**Student Activity 2: pH Calculations**

**Section 1: Setting Up the Dissociation Reaction**

Strong acids: HCl, HBr, HI, HNO₃, HClO₄, H₂SO₄ --> all other acids are WEAK

Strong bases: Arrhenius Bases: LiOH, NaOH, KOH, RbOH, Ca(OH)₂ etc. (have OH⁻ in its structure) → all other bases are WEAK

*Make sure to pay attention to the reaction arrow.*

*Samples highlighted the species in the product that controls the pH. You should circle those in yours.*

<table>
<thead>
<tr>
<th>Category</th>
<th>Equation:</th>
<th>Conjugate Acid or Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ex: HNO₂</td>
<td>Weak Acid</td>
<td>Conjugate Base: NO₂⁻</td>
</tr>
<tr>
<td>Ex: NaOH</td>
<td>Strong Base</td>
<td>None</td>
</tr>
<tr>
<td>Ex: C₆H₅NH₂</td>
<td>Weak Base</td>
<td>Conj. Acid: C₆H₅NH₃⁺</td>
</tr>
<tr>
<td>Ex: HBr</td>
<td>Strong Acid</td>
<td>Technically: Br⁻, albeit very weak CB</td>
</tr>
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</table>

KOH

(CH₃)₂NH

HClO₄

HClO

C₆H₅COOH
Section 2: Calculating the pH of Strong/Weak Acids/Bases
You will need your Ka/Kb reference sheet here!

Ex: What is the pH of 0.01M HCl?
Strong acid so just take the -log[H⁺] = -log[0.01] = 2

Ex: What is the pH of a 0.001M NaOH solution?
Strong base. It’s a base so we have to work through pOH =-log[OH⁻]=-log[0.001] = 3
pH=14-3= 11

Ex: What is the pH of a 0.5M HNO₂?
Weak acid so must use Ka to solve because the concentration of the H⁺ does NOT equal the concentration of the HNO₂.

Reaction:  HNO₂ + H₂O ⇌ H₃O⁺ + NO₂⁻

Ka=\[\text{H}_3\text{O}^+]\[\text{NO}_2^-\]
[\text{HNO}_2]

<table>
<thead>
<tr>
<th></th>
<th>HNO₂</th>
<th>H₃O⁺</th>
<th>NO₂⁻</th>
</tr>
</thead>
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<tr>
<td>I (initial)</td>
<td>0.5 from the problem</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>C (change)</td>
<td>-x</td>
<td>+x</td>
<td>+x</td>
</tr>
<tr>
<td>E (equilibrium)</td>
<td>0.5-x ~ 0.5</td>
<td>x</td>
<td>x</td>
</tr>
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</table>

Plug the equilibrium values into the Ka equation.
Ka value = 7.2 x 10⁻⁴

7.2 x 10⁻⁴ = \[\frac{x}{0.5}\]

X² = 3.6 x 10⁻⁴ (take square root)
X=0.0190M = [H₃O⁺]

NOW you can plug it into pH=-log[H₃O⁺] = -log[0.019] = 1.72 ◼️ answer for the pH
Ex: What is the pH of a 0.2M C₂H₅NH₂ solution?
This is a weak base. You can’t just use the pOH equation.
This is the SAME methodology as the weak acid above EXCEPT that the base will be a Kb
and you will solve for a pOH first.

Reaction: C₂H₅NH₂ + H₂O ⇌ C₂H₅NH₃⁺ + OH⁻

We have a Kb because this is a base.
Kb = \frac{[C₂H₅NH₃⁺][OH⁻]}{[C₂H₅NH₂]}

<table>
<thead>
<tr>
<th></th>
<th>C₂H₅NH₂</th>
<th>OH⁻</th>
<th>C₂H₅NH₃⁺</th>
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<td>+x</td>
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</tr>
<tr>
<td>E (equilibrium)</td>
<td>0.2-x ~ 0.2 Remember you can approximate away the – x because it is soooo small</td>
<td>x this is the one that really matters because it is the one that determines the pH</td>
<td>x</td>
</tr>
</tbody>
</table>

Plug the equilibrium values into the Kb expression.
I looked the Kb value up on the BACK of the pink reference sheet = 4.38x10⁻⁴

4.38 x 10⁻⁴ = \frac{[x][x]}{0.2}

(multiply and then take the square root to solve for x)
X=0.00936M=[OH⁻]

Now take the pOH=-\log[OH⁻]=-\log[0.00936]=2.03
To get to pH (always with a base)... pH=14-pOH = 14-2.02 = 11.97  final pH answer

NOW YOU TRY:
1. What is the pH of a 0.025M KOH solution?
2. What is the pH of a 0.15M C₆H₅COOH solution (the Ka is hard to read for this one... it’s 6.3 x 10⁻⁵)?

3. What is the pH of a 0.3M NH₃ solution?

4. What is the pH of 0.08M HI solution?