Demo of Hydrolysis AlCl₃, CaCl₂, Na₂CO₃, etc.

Hydrolysis:

- Reaction between a salt (ion or ions in a salt) and water to produce an acidic or basic solution.
- Net ionic equations for hydrolysis:

An ion + water $\rightarrow$ a molecule or ion + H₂O⁺ or OH⁻

SPECTATORS - ions which do NOT hydrolyze (need periodic table and acid table to find these)

<table>
<thead>
<tr>
<th>Spectator Cations (look on per. table)</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Group 1</strong> (Alkali Metal ions) eg. Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Fr⁺</td>
</tr>
<tr>
<td><strong>Group 2</strong> (Alkaline Earth ions) eg. Be²⁺, Mg²⁺, Ca²⁺, Ba²⁺, Sr²⁺, Ra²⁺</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Spectator Anions (look on acid table)</th>
</tr>
</thead>
<tbody>
<tr>
<td>- Conjugate bases of strong acids.</td>
</tr>
<tr>
<td>- <strong>Top 5 ions on the right</strong> side of table.</td>
</tr>
<tr>
<td>- ClO₄⁻, F⁻, Br⁻, Cl⁻, NO₃⁻. (HSO₄⁻ is not a spectator – it is amphiprotic – will be dealt with later)</td>
</tr>
</tbody>
</table>

- **spectators are eliminated** in net ionic equations (NIE’s) for hydrolysis!

Process – if given salt (dissociate $\rightarrow$ eliminate $\rightarrow$ evaluate)

1. Write dissociation equation
2. Eliminate spectactors
3. Remaining ions $\rightarrow$ left side of table – undergo acidic hydrolysis is $\rightarrow$ produce H₂O⁺ $\rightarrow$ right side of table – undergo base hydrolysis – produce OH⁻ $\rightarrow$ amphiprotic – determine Kₐ and Kₐ to find dominant hydrolysis.

Examples:

Determining A, B, or N

Is the salt NaF Acidic, basic or neutral in water?

Dissociation: NaF $\rightarrow$ Na⁺ + F⁻

- so NaF is basic

Is the salt NH₄NO₃ acidic, basic or neutral in aqueous solution?

Dissociation: NH₄NO₃ $\rightarrow$ NH₄⁺ + NO₃⁻

- so NH₄NO₃ is acidic.
Is the salt KCl acidic, basic or neutral?

Dissociation: KCl $\rightarrow$ K$^+$ + Cl$^-$

- since neither ion undergoes hydrolysis, this salt is **NEUTRAL**.

**Cations Which Hydrolyze**

- **Hydrated cations**
  - metals from center of the periodic table (transition metals) are smaller ions and have larger charges
  - this attracts H$_2$O molecules

eg.) Fe$^{3+}$ (iron (III) or ferric ion)

Hydration: Fe$^{3+}$ + 6H$_2$O $\rightarrow$ Fe(H$_2$O)$_6^{3+}$

This ion acts as a weak acid (see it ~ 13th down on the acid table.)

The equation for the hydrolysis of hexaaquoiron or ferric ion is:

Fe(H$_2$O)$_6^{3+}$(aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+$(aq) + Fe(H$_2$O)$_5$(OH$^2+$(aq)

3 Common **Hydrated** cations (on left of acid chart):

iron(III) Fe$^{3+}$ forms Fe(H$_2$O)$_6^{3+}$ hexaaquoiron(III)
Chromium(III) Cr$^{3+}$ forms Cr(H$_2$O)$_6^{3+}$ hexaaquochromium(III)
Aluminum Al$^{3+}$ forms Al(H$_2$O)$_6^{3+}$ hexaaquoaluminum

Can appear in either form in salts

Eg.) AlCl$_3$ is the same as Al(H$_2$O)$_6$Cl$_3$

**Another Acidic Cation**

NH$_4^+$

**Hydrolysis equation:** NH$_4^+$(aq) + H$_2$O(l) $\rightleftharpoons$ H$_3$O$^+$(aq) + NH$_3$(aq)

List the 4 hydrolyzing cations on the acid table:

<table>
<thead>
<tr>
<th>Cation</th>
<th>Hydrolysis Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>NH$_4^+$</td>
<td>$\text{NH}_4^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{NH}_3$</td>
</tr>
</tbody>
</table>
ANIONS WHICH HYDROLYZE
Looking on the RIGHT side of the ACID TABLE:

<table>
<thead>
<tr>
<th>Base</th>
<th>( K_a )</th>
<th>Strength of base</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}^+ + \text{ClO}_4^- )</td>
<td>very large</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{I}^- )</td>
<td>very large</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{Br}^- )</td>
<td>very large</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{Cl}^- )</td>
<td>very large</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{NO}_3^- )</td>
<td>very large</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{HSO}_4^- )</td>
<td>very large</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{H}_2\text{O} )</td>
<td>1.0</td>
<td>Strong</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{IO}_3^- )</td>
<td>(1.7 \times 10^{-1} )</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{HOOCCOO}^- )</td>
<td>(5.4 \times 10^{-2} )</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{HSO}_3^- )</td>
<td>(1.7 \times 10^{-2} )</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{PO}_4^{3-} )</td>
<td>(4.4 \times 10^{-13} )</td>
<td>Weak</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{OH}^- )</td>
<td>(1.0 \times 10^{-14} )</td>
<td>Strong</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{O}^{2-} )</td>
<td>very small</td>
<td>Strong</td>
</tr>
<tr>
<td>( \text{H}^+ + \text{NH}_2^- )</td>
<td>very small</td>
<td>Strong</td>
</tr>
</tbody>
</table>

All of the anions in this section from \( \text{IO}_3^- \) down to \( \text{S}_2^- \) will undergo base hydrolysis. Anions that are NOT amphiprotic will act as weak bases in water. We will deal with amphiprotic anions (eg. \( \text{HCOO}^- \)) later.

Some examples of net-ionic hydrolysis equations for these would be:

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

\[
\text{NO}_2^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HNO}_2(aq) + \text{OH}^-(aq)
\]

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

Salts which contain these anions may also be basic (depending on the cation). When you get a salt, you must dissociate it, eliminate spectators and then look for hydrolysis of any remaining ions.

Eg.) Determine whether the salt sodium carbonate (\( \text{Na}_2\text{CO}_3 \)) is acidic, basic or neutral in aqueous solution.

First dissociate the salt: \( \text{Na}_2\text{CO}_3 \rightarrow 2\text{Na}^+(aq) + \text{CO}_3^{2-}(aq) \)

Spectator Cation Eliminate!  

Hydrolyzing Anion(weak base)

The net-ionic equation for the hydrolysis taking place in this salt would be:

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

and the salt would act as a weak base in water. Remember that “net-ionic” means that any spectator ions have been removed!
Write the net-ionic equation for the hydrolysis taking place in aqueous magnesium sulphate:

\[
\text{Magnesium sulphate} \rightarrow \text{Mg}^{2+} + \text{SO}_4^{2-}
\]

Hydrolysis When BOTH Cation and Anion hydrolyze

Eg. Is the salt ammonium nitrite \(\text{NH}_4\text{NO}_2\) acidic, basic or neutral?

Of course we start out by dissociating: \(\text{NH}_4\text{NO}_2 \rightarrow \text{NH}_4^{+} + \text{NO}_2^{-}\)

Remember that \(\text{NH}_4^{+}\) produces \(\text{H}_3\text{O}^{+}\) (\(\text{NH}_4^{+} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^{+} + \text{NH}_3\)) (equation 1)

And \(\text{NO}_2^{-}\) produces \(\text{OH}^{-}\) (\(\text{NO}_2^{-} + \text{H}_2\text{O} \rightleftharpoons \text{HNO}_2 + \text{OH}^{-}\)) (equation 2)

The Ka for \(\text{NH}_4^{+}\) tells how much \(\text{H}_3\text{O}^{+}\) it produces (The \(K_{eq}\) for equation 1 is the Ka of \(\text{NH}_4^{+}\))

The Kb for \(\text{NO}_2^{-}\) tells how much \(\text{OH}^{-}\) it produces (The \(K_{eq}\) for equation 2 is the Kb of \(\text{NO}_2^{-}\))

The Ka for \(\text{NH}_4^{+}\) is \(5.6 \times 10^{-10}\) (look up \(\text{NH}_4^{+}\) on the left side of the table and it’s Ka is on the right)

The Kb for \(\text{NO}_2^{-}\) must be calculated: \(\text{Kb (NO}_2^{-}) = \frac{\text{Kw}}{\text{Ka (HNO}_2)} = \frac{1.0 \times 10^{14}}{4.6 \times 10^{4}} = 2.2 \times 10^{11}\)

Since the Ka of \(\text{NH}_4^{+}\) > Kb of \(\text{NO}_2^{-}\) We can say that this salt is ACIDIC

So, in summary:

<table>
<thead>
<tr>
<th>If</th>
<th>Then the salt is:</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\text{Ka (cation)} &gt; \text{Kb (anion)})</td>
<td>Acidic</td>
</tr>
<tr>
<td>(\text{Kb (anion)} &gt; \text{Ka (cation)})</td>
<td>Basic</td>
</tr>
<tr>
<td>(\text{Ka (cation)} = \text{Kb (anion)})</td>
<td>Neutral</td>
</tr>
</tbody>
</table>

Determine whether the salt \(\text{NH}_4\text{CN}\) (ammonium cyanide) is acidic, basic or neutral.

Hydrolysis of Amphiprotic Anions

Amphiprotic Anions \(\rightarrow\) Start with “H” and have a “+” charge.

\(\text{Eg. } \text{HSO}_4^{-}, \text{HSO}_3^{-}, \text{H}_2\text{PO}_4^{-}, \text{HPO}_4^{2-}, \text{HS}^{-}\text{ etc.}\)

Amphiprotic Anions hydrolyze as acids to produce \(\text{H}_3\text{O}^{+}\) but they also hydrolyze as bases to produce \(\text{OH}\)

So, how can we tell whether they are acidic or basic or neutral? We need to determine the predominant hydrolysis. See the next page…
Find the Ka of the ion. (Look for ion on the LEFT SIDE of the acid table, read Ka on the right.)

Find the Kb of the ion. (Look for the ion on the RIGHT SIDE of the table and use:
\[ \text{Kb} = \frac{\text{Kw}}{\text{Ka(conj. acid)}} \]

<table>
<thead>
<tr>
<th>If</th>
<th>Then the predominant hydrolysis is:</th>
<th>And, in aqueous solution, the ion:</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ka (the ion) &gt; Kb (the ion)</td>
<td><strong>ACID HYDROLYSIS</strong></td>
<td>Acts as an Acid</td>
</tr>
<tr>
<td>Kb (the ion) &gt; Ka (the ion)</td>
<td><strong>BASE HYDROLYSIS</strong></td>
<td>Acts as a Base</td>
</tr>
</tbody>
</table>

Eg. Find the predominant hydrolysis of the hydrogen carbonate ion (HCO$_3^-$) and write the net-ionic equation for it.

To find the Ka of HCO$_3^-$, look it up on the left side of table (6th from the bottom). It’s Ka = 5.6 x 10$^{-11}$

To find the Kb of HCO$_3^-$, look it up on the right side of table. (15th from the bottom)

\[ (\text{It’s conjugate acid is H}_2\text{CO}_3 \text{ and the Ka of H}_2\text{CO}_3 = 4.3 \times 10^{-7}) \]

So we calculate the Kb of HCO$_3^-$ using:
\[ \text{Kb(HCO}_3^-) = \frac{\text{Kw}}{\text{Ka(H}_2\text{CO}_3)} = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8} \]

\[
\begin{align*}
\text{So, since Kb (HCO}_3^-) & > \text{ Ka (HCO}_3^-) \text{, the ion HCO}_3^- \text{ predominantly undergoes BASE HYDROLYSIS.} \\
& (2.3 \times 10^{-8}) \quad (5.6 \times 10^{-11})
\end{align*}
\]

And the net-ionic equation for the predominant hydrolysis is:

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

Read p. 144 – 147 in SW & Do Ex. 69 (a-f) and Ex. 70 (a – j), 71, 72 & 73 on p. 148 SW.

Do Worksheet 4-5 (Hydrolysis) & Do Experiment 20-D (Hydrolysis)

**Putting it all Together—Finding the pH in a Salt Solution**

Eg. Calculate the pH of 0.30 M Na$_2$CO$_3$

Step 1: Dissociate and Eliminate any spectators. Identify any ions left as weak acids or weak bases.

\[ \text{Na}_2\text{CO}_3 \rightarrow 2\text{Na}^+ + \text{CO}_3^{2-} \]

Found on RIGHT side of acid table 6th from the bottom.
Undergoes BASE HYDROLYSIS
Step 2: Write HYDROLYSIS EQUATION (Don’t forget that CO$_3^{2-}$ undergoes BASE hydrolysis!)
And an ICE table underneath it:

\[
\begin{array}{c|c|c|c}
[I] & CO_3^{2-} & + & H_2O(l) & \rightleftharpoons & HCO_3^{-} & + & OH^{-} \\
[C] & 0.30 & - x & + x & + x \\
[E] & 0.30 - x & x & x \\
\end{array}
\]

Step 3: Since CO$_3^{2-}$ is a WEAK BASE, we need to calculate the value of $K_b$ for CO$_3^{2-}$:

\[
K_b (CO_3^{2-}) = \frac{K_w}{Ka (HCO_3^{-})} = \frac{1.0 \times 10^{-14}}{5.6 \times 10^{-11}} = 1.786 \times 10^{-4} \text{ (use unrounded value in the next calculation)}
\]

Step 4: Write the $K_b$ expression for the hydrolysis of CO$_3^{2-}$:

\[
K_b = \frac{[HCO_3^{-}] [OH^-]}{[CO_3^{2-}]}
\]

Step 5: Insert equilibrium concentration [E] values from the ICE table into the $K_b$ expression. State any valid assumptions:

\[
K_b = \frac{x^2}{(0.30 - x)} \quad \text{Assume } 0.30 - x \approx 0.30
\]

Step 6: Calculate the value of x. Remember in the ICE table, that x = [OH$^-$]

\[
1.786 \times 10^{-4} = \frac{x^2}{0.30}
\]

\[
x^2 = 0.30 \times 1.786 \times 10^{-4}
\]

\[
[OH^-] = x = \sqrt{0.30 \times 1.786 \times 10^{-4}} = 7.319 \times 10^{-3} \text{ M}
\]

Step 7: Calculate pOH (pOH = - log [OH$^-$])

\[
pOH = -\log (7.319 \times 10^{-3}) = 2.1355
\]

Step 8: Convert to pH (pH = 14.00 - pOH). Express in the correct # of SD’s as justified by data:

\[
pH = 14.00 - 2.1355 = 11.86
\]

Step 9: Make sure your answer makes sense. The salt was a WEAK BASE, so a pH of 11.86 is reasonable!

Now it’s your turn!
Question: Calculate the pH of a 0.24 M solution of the salt aluminum nitrate. Show all your steps. State any assumptions used.

Metal, Non-metal and Metalloid Oxides (also called Anhydrides)

Demonstration of the pH’s of metal and non-metal oxides.

<table>
<thead>
<tr>
<th>Compound</th>
<th>Metal or Non-metal Oxide</th>
<th>Colour in Universal Indicator</th>
<th>Approximate pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aqueous MgO</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous CaO</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous ZnO</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous CO₂</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous NO₂</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aqueous SO₂</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Conclusions:
Metal oxides act as (acids/bases)___________ in aqueous solution.
Non-Metal oxides act as (acids/bases)___________ in aqueous solution.

Explanation:

Group 1 and Group 2 Oxides are ionic. They dissociate to form the oxide ion (O²⁻).

Eg. \( \text{Na}_2\text{O(s)} \rightarrow 2 \text{Na}^+_{(aq)} + \text{O}^{2-}_{(aq)} \)
The oxide (O\textsuperscript{2-}) ion then undergoes 100% hydrolysis (because it’s a strong base)

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

**Another example:** BaO \(\rightarrow\) Ba\textsuperscript{2+}(aq) + O\textsuperscript{2-}(aq)

The oxide (O\textsuperscript{2-}) ion then undergoes 100% hydrolysis (because it’s a strong base)

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

We can also summarize the reactions of group 1 and group 2 metals with water in the form of *formula equations*:

Na\textsubscript{2}O + H\textsubscript{2}O \(\rightarrow\) 2 NaOH  
BaO + H\textsubscript{2}O \(\rightarrow\) Ba(OH)\textsubscript{2}

Write a balanced *formula equation* for the **overall reactions** of the following oxides with water:

**Calcium oxide:** ____________________________________________

**Lithium oxide:** ____________________________________________

**Non-Metal Oxides act as ACIDS in aqueous solution:**

Some common examples of **non-metal oxides**: NO\textsubscript{2}, N\textsubscript{2}O\textsubscript{5}, SO\textsubscript{2}, SO\textsubscript{3}, CO\textsubscript{2}, Cl\textsubscript{2}O

These compounds react with water to form **ACIDS**.

The *formula equations* for some of these are:

\[
\text{SO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_3(aq) \text{ (sulphurous acid)}
\]

\[
2\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(aq) + \text{HNO}_2(aq) \text{ (nitric and nitrous acids)}
\]

Once these acids are formed, they can ionize (strong ones 100%, weak ones < 100%) to form H\textsubscript{3}O\textsuperscript{+} ions.

Eg. \ H\textsubscript{2}SO\textsubscript{3}(aq) + H\textsubscript{2}O(l) \xleftrightarrow{\text{equilibrium}} H\textsubscript{3}O\textsuperscript{+}(aq) + HSO\textsubscript{3}-(aq) ( < 100% ionization since H\textsubscript{2}SO\textsubscript{3} is a weak acid)

Eg. \ HNO\textsubscript{3}(aq) + H\textsubscript{2}O(l) \(\rightarrow\) H\textsubscript{3}O\textsuperscript{+}(aq) + NO\textsubscript{3}-(aq) (100% ionization since HNO\textsubscript{3} is a strong acid)

**Metalloid Oxides** (by staircase)

Eg. Al\textsubscript{2}O\textsubscript{3}, Ga\textsubscript{2}O\textsubscript{3}, GeO\textsubscript{2} These compounds usually have low solubility so not many ions are freed to undergo hydrolysis. So very little hydrolysis occurs so they do not act as acids or bases.

These compounds can react WITH acids or bases. Compounds that can do this are called **amphoteric**.
Anhydrides

Oxide compounds that react as acids or bases in aqueous solution are also called Anhydrides. (an-hydride translates to “without water”) These are compounds that react WITH water to form acidic or basic solutions.

**Acidic Anhydride** — An oxide (“O” containing) compound which reacts with water to form an ACIDIC SOLUTION.

Acidic anhydrides are oxides of elements on the RIGHT side of the periodic table.

Some examples of acidic anhydrides are: \( \text{SO}_2 \), \( \text{SO}_3 \), \( \text{Cl}_2\text{O} \) etc.

And some of their reactions with water are:

\[
\text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(aq) \quad \text{(sulphuric acid—a strong acid)}
\]

\[
2\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_2(aq) + \text{HNO}_3(aq) \quad \text{(nitrous and nitric acids)}
\]

\[
\text{Cl}_2\text{O}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{2HClO} \quad \text{(hypochlorous acid)}
\]

(Note: You should KNOW these equations!)

**Basic Anhydride** — An oxide (“O” containing) compound which reacts with water to form a BASIC SOLUTION.

NOTE: Basic Anhydrides are METAL (LEFT side of Periodic Table) oxides.

Some examples are: \( \text{Na}_2\text{O} \), \( \text{CaO} \), \( \text{MgO} \), \( \text{CaO} \) etc.

Formula equations for some Basic Anhydrides reacting with water:

\[
\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH} \quad \text{(the base is called sodium hydroxide)}
\]

\[
\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 \quad \text{(the base is called calcium hydroxide – sometimes called “hydrated lime”)}
\]

Read p. 184-185 in SW.
Do Ex. 144-145 on p. 185 of SW.

**Acid Rain**

Since our atmosphere naturally contains \( \text{CO}_2 \) (an acidic anhydride), some of this reacts with water (rain) to make the rain slightly acidic:

\[
\text{CO}_2(g) + 2\text{H}_2\text{O}(l) \leftrightarrow \text{H}_3\text{O}^+(aq) + \text{HCO}_3^-(aq)
\]

So natural rainwater (unaffected by human activities) can have a pH as low as 5.6 (due to the \( \text{CO}_2 \) in air)

If rain has a \( \text{pH} < 5.6 \) it is called ACID RAIN.
Acid Rain is caused by Acidic Anhydrides (not counting CO₂) in the air.

The main human sources of acid rain are:

1. Burning fuels containing sulphur.
2. Car exhaust

1. When burning coal or other fuels containing sulphur, the sulphur burns too forming sulphur dioxide:

\[ S(s) + O_2(g) \rightarrow SO_2(g) \]  
(an acidic anhydride)

In the atmosphere further oxidation can occur producing sulphur trioxide:

\[ 2SO_2(g) + O_2(g) \rightarrow 2SO_3(g) \]  
(an acidic anhydride)

In rainwater or cloud droplets two reactions can then occur:

\[ SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq) \]  
(sulphurous acid – a weak acid)

\[ SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq) \]  
(sulphuric acid – a strong acid)

2. In the hot cylinders in an internal combustion engine, N₂ from the air and O₂ from the air react to form nitrogen oxides:

\[ N_2(g) + O_2(g) \rightarrow 2NO(g) \]  
(nitrogen monoxide – an acidic anhydride)

\[ N_2(g) + 2O_2(g) \rightarrow 2NO_2(g) \]  
(nitrogen dioxide – an acidic anhydride)

Nitrogen monoxide can further oxidize in the air to produce nitrogen dioxide:

\[ 2NO(g) + O_2(g) \rightarrow 2NO_2(g) \]

Nitrogen dioxide reacts with rain water or cloud droplets to produce both nitrous and nitric acids:

\[ 2NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq) \]

Some natural sources of Acid Rain

Volcanoes can produce SO₂ into the atmosphere (which produces both H₂SO₃ and H₂SO₄).

Lightning can provide enough energy to cause nitrogen and oxygen in the air to react and form NO₂ (which produces HNO₂ and HNO₃).

Natural Protection

Some areas are not as sensitive to acid rain as others—even when there are major sources of acid rain present! This is because these areas have rocks and soils that are high in “carbonates” or compounds containing carbonate (CO₃²⁻). The CO₃²⁻ acts as a weak base and neutralizes the acid rain to a certain extent:

\[ \text{E.g.,} \quad H_2SO_4(aq) + CaCO_3(s) \rightarrow CaSO_4(s) + CO_2(g) + H_2O(l) \]  
(This neutralizes the H₂SO₄ caused by acid rain)

In some cases, powdered limestone (CaCO₃) is dumped onto lakes that are too acidic and this helps neutralize the acid rain. But the process is expensive.
Problems Associated with Acid Rain (see p. 187-188 in SW.)

- Aquatic life is affected (from bottom of food chain up)
- Forests weakened and killed (Quebec, Germany, Scandinavia)
- Minerals are leached out of the topsoil to lower levels. $\text{Al}^{3+}$ ions are released which are very toxic to fish (mucous in gills) and plants (prevents uptake of other important minerals)
- Metal and stone buildings and statues (especially marble ($\text{CaCO}_3$)) are damaged.
- Acid Rain is carried over large distances (due to high smoke stacks) – can cross international borders

Possible Solutions to the Problem

- International conferences and agreements to limit sulphur in fuels and $\text{NO}_x$ in car exhaust.
- Alternate, less polluting energy sources being used. (geothermal, solar, wind etc.)
- Industrial processes are being modernized (eg. new process for pulp mills involving $\text{H}_2\text{O}_2$ instead of sulphites.)
- Devices to remove gases like $\text{SO}_3$ from smoke stacks. (scrubbers). $\text{SO}_3 + \text{CaO} \rightarrow \text{CaSO}_4(s)$. The $\text{CaSO}_4$ is removed by electrostatic precipitation.
- Recycling

Review of Titrations

The “heart” of titrations is (moles in the middle):

Moles of standard $\xrightarrow{\text{Coefficient ratio}}$ Moles of sample

The moles of standard can be calculated by knowing the concentration of the standard and by measuring the volume used in the titration. Then the equation: $\text{mol} = M \times L$ can then be used.

The coefficient ratio in the balanced equation can then be used to calculate the moles of sample in the flask.

Knowing the volume of the sample and using the moles from the last step, one can then calculate the concentration of the sample: $M = \text{mol}/L$
**Equivalence Point (Stoichiometric Point)**

The point at where the actual mole ratio of Sample/Standard is the same as the coefficient ratio in the balanced equation.

Eg. Using the following reaction for a titration:

\[
\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}
\]

At equivalence point: moles of NaOH / moles of HCl = 1/1 (or mol NaOH = mol HCl)

Using the following reaction for a titration:

\[
2\text{HCl} + \text{Ba(OH)}_2 \rightarrow 2\text{H}_2\text{O} + \text{BaCl}_2
\]

At equivalence point: moles of Ba(OH)_2 / moles of HCl = 1/2 (or mol HCl = 2 x mol Ba(OH)_2)

In the most common type of titration question, we calculate: moles of standard → moles of sample → concentration of sample

Here’s an example:

A solution of HCl of unknown concentration was titrated with 0.150 M Ba(OH)_2. The equivalence point is reached when 14.83 mL of Ba(OH)_2 is added to 50.00 mL of the HCl solution. Find the [HCl] in the original sample.

(> The standard solution is the one of known concentration > in this case the 0.150 M Ba(OH)_2.)

1. Moles of Ba(OH)_2 = 0.150 M x 0.01483 L = 0.0022245 mol Ba(OH)_2

2. Moles of HCl:
   Using the balanced equation: \(2\text{HCl} + \text{Ba(OH)}_2 \rightarrow 2\text{H}_2\text{O} + \text{BaCl}_2\)

   \[
   \frac{0.0022245 \text{ mol Ba(OH)}_2 \times 2 \text{ mol HCl}}{1 \text{ mol Ba(OH)}_2} = 0.004449 \text{ mol HCl}
   \]

3. \([\text{HCl}] = \frac{0.004449 \text{ mol HCl}}{0.05000 \text{ L HCl}} = 0.0890 \text{ M}\)

Don’t forget, if a series of volume readings for different “Trials” are given, you may have to discard a reading that is more than 0.02 or so mL different from the rest of them. This ability to discard “far off” volume readings and then to calculate the “best average” volume will be tested!

Also markers always seem to use titration questions to test your ability to handle Significant Digits. Remember, when subtracting (or adding) use decimal places and when multiplying or dividing, use # of SD’s!
Here’s a question for you to try:

0.200 M NaOH is used to titrated 3 separate 50.0 mL samples of a solution of H₂SO₄ of unknown concentration. The NaOH is in the burette. Use the following data table to calculate the [H₂SO₄] in the original H₂SO₄ solution. Show all of your steps clearly, including the balanced formula equation for the reaction.

<table>
<thead>
<tr>
<th>Trial</th>
<th>Initial Burette Reading (mL)</th>
<th>Final Burette Reading (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Trial 1</td>
<td>0.00</td>
<td>9.02</td>
</tr>
<tr>
<td>Trial 2</td>
<td>9.02</td>
<td>17.95</td>
</tr>
<tr>
<td>Trial 3</td>
<td>17.95</td>
<td>26.89</td>
</tr>
</tbody>
</table>

Read “Experimental Note” on the bottom of page 157 of SW.
Do Ex. 94 – 97 on p. 158 of SW.

NOTE: If you check your answers on p. 291 of SW, you’ll notice, he uses “mmol’s” (millimoles) in his work. See p. 154 – 155 of doing these using mmols. You don’t have to use mmols. You can if you like. Just be careful with your units.

Do Worksheet 4-6

**Indicators**

- See demonstration of some indicators and their colours in acid and base solutions.

Acid-Base Indicators consist of equilibrium mixtures of:

\[
\text{A weak acid} \quad \text{and it’s conjugate base}
\]

\[
\text{HInd} \quad \text{Ind}^-
\]

which are:

Called the **acid form** of the indicator.

Called the **base form** of the indicator.

Different Colours
Eg.) An indicator HInd has a yellow acid form (HInd) and a red base form (Ind').

The equilibrium equation representing this indicator is:

\[
\text{HInd (yellow)} \quad + \quad \text{H}_2\text{O} \quad \Leftrightarrow \quad \text{H}_3\text{O}^+ \quad + \quad \text{Ind'} (\text{red})
\]

If reactants are favoured (equilibrium shifts to the LEFT), then: \([\text{HInd}] > [\text{Ind'}]\)

So: yellow is much greater than red.

And the solution will be **YELLOW**

If products are favoured (equilibrium shifts to the RIGHT), then: \([\text{Ind'}] > [\text{HInd}]\)

So: red is much greater than yellow.

And the solution will be **RED**

If there are equal amounts of reactant and products (equilibrium favours neither reactants nor products), then:

\([\text{HInd}] = [\text{Ind'}]\)

So: there is an equal mixture of yellow and red.

And the solution will be **ORANGE**

Make sure you read the material above a few times and make sure you understand how the shifts affect the colours of the solution in each case. This understanding is VERY important in dealing with indicators!

**Finding Colours of Acid and Base forms of indicators experimentally**

Looking at the equilibrium equation representing any indicator:

\[
\text{HInd} \quad + \quad \text{H}_2\text{O} \quad \Leftrightarrow \quad \text{H}_3\text{O}^+ \quad + \quad \text{Ind'}
\]

Adding any strong acid (eg. 1 M HCl) to an indicator (mixture of HInd and Ind') will:

- Increase the \([\text{H}_3\text{O}^+]\)
- Cause the equilibrium to shift to the LEFT
- Make \([\text{HInd}] > [\text{Ind'}]\)
- Cause the solution to turn the colour of HInd.

Go over those last 4 points while looking at the equilibrium equation and make sure you understand them. (Remember Le Chatelier’s Principle!)

So when HCl is added to an indicator (mixture of HInd and Ind'), it increases \([\text{H}_3\text{O}^+]\) causing the equilibrium:

\[
\text{HInd} \quad + \quad \text{H}_2\text{O} \quad \Leftrightarrow \quad \text{H}_3\text{O}^+ \quad + \quad \text{Ind'}
\]

To shift LEFT so \([\text{HInd}] > [\text{Ind'}]\) and the colour of HInd predominates.

In summary: If you add any strong acid (eg. HCl) to an indicator, it will turn the colour of the ACID FORM (HInd).
What about the Base Form (Ind\textsuperscript{−})?

Looking at the equilibrium equation representing any indicator:

\[
\text{HInd} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Ind}^-
\]

Adding any strong base (e.g., 1 M NaOH) to an indicator (mixture of HInd and Ind\textsuperscript{−}) will:

- Decrease the [\text{H}_3\text{O}^+] (because the OH\textsuperscript{−} neutralizes H\text{O}^+ to form H\text{O}₂: \text{H}_3\text{O}^+ + \text{OH}^− \rightarrow \text{H}_2\text{O})
- Cause the equilibrium to shift to the RIGHT
- Make [\text{Ind}^−] > [\text{HInd}]
- Cause the solution to turn the colour of Ind\textsuperscript{−}.

So when NaOH is added to an indicator (mixture of HInd and Ind\textsuperscript{−}), it decreases [\text{H}_3\text{O}^+] causing the equilibrium:

\[
\text{HInd} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Ind}^-
\]

Shifting Right

To shift RIGHT so [\text{Ind}^−] > [\text{HInd}] and the colour of Ind\textsuperscript{−} predominates

In summary: If you add any strong base (e.g., NaOH) to an indicator, it will turn the color of the BASE FORM (Ind\textsuperscript{−}).

**Question**: When a drop of 0.1M HCl is added to the indicator bromcresol green, the colour is yellow. When a drop of 0.10M NaOH is added to the indicator, the colour is blue.

a. What colour is the acid form of bromcresol green (HInd)? _______________________

b. What colour is the base form of bromcresol green (Ind\textsuperscript{−})? _______________________

c. What would the colour be if [HInd] = [Ind\textsuperscript{−}] for bromcresol green? _______________________

**Transition Point**

The **Transition Point** for an indicator is reached when [\text{HInd}] = [\text{Ind}^−]

This is where you have equal amounts of the colour of HInd and the colour of Ind\textsuperscript{−}.

(See the next page…)
Eg.) Looking on the “Acid-Base Indicators” Table (on the back of your Acid Table):

The two colours on the right side of the table for each indicator lists the colour of the ACID FORM first and then the colour of the BASE FORM.

So the Acid Form of methyl violet (HInd) is YELLOW and the Base Form of methyl violet (Ind⁻) is BLUE.

The colour at the TRANSITION POINT of

Methyl violet would be _______________

The colour at the TRANSITION POINT of

Bromcresol green would be ______________

The colour at the TRANSITION POINT of

Indigo carmine would be ________________

**Transition Point and Ka of Indicator**

The equilibrium equation for an indicator (HInd) is, as you know:

\[
\text{HInd} \quad + \quad \text{H}_2\text{O} \quad \rightleftharpoons \quad \text{H}_3\text{O}^+ \quad + \quad \text{Ind}^- 
\]

So the acid form (HInd) can be thought of as a weak acid. And weak acids, as you know, have a “Ka”.

The Ka expression for the weak acid HInd would be:

\[
\text{Ka} = \frac{[\text{H}_3\text{O}^+][\text{Ind}^-]}{[\text{HInd}]} 
\]

If we look back to the definition of transition point:
The Transition Point for an indicator is reached when \([\text{HInd}] = [\text{Ind}^-]\\)

Since \([\text{HInd}] = [\text{Ind}^-]\) at the transition point, we can cancel them out in the Ka expression:

So AT THE TRANSITION POINT: \(\text{Ka} = \frac{[\text{H}_3\text{O}^+][\text{Ind}^-]}{[\text{HInd}]}\\)

or AT THE TRANSITION POINT: \(\text{Ka} = [\text{H}_3\text{O}^+]\\)

Now, I’d like to throw in another definition. It’s a quantity called “pKa”:

\[
p\text{Ka} = -\log \text{Ka} 
\]
Going back to the Ka at transition point:

**AT THE TRANSITION POINT:** \( Ka = [H_3O^+] \)

We take the \(-\log\) of both sides: \(-\log Ka = -\log [H_3O^+]\) or as we know from definitions:

**AT THE TRANSITION POINT:** \( pKa = pH \)

Remember, this is **ONLY true at the TRANSITION POINT** of the indicator. (When \([\text{HInd}] = [\text{Ind}^-]\) and the two colours are equal)

So now, we can summarize **FOUR** things we know to be true at the Transition Point:

**AT THE TRANSITION POINT**

- \([\text{HInd}] = [\text{Ind}^-]\)
- \( Ka \) (indicator) = \([H_3O^+]\)
- \( pKa = pH \)
- The colour is a 50/50 mixture of the acid and base colours

You need to remember all that really well to understand the next material!

**Transition Range and Transition Point**

If you look at the Indicator Table on the back of the Acid Table, there is a column entitled “pH Range in which Colour Change Occurs”.

As the pH is gradually raised, the colour does not instantaneously change from acid colour to base colour. There is a gradual change over a range of pH’s.

For example:

<table>
<thead>
<tr>
<th>Indicator</th>
<th>pH Range in Which Colour Change Occurs</th>
<th>Colour Change as pH Increases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methyl violet</td>
<td>0.0 – 1.6</td>
<td>yellow to blue</td>
</tr>
<tr>
<td>Thymol blue</td>
<td>1.2 – 2.8</td>
<td>red to yellow</td>
</tr>
<tr>
<td>Orange IV</td>
<td>1.4 – 2.3</td>
<td>red to yellow</td>
</tr>
</tbody>
</table>

It says that Methyl Violet gradually changes from yellow to blue in the pH range of 0.0 – 1.6.

This means when pH is at or below 0.0, the colour of methyl violet is yellow.

When pH is 1.6 or above, the colour of methyl violet is blue. But what about between?

Between pH of 0.0 and 1.6, there is a mixture of the yellow and the blue form of methyl violet, so the colour is GREEN. We can refine it even further by saying that between pH of 0.0 and 0.8, the colour is more of a yellow green and between pH 0.8 and 1.6, it is more of a blue green. At a pH of 0.8 (half way between 0.0 and 1.6), the colour would be simply Green! See if you can fill these colours in the table on the next page...
Finding the Ka of an Indicator

To find the pH at Transition Point:
1. Look on the Indicator table
2. Find the midpoint of the pH range by adding the two numbers and dividing by two.

Remember from the last page:
\[ \text{pKa} = \text{pH at the Transition Point.} \]

Since \( \text{pKa} = -\log \text{Ka} \)

**Ka = antilog (-pKa)**

Let’s do an example. Find the Ka of Phenol Red:

<table>
<thead>
<tr>
<th>Indicator</th>
<th>pH Range in Which Colour Change Occurs</th>
<th>Colour Change as pH Increases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Phenol red</td>
<td>6.6 - 8.0</td>
<td>yellow to red</td>
</tr>
</tbody>
</table>

The pH at the Transition Point is: \[ \frac{6.6 + 8.0}{2} = 7.3 \]

Since pH at TP = pKa, then pKa = 7.3

\[ \text{Ka} = \text{antilog (-7.3)} = 5 \times 10^{-8} \]

Find the Ka of Alizarin Yellow:

**Thymol Blue (A diprotic Indicator)**

You’ll notice that Thymol Blue appears twice on the Indicator Table:

<table>
<thead>
<tr>
<th>Indicator</th>
<th>pH Range in Which Colour Change Occurs</th>
<th>Colour Change as pH Increases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Thymol blue</td>
<td>1.2 - 2.8</td>
<td>red to yellow</td>
</tr>
<tr>
<td>Thymol blue</td>
<td>8.0 - 9.6</td>
<td>yellow to blue</td>
</tr>
</tbody>
</table>

pH’s are only 1 decimal place which means 1 SD
This is because Thymol Blue is a diprotic acid. Each time it loses a proton, it goes through a color change.

We can call Thymol Blue (Tb) a weak acid $H_2Tb$

The equilibrium equation for the first ionization is: $H_2Tb + H_2O \rightleftharpoons H_3O^+ + HTb^-$

Using the table above, fill in the colours: __________  __________

The equilibrium equation for the second ionization is: $HTb^- + H_2O \rightleftharpoons H_3O^+ + Tb^{2-}$

Using the table below, fill in the colours: __________  __________

Looking at the pH ranges above, try to fill in the following information:

<table>
<thead>
<tr>
<th>pH</th>
<th>Form(s) which predominate(s) (H$_2$Tb, HTb$^-$ or Tb$^{2-}$)</th>
<th>Approximate Colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2.0</td>
<td>&amp; are equal</td>
<td></td>
</tr>
<tr>
<td>3.0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7.0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8.8</td>
<td>&amp; are equal</td>
<td></td>
</tr>
<tr>
<td>10.0</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Also, fill in the colours on the following diagram:

Colours of Thymol Blue:

pH       < 1.2 | 2.8 - 8.0 | > 9.6

Find the pH’s and the colours of the given indicators in the following solutions (assume temp. = 25°C):

<table>
<thead>
<tr>
<th>Solution</th>
<th>pH</th>
<th>Colour in Thymol Blue</th>
<th>Colour in Methyl Red</th>
<th>Colour in Alizarin Yellow</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.2 M HCl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.01 M HCl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.0005 M HCl</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Pure water</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.0001 M NaOH</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>0.2 M NaOH</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
A variety of indicators can also be used to narrow the known pH range for a solution and help identify the solution:

For example, a solution displays the following colours in the indicators shown. See if you can narrow the pH down to a range:

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Colour of Solution</th>
<th>Approximate pH Range</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bromthymol blue</td>
<td>Blue</td>
<td></td>
</tr>
<tr>
<td>Thymol blue</td>
<td>Yellow</td>
<td></td>
</tr>
<tr>
<td>Phenolphthalein</td>
<td>Colourless</td>
<td></td>
</tr>
</tbody>
</table>

Approximate pH range of the solution using all information:

Let’s try another one:

For example, a solution displays the following colours in the indicators shown. See if you can narrow the pH down to a range:

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Colour of Solution</th>
<th>Approximate pH Range</th>
</tr>
</thead>
<tbody>
<tr>
<td>Orange IV</td>
<td>Yellow</td>
<td></td>
</tr>
<tr>
<td>Methyl red</td>
<td>Red</td>
<td></td>
</tr>
<tr>
<td>Methyl Orange</td>
<td>Red</td>
<td></td>
</tr>
</tbody>
</table>

Approximate pH range of the solution using all information:

For example, a solution displays the following colours in the indicators shown. See if you can narrow the pH down to a range:

<table>
<thead>
<tr>
<th>Indicator</th>
<th>Colour of Solution</th>
<th>Approximate pH Range</th>
</tr>
</thead>
<tbody>
<tr>
<td>Methyl Orange</td>
<td>Yellow</td>
<td></td>
</tr>
<tr>
<td>Alizarin Yellow</td>
<td>Yellow</td>
<td></td>
</tr>
<tr>
<td>Thymol Blue</td>
<td>Green</td>
<td></td>
</tr>
</tbody>
</table>

Approximate pH range of the solution using all information:

Universal Indicators – Give a variety of colours over a larger pH range

If several indicators are mixed, the combinations of colours can lead to many different colours as we move from one pH to another. Study the 3 tables given on page 162 of SW to give you an idea of how universal indicators can be made. The second table is somewhat simplified as it does not include the colours of indicators in their transition ranges. The third table is more precise.

Read p. 159 – 162 in SW.
Do Ex. 108 – 112 and 114 – 120 on p. 162 – 163 of SW
Using Indicators to Rank Weak Acids in Order of Strengths

To understand this section, recall that equilibrium always favours the side with the WEAKER acid (or weaker base).

Let’s say an indicator HInd is Red in 0.1M HCl and Blue in 0.1M NaOH.

Give the equilibrium equation for this indicator and write the colour of each form (HInd) and (Ind⁻) underneath it:

A few drops of this indicator (a mixture of HInd and Ind⁻) is added to a weak acid called HA₁ and the colour is Blue.

Which is the stronger acid, HA₁ or HInd?
To find out, we write an equilibrium equation (NOT with H₂O this time!). For reactants, we use the weak acid HA₁ and the base form of the indicator, Ind⁻. (two acids are not written on the same side of equilibrium equations!)
I’m sure you can fill in the two products: Write the colours of Ind⁻ and HInd right underneath each one.

\[
\text{HA}_1 + \text{Ind}^- \rightleftharpoons +
\]

Since the colour of the indicator was Blue, it means that the form of the indicator (HInd or Ind⁻) _______ is predominating (favoured by the equilibrium). So the (reactants/products) ____________________________ of the equation above are favoured, meaning \((\text{HA}_1/\text{HInd})_______________________\) is the Weaker acid or \((\text{HA}_1/\text{HInd})_______________________\) is the Stronger Acid.

Now let’s look at another experiment involving the same indicator and a different weak acid HA₂.

A few drops of this indicator (a mixture of HInd and Ind⁻) is added to a weak acid called HA₂ and the colour is Red.

Which is the stronger acid, HA₂ or HInd?
To find out, we write an equilibrium equation (NOT with H₂O this time!). For reactants, we use the weak acid HA₂ and the base form of the indicator, Ind⁻. (two acids are not written on the same side of equilibrium equations!)
I’m sure you can fill in the two products: Write the colours of Ind⁻ and HInd right underneath each one.

\[
\text{HA}_2 + \text{Ind}^- \rightleftharpoons +
\]

Since the colour of the indicator was Red, it means that the form of the indicator (HInd or Ind⁻) _______ is predominating (favoured by the equilibrium). So the (reactants/products) ____________________________ of the equation above are favoured, meaning \((\text{HA}_2/\text{HInd})_____________________\) is the Weaker acid or \((\text{HA}_2/\text{HInd})_____________________\) is the Stronger Acid.

So, to summarize the results of both experiments:

Experiment 1: __________ > __________

Experiment 2: __________ > __________

So, in comparing strengths of HA₁ and HA₂, we can say that __________ > __________
Now, make a little mini acid table with the acids on the left, a \( \rightleftharpoons \) in the middle and \( \text{H}^+ + \) conj. base on the right.

Put the acids in order of strongest \( \rightarrow \) weakest.

<table>
<thead>
<tr>
<th>Acid</th>
<th>Base</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>( \rightleftharpoons \text{H}^+ + )</td>
</tr>
<tr>
<td></td>
<td>( \rightleftharpoons \text{H}^+ + )</td>
</tr>
<tr>
<td></td>
<td>( \rightleftharpoons \text{H}^+ + )</td>
</tr>
</tbody>
</table>

Do Worksheet 4 - 7—Indicators