{OH}^-]\) in basic solution

- https://www.youtube.com/watch?v=g8jdCWC10vQ

- Acid base indicator: substances whose color is sensitive to pH
- Titration: controlled addition and measurement of amount of solution of a known concentration required to react completely with a known amount of solution of unknown concentration to reach an equivalence point.

Pg 518-519
Naming Acids

All acids must contain **Hydrogen**, but not everything with H is an acid (H₂O!!)

- All acid formulas will be in the format of HX
  - H = hydrogen
  - X = the **anion** used to make the acid

3 Rules to Follow when naming acids:
1. If X is an anion that ends in –ide, the acid name will begin with hydro- and end in the suffix –ic (*binary acid*)
   - None of these will have Oxygen

   - HCl = **Hydrochloric acid**
   - H₂S = **Hydrosulfuric acid**
   - HI = **Hydroiodic acid**
   - HBr = **Hydrobromic acid**
3 Rules to Follow when naming acids:

2. If X is a polyatomic ion that ends in an –ite suffix, the acid name will end in –ous and the root of the name will be the name of the ion.
   - $\text{H}_2\text{SO}_3$
     - Sulfurous acid
   - $\text{HNO}_2$
     - Nitrous acid

3. If X is a polyatomic ion that ends in an –ate suffix, the acid name will end in –ic and the root of the name will be the name of the ion.
   - $\text{HNO}_3$
     - Nitric acid
   - $\text{H}_2\text{SO}_4$
     - Sulfuric acid

• **oxyacid** is an acid that is a compound of hydrogen, oxygen, and a third element.
More practice writing and naming acids

- HCN =
  - Hydrocyanic acid
- Hydrobromic acid
  - HBr
- Phosphorous acid =
  - $\text{H}_3\text{PO}_3$
- $\text{H}_3\text{PO}_4$ =
  - Phosphoric acid
- Carbonic acid =
  - $\text{H}_2\text{CO}_3$
- HF =
  - Hydrofluoric acid
Naming Bases

• Bases produce hydroxide (OH-) in water.
• Bases are named the same as other ionic compounds.

• Practice naming these:
  • NaOH
    – Sodium hydroxide
  • Aluminum hydroxide
    – Al(OH)₃
  • Barium hydroxide
    – Ba(OH)₂
  • LiOH
    – Lithium hydroxide
Warm up:
Make a chart comparing and contrasting acids and bases
More properties

ACIDS

• Tart or sour in taste
• Corrosive (reacts with metals)
• Turns acid/base indicator PINK
• Contains H+ ions
• Examples:
  – Vinegar (acetic acid)
  – Citrus fruits (citric acid)
  – Tea (tannic acid)
  – Carbonated drinks (carbonic acid)

BASES a.k.a. alkaline

• Bitter in taste
• Caustic (burns or dissolves)
• Turns acid/base indicator BLUE
• Contains OH- ions
• Slippery
• Examples:
  – Soap
  – Bleach
  – Stomach meds (ie: TUMS, Rolaids, Milk of Magnesia)

• Both? - dissociate in water (form ions), liquid state, pH etc
# Definitions of Acids and Bases

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<thead>
<tr>
<th>Type</th>
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<th>Base</th>
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<td>$\text{OH}^-$ producer</td>
</tr>
<tr>
<td>Brønsted-Lowry</td>
<td>proton ($\text{H}^+$) donor</td>
<td>proton ($\text{H}^+$) acceptor</td>
</tr>
<tr>
<td>Lewis</td>
<td>electron-pair acceptor</td>
<td>electron-pair donor</td>
</tr>
</tbody>
</table>
Acids taste does not taste Bases taste does taste.
Properties of Acids

1. sour taste.
2. Acids change litmus paper red
3. Some react with active metals and release hydrogen gas, $\text{H}_2$.

$$\text{Ex: } \text{Ba}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + \text{H}_2(g)$$

4. react with bases to produce salts and water.
5. Acids conduct electric current.
6. pH less than 7.
Some Common Industrial Acids

- **Sulfuric Acid**
  - Sulfuric acid is the most commonly produced industrial chemical in the world.

- **Nitric Acid**

- **Phosphoric Acid**

- **Hydrochloric Acid**
  - Concentrated solutions of hydrochloric acid are commonly referred to as *muriatic acid*.

- **Acetic Acid**
  - Pure acetic acid is a clear, colorless, and pungent-smelling liquid known as *glacial acetic acid*. 
Properties of Bases:

1. Aqueous solutions of bases taste bitter.
2. Bases change the color of acid-base indicators.
3. Dilute aqueous solutions of bases feel slippery.
4. Bases react with acids to produce salts and water.
5. Bases conduct electric current.
1. Correct acid and base chart, staple into notebook
2. Finish acid base PHet computer simulation (under “chem notebook” tab on Mrs. D’s website)
3. Work on ch 14-15 review (due Wednesday!)
Arrhenius Acids and Bases

- **Arrhenius acid**: compound that increases the concentration of hydrogen ions, $H^+$, in aqueous solution.

- **Arrhenius base**: substance that increases the concentration of hydroxide ions, $OH^-$, in aqueous solution.

- **All aqueous acids are electrolytes.**
**Strength of Acids**

- **strong acid**: ionizes completely in aqueous solution.
  - a strong acid is a strong electrolyte (conducts)
  - Ex: $\text{HClO}_4$, $\text{HCl}$, $\text{HNO}_3$

- **weak acid** releases few hydrogen ions in aqueous solution.
  - hydronium ions, anions, and dissolved acid molecules in aqueous solution
  - *Organic acids* (—COOH), such as acetic acid
<table>
<thead>
<tr>
<th>Strong acids</th>
<th>Weak acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{HI} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{I}^-$</td>
<td>$\text{HSO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{SO}_4^{2-}$</td>
</tr>
<tr>
<td>$\text{HClO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{ClO}_4^-$</td>
<td>$\text{H}_3\text{PO}_4 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{H}_2\text{PO}_4^-$</td>
</tr>
<tr>
<td>$\text{HBr} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Br}^-$</td>
<td>$\text{HF} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{F}^-$</td>
</tr>
<tr>
<td>$\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-$</td>
<td>$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-$</td>
</tr>
<tr>
<td>$\text{H}_2\text{SO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HSO}_4^-$</td>
<td>$\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HCO}_3^-$</td>
</tr>
<tr>
<td>$\text{HClO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{ClO}_3^-$</td>
<td>$\text{H}_2\text{S} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{HS}^-$</td>
</tr>
<tr>
<td></td>
<td>$\text{HCN} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CN}^-$</td>
</tr>
<tr>
<td></td>
<td>$\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CO}_3^{2-}$</td>
</tr>
</tbody>
</table>
Aqueous Solutions of Bases

- **Strong Bases**: Most bases are ionic compounds containing metal cations and the hydroxide anion, $\text{OH}^-$.  
  - dissociate in water

\[
\text{NaOH(s)} \xrightarrow{H_2O} \text{Na}^+(aq) + \text{OH}^-(aq)
\]

- **Weak Base**: Ammonia, $\text{NH}_3$, is molecular (not an ion)  
  - Ammonia produces hydroxide ions when it reacts with water molecules.

\[
\text{NH}_3(aq) + \text{H}_2\text{O(l)} \leftrightarrow \text{NH}_4^+(aq) + \text{OH}^-(aq)
\]
<table>
<thead>
<tr>
<th>Strong bases</th>
<th>Weak bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ca(OH)$_2$ → Ca$^{2+}$ + 2OH$^-$</td>
<td>NH$_3$ + H$_2$O ⇌ NH$_4^+$ + OH$^-$</td>
</tr>
<tr>
<td>Sr(OH)$_2$ → Sr$^{2+}$ + 2OH$^-$</td>
<td>C$_6$H$_5$NH$_2$ + H$_2$O ⇌ C$_6$H$_5$NH$_3^+$ + OH$^-$</td>
</tr>
<tr>
<td>Ba(OH)$_2$ → Ba$^{2+}$ + 2OH$^-$</td>
<td></td>
</tr>
<tr>
<td>NaOH → Na$^+$ + OH$^-$</td>
<td></td>
</tr>
<tr>
<td>KOH → K$^+$ + OH$^-$</td>
<td></td>
</tr>
<tr>
<td>RbOH → Rb$^+$ + OH$^-$</td>
<td></td>
</tr>
<tr>
<td>CsOH → Cs$^+$ + OH$^-$</td>
<td></td>
</tr>
</tbody>
</table>
Brønsted-Lowry Acids and Bases

- **Brønsted-Lowry acid**: molecule or ion that is a *proton donor.*
- **Brønsted-Lowry base**: molecule or ion that is a *proton acceptor.*

- Hydrogen chloride acts as a Brønsted-Lowry acid when it reacts with ammonia.
- Ammonia accepts a proton from the hydrochloric acid. It acts as a Brønsted-Lowry base.

\[
\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{Cl}^-
\]
Brønsted-Lowry Acids and Bases, continued

• Which is the acid and which is the base?

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

• Water can act as a Brønsted-Lowry acid, it give/donates a proton
• The OH\(^-\) ion produced in solution by Arrhenius hydroxide bases (NaOH) is the Brønsted-Lowry base.
  • The OH\(^-\) ion can accept a proton
Lewis Acids and Bases

- The Lewis definition is the broadest of the three acid definitions.

- **Lewis acid**: accepts an electron pair to form a covalent bond.
  - A bare proton (hydrogen ion) is a Lewis acid
    \[
    \text{H}^+(aq) + :\text{NH}_3(aq) \rightarrow [\text{H}—\text{NH}_3]^+(aq) \text{ or } [\text{NH}_4]^+(aq)
    \]

- **Lewis base**: donates an electron pair to form a covalent bond
Compare acids and bases in terms of ions produced, protons, and electrons

<table>
<thead>
<tr>
<th>Acid</th>
<th>Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>• $\text{H}_3\text{O}^+$</td>
<td>• $\text{OH}^-$</td>
</tr>
<tr>
<td>• $\text{H}^+$</td>
<td>• Proton acceptor</td>
</tr>
<tr>
<td>• proton donor</td>
<td>• Electron pair donor</td>
</tr>
<tr>
<td>• Electron pair acceptor</td>
<td></td>
</tr>
</tbody>
</table>
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</tr>
</tbody>
</table>
Conjugate Acids and Bases

• The species that remains after a Brønsted-Lowry acid has given up a proton is the **conjugate base** of that acid.

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

acid \hspace{1cm} \text{conjugate base}

• Brønsted-Lowry acid-base reactions involve two acid-base pairs, known as conjugate acid-base pairs.

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

\[
\text{acid}_1 \hspace{1cm} \text{base}_2 \hspace{1cm} \text{base}_1 \hspace{1cm} \text{acid}_2
\]
The stronger an acid is, the weaker its conjugate base.

The stronger a base is, the weaker its conjugate acid.

\[
\text{HCl}(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]
• Proton transfer reactions favor the production of the weaker acid and the weaker base.

\[
\text{HClO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{ClO}_4^-(aq)
\]

stronger acid     stronger base     weaker acid     weaker base

• The reaction to the right is more favorable

\[
\text{CH}_3\text{COOH}(aq) + \text{H}_2\text{O}(l) \leftrightarrow \text{H}_3\text{O}^+(aq) + \text{CH}_3\text{COO}^{-}(aq)
\]

weaker acid     weaker base     stronger acid     stronger base

• The reaction to the left is more favorable
<table>
<thead>
<tr>
<th>Conjugate acid</th>
<th>Formula</th>
<th>Conjugate base</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydriodic acid*</td>
<td>HI</td>
<td>iodide ion</td>
<td>I(^-)</td>
</tr>
<tr>
<td>perchloric acid*</td>
<td>HClO(_4)</td>
<td>perchlorate ion</td>
<td>ClO(_4)</td>
</tr>
<tr>
<td>hydrobromic acid*</td>
<td>HBr</td>
<td>bromide ion</td>
<td>Br(^-)</td>
</tr>
<tr>
<td>hydrochloric acid*</td>
<td>HCl</td>
<td>chloride ion</td>
<td>Cl(^-)</td>
</tr>
<tr>
<td>sulfuric acid*</td>
<td>H(_2)SO(_4)</td>
<td>hydrogen sulfate ion</td>
<td>HSO(_4)</td>
</tr>
<tr>
<td>chloric acid*</td>
<td>HClO(_3)</td>
<td>chlorate ion</td>
<td>ClO(_3)</td>
</tr>
<tr>
<td>nitric acid*</td>
<td>HNO(_3)</td>
<td>nitrate ion</td>
<td>NO(_3)</td>
</tr>
<tr>
<td>hydronium ion</td>
<td>H(_3)O(^+)</td>
<td>water</td>
<td>H(_2)O</td>
</tr>
<tr>
<td>chlorous acid</td>
<td>HClO(_2)</td>
<td>chlorite ion</td>
<td>ClO(_2)</td>
</tr>
<tr>
<td>hydrogen sulfate ion</td>
<td>HSO(_4)</td>
<td>sulfate ion</td>
<td>SO(_4)(^-)</td>
</tr>
<tr>
<td>phosphoric acid</td>
<td>H(_3)PO(_4)</td>
<td>dihydrogen phosphate ion</td>
<td>H(_2)PO(_4)</td>
</tr>
<tr>
<td>hydrofluoric acid</td>
<td>HF</td>
<td>fluoride ion</td>
<td>F(^-)</td>
</tr>
<tr>
<td>acetic acid</td>
<td>CH(_3)COOH</td>
<td>acetate ion</td>
<td>CH(_3)COO(^-)</td>
</tr>
<tr>
<td>carbonic acid</td>
<td>H(_2)CO(_3)</td>
<td>hydrogen carbonate ion</td>
<td>HCO(_3)</td>
</tr>
<tr>
<td>hydrosulfuric acid</td>
<td>H(_2)S</td>
<td>hydrosulfide ion</td>
<td>HS(^-)</td>
</tr>
<tr>
<td>dihydrogen phosphate ion</td>
<td>H(_2)PO(_4)</td>
<td>hydrogen phosphate ion</td>
<td>HPO(_2)(^-)</td>
</tr>
<tr>
<td>hypochlorous acid</td>
<td>HClO</td>
<td>hypochlorite ion</td>
<td>ClO(_3)</td>
</tr>
<tr>
<td>ammonium ion</td>
<td>NH(_4)</td>
<td>ammonia</td>
<td>NH(_3)</td>
</tr>
<tr>
<td>hydrogen carbonate ion</td>
<td>HCO(_3)</td>
<td>carbonate ion</td>
<td>CO(_3)(^-)</td>
</tr>
<tr>
<td>hydrogen phosphate ion</td>
<td>HPO(_2)(^-)</td>
<td>phosphate ion</td>
<td>PO(_4)(^3-)</td>
</tr>
<tr>
<td>water</td>
<td>H(_2)O</td>
<td>hydroxide ion</td>
<td>OH(^-)</td>
</tr>
<tr>
<td>ammonia</td>
<td>NH(_3)</td>
<td>amide ion†</td>
<td>NH(_3)</td>
</tr>
<tr>
<td>hydrogen</td>
<td>H(_2)</td>
<td>hydride ion†</td>
<td>H(^-)</td>
</tr>
</tbody>
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* Strong acids
† Strong bases
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</table>
For the 2 equation, label the acids and bases and their conjugates (can also use notation \text{acid}_1, \text{base}_1, \text{acid}_2 \text{ base}_2)

1. \( \text{HClO}_4 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{ClO}_4^- \)

2. \( \text{C}_5\text{H}_5\text{N} + \text{H}_2\text{O} \rightarrow [\text{C}_5\text{H}_5\text{NH}]^+ + \text{OH}^- \)

3. What is pH?
Neutralization Reactions
Strong Acid-Strong Base Neutralization

- In aqueous solutions, **neutralization** is the reaction of hydronium ions and hydroxide ions to form water molecules.

- A **salt** is an ionic compound composed of a cation from a base and an anion from an acid.

\[
\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)
\]
Acid Rain

- NO, NO$_2$, CO$_2$, SO$_2$, and SO$_3$ gases from industrial processes can dissolve in atmospheric water to produce acidic solutions.

  - example: $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

- Acid Rain: very acidic rain
- Acid rain can erode statues and affect ecosystems.
Chapter 15
Measuring pH

- **pH =** power of Hydrogen

- Scale goes from 0-14
  - 0-6 = acid
  - **7.0 = neutral**
  - 8-14 = base
  - From 1 to 2 = increased by a factor of 10
pH = −\log [H_3O^+]

pOH = −\log [OH^-]

eample 1: a neutral solution has a [H_3O^+] = 1 \times 10^{-7}, find pH

pH = −\log [H_3O^+] = −\log(1 \times 10^{-7}) = −(-7.0) = 7.0

Example 2: [H_3O^+] = 1 \times 10^{-11}

= −\log(1 \times 10^{-11})

= 11

• For every number increase or decrease on the pH scale, the concentration of H^= increase or decrease by a power of 10
## pH Values as Specified [H₃O⁺]

<table>
<thead>
<tr>
<th>Solution</th>
<th>[H₃O⁺] (M)</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.00 L of H₂O</td>
<td>1.00 × 10⁻⁷</td>
<td>7.00</td>
</tr>
<tr>
<td>0.100 mol HCl in 1.00 L of H₂O</td>
<td>1.00 × 10⁻¹</td>
<td>1.00</td>
</tr>
<tr>
<td>0.0100 mol HCl in 1.00 L of H₂O</td>
<td>1.00 × 10⁻²</td>
<td>2.00</td>
</tr>
<tr>
<td>0.100 mol NaCl in 1.00 L of H₂O</td>
<td>1.00 × 10⁻⁷</td>
<td>7.00</td>
</tr>
<tr>
<td>0.0100 mol NaOH in 1.00 L of H₂O</td>
<td>1.00 × 10⁻¹²</td>
<td>12.00</td>
</tr>
<tr>
<td>0.100 mol NaOH in 1.00 L of H₂O</td>
<td>1.00 × 10⁻¹³</td>
<td>13.00</td>
</tr>
</tbody>
</table>
Concentrations and $K_w$ (ionization constant)

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

<table>
<thead>
<tr>
<th>Solution</th>
<th>$[H_3O^+]$ (M)</th>
<th>$[OH^-]$ (M)</th>
<th>$K_w = [H_3O^+][OH^-]$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pure water</td>
<td>$1.0 \times 10^{-7}$</td>
<td>$1.0 \times 10^{-7}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
<tr>
<td>0.10 M strong acid</td>
<td>$1.0 \times 10^{-1}$</td>
<td>$1.0 \times 10^{-13}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
<tr>
<td>0.010 M strong acid</td>
<td>$1.0 \times 10^{-2}$</td>
<td>$1.0 \times 10^{-12}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
<tr>
<td>0.10 M strong base</td>
<td>$1.0 \times 10^{-13}$</td>
<td>$1.0 \times 10^{-1}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
<tr>
<td>0.010 M strong base</td>
<td>$1.0 \times 10^{-12}$</td>
<td>$1.0 \times 10^{-2}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
<tr>
<td>0.025 M strong acid</td>
<td>$2.5 \times 10^{-2}$</td>
<td>$4.0 \times 10^{-13}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
<tr>
<td>0.025 M strong base</td>
<td>$4.0 \times 10^{-13}$</td>
<td>$2.5 \times 10^{-2}$</td>
<td>$1.0 \times 10^{-14}$</td>
</tr>
</tbody>
</table>
Today

• Finish acid base lab
• Work on PhET simulation
  – This becomes homework due Monday
Hydronium Ions and Hydroxide Ions

Self-Ionization of Water

• **self-ionization of water:** two water molecules produce a hydronium ion and a hydroxide ion by transfer of a proton.

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

• In water at 25°C, \([\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{ M and } [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M.}\)

• The *ionization constant of water*, \(K_w\), is expressed by the following equation.

\[
K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}
\]
**Example**

1. Calculate the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ concentrations of a solution of $1.0 \times 10^{-4}$ M HCl?

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

   - 1 to 1 mole ratio, so $[\text{HCl}]$ will equal $[\text{H}_3\text{O}^+] = 1.0 \times 10^{-4}$ M

   - $[1.0 \times 10^{-4}] \times [\text{OH}^-] = 1.0 \times 10^{-14}$

   - $[1.0 \times 10^{-14}] / [1.0 \times 10^{-4}] = 1 \times 10^{-10}$ M HCl

2. Calculate the $[\text{H}_3\text{O}^+]$ of a solution given that $[\text{OH}^-] = 1.0 \times 10^{-2}$

   - $[\text{H}_3\text{O}^+][1.0 \times 10^{-2}] = 1.0 \times 10^{-14}$
1. Correct acid and base chart, staple into notebook

2. Finish acid base PHet computer simulation (under “chem notebook” tab on Mrs. D’s website)

3. Work on ch 14-15 review (due Wednesday, test Friday!)
• Solutions in which \([H_3O^+] = [OH^-]\) is *neutral*.

• Solutions in which the \([H_3O^+] > [OH^-]\) are *acidic*.
  
  • \([H_3O^+] > 1.0 \times 10^{-7} \text{ M}\)

• Solutions in which the \([OH^-] > [H_3O^+]\) are *basic*.
  
  • \([OH^-] > 1.0 \times 10^{-7} \text{ M}\)
Titration

- **Neutralization**: bring solution closer to neutral
  - occurs when hydronium ions and hydroxide ions are supplied in equal numbers by reactants.

\[
\text{H}_3\text{O}^+(aq) + \text{OH}^-(aq) \rightarrow 2\text{H}_2\text{O}(l)
\]

- **Titration**: controlled addition and measurement of the amount of a solution of known concentration required to react completely with a measured amount of a solution of unknown concentration.
• **equivalence point:** point at which the two solutions used in a titration are present in chemically equivalent amounts
  • (pink color of Phenolphthalein appears).
• The point in a titration at which an indicator changes color is called the **end point** of the indicator. pH=7 at this point

• Kinda how an antacid works!
Titration Curve for a Strong Acid and a Strong Base

- [http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s5_5.swf](http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s5_5.swf)
• Online links

• PHET simulation:
  • http://phet.colorado.edu/sims/html/ph-scale/latest/ph-scale_en.html

• General info and quiz:
  http://www.elmhurst.edu/~chm/vchembook/184ph.html

• Titration:
  http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s5_5.swf
Objectives: Ch 14

- List five general properties of aqueous acids and bases.
- Name common binary acids and oxyacids, given their chemical formulas.
- List five acids commonly used in industry and the laboratory, and give two properties of each.
- Define acid and base according to Arrhenius’ s theory of ionization.
- Explain the differences between strong and weak acids and bases.
- Define and recognize Brønsted-Lowry acids and bases.
- Define a Lewis acid and a Lewis base.
- Name compounds that are acids under the Lewis definition but are not acids under the Brønsted-Lowry definition.
- Describe a conjugate acid, a conjugate base, and an amphoteric compound.
- Explain the process of neutralization.
- Define acid rain, give examples of compounds that can cause acid rain, and describe effects of acid rain.
Objectives Ch 15

- **Describe** the self-ionization of water.
- **Define** pH, and give the pH of a neutral solution at 25°C.
- **Explain** and use the pH scale.
- **Given** \([H_3O^+]/[OH^-]\), **find** pH.
- **Given** pH, **find** \([H_3O^+]/[OH^-]\).

- **Describe** how an acid-base indicator functions.
- **Explain** how to carry out an acid-base titration.
- **Calculate** the molarity of a solution from titration data.

- **Describe** the self-ionization of water.
- **Define** pH, and give the pH of a neutral solution at 25°C.
- **Explain** and use the pH scale.
- **Given** \([H_3O^+]/[OH^-]\), **find** pH.
- **Given** pH, **find** \([H_3O^+]/[OH^-]\).
Hey, Litmus Paper, why so blue?

SIGN #93 THAT YOU'RE DEALING WITH SOMEBODY WHO'S HAZY ON THE SCIENCE.

"WHAT'S THE BIG DEAL ABOUT ACID RAIN? CAN'T WE JUST MAKE ALKALINE RAIN TO COUNTERACT IT?"
CA State Standards

• 5a: Students know the observable properties of acids, bases, and salt solutions.
• 5d: Students know how to use the pH scale to characterize acid and base solutions.
Common Core Standards

Physical Science

• HS-PS1-3. Plan and conduct an investigation to gather evidence to compare the structure of substances at the bulk scale to infer the strength of electrical forces between particles.

CA State Standards

• 5a: Students know the observable properties of acids, bases, and salt solutions.

• 5d: Students know how to use the pH scale to characterize acid and base solutions.
Common Core Standards
Physical Science

• HS-PS2-6 Communicate scientific and technical information about why the molecular-level structure is important in the functioning of designed materials.

• HS-PS1-6. Refine the design of a chemical system by specifying a change in conditions that would produce increased amounts of products at equilibrium.*