

Acids and Bases – Formula Sheet:

<p>Arrhenius Definition:</p> <ul style="list-style-type: none"> • Acids produce H^+ ions in solutions. • Bases produce OH^- ions in solutions. <p>Bronsted-Lowry Definition:</p> <ul style="list-style-type: none"> • Acids are proton donors. • Bases are proton acceptors. <p>Lewis Definition:</p> <ul style="list-style-type: none"> • Acids are electron pair acceptors. • Bases are electron pair donors. 	<p>Acid-Base Equations: (0.1 M HCl or 0.15M KOH)</p> $pH = -\log[H^+] \quad pOH = -\log[OH^-]$ $pH + pOH = 14$ $[H^+] = 10^{-pH} \quad [OH^-] = 10^{-pOH}$ $[H^+][OH^-] = 1 \times 10^{-14}$ $[H^+] = [H_3O^+]$																					
<p>Strong Acids: HCl / HBr / HI / HNO₃ / HClO₄ / H₂SO₄</p> <p>Weak Acids: HF / HNO₂ / HClO / HCN / HC₂H₃O₂</p> <p>Strong Bases: NaOH / KOH Weak Bases: NH₃</p>	<p>Autoionization of Water:</p> $H_2O_{(l)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$ $K_w = [H^+][OH^-]$ $K_w = 1 \times 10^{-14} \text{ at } 25^\circ C$																					
<p>pH of a Weak Base: (0.25 M NH₃)</p> $B_{(aq)} + H_2O_{(l)} \rightarrow HB^+_{(aq)} + OH^-_{(aq)}$ $K_b = \frac{[HB^+][OH^-]}{[B]}$ <p>If $K_b < 1 \times 10^{-4}$, then $\rightarrow [OH^-] \approx \sqrt{[B] \cdot K_b}$</p> $pOH = -\log[OH^-] \quad pH = 14 - pOH$ $pH = \frac{1}{2}(14 + pK_a + \log[B])$	<p>pH of a Weak Acid: (0.5M HC₂H₃O₂)</p> $HA_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + A^-_{(aq)}$ $K_a = \frac{[H_3O^+][A^-]}{[HA]}$ <p>If $K_a < 1 \times 10^{-4}$, then $\rightarrow [H^+] \approx \sqrt{[HA] \cdot K_a}$</p> $pH = -\log[H^+]$ $pH = \frac{1}{2}(pK_a - \log[HA])$																					
<table border="1" style="width: 100%; border-collapse: collapse; text-align: center;"> <thead> <tr> <th style="color: red;">Acidic Ions:</th> <th style="color: blue;">Neutral Ions:</th> <th style="color: blue;">Weak Basic Ions:</th> </tr> </thead> <tbody> <tr> <td>NH_4^+</td> <td>Cl^-</td> <td>F^-</td> </tr> <tr> <td>Al^{3+}</td> <td>Br^-</td> <td>CN^-</td> </tr> <tr> <td>Fe^{3+}</td> <td>I^-</td> <td>$C_2H_3O_2^-$</td> </tr> <tr> <td>Cu^{2+}</td> <td>NO_3^-</td> <td>NO_2^-</td> </tr> <tr> <td></td> <td>ClO_4^-</td> <td>ClO^-</td> </tr> <tr> <td></td> <td>HSO_4^-</td> <td>CO_3^{2-}</td> </tr> </tbody> </table> <p>Strong Base Ions: OH^- O^{2-} H^- NH_2^-</p>	Acidic Ions:	Neutral Ions:	Weak Basic Ions:	NH_4^+	Cl^-	F^-	Al^{3+}	Br^-	CN^-	Fe^{3+}	I^-	$C_2H_3O_2^-$	Cu^{2+}	NO_3^-	NO_2^-		ClO_4^-	ClO^-		HSO_4^-	CO_3^{2-}	<p>Percent Ionization for Acids:</p> $\% \text{ Ionization} = \frac{[H^+]}{[HA]} \times 100\%$ <p>Percent Ionization for Bases:</p> $\% \text{ Ionization} = \frac{[OH^-]}{[B]} \times 100\%$
Acidic Ions:	Neutral Ions:	Weak Basic Ions:																				
NH_4^+	Cl^-	F^-																				
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	ClO_4^-	ClO^-																				
	HSO_4^-	CO_3^{2-}																				

Acid:	K _a Value:	Acid-Base Dissociation Constant Equations:							
<i>HNO₂</i>	4.0×10^{-4}	$pK_a = -\log K_a \quad pK_b = -\log K_b$							
<i>HF</i>	7.2×10^{-4}	$pK_a + pK_b = 14$							
<i>HC₂H₃O₂</i>	1.8×10^{-5}	$K_a = 10^{-pK_a} \quad K_b = 10^{-pK_b}$							
<i>HClO</i>	3.5×10^{-8}	$K_a \times K_b = 1 \times 10^{-14}$							
<i>H₂PO₄⁻</i>	6.2×10^{-8}								
<i>NH₄⁺</i>	5.6×10^{-10}								
<i>HCN</i>	6.2×10^{-10}								
<p>Note: A buffer solution is made up of a weak acid and its conjugate weak base. Buffer solutions resist changes to its pH.</p>		<p>pH - Buffer Solution: (0.5M NH₄Cl / 0.4M NH₃) Henderson-Hasselbalch Equation:</p>							
<p>Examples of Buffer Solutions:</p>		$pH = pK_a + \log \left(\frac{[A^-]}{[HA]} \right)$							
<ol style="list-style-type: none"> 1. HF / NaF 2. NH₄Cl / NH₃ 3. HC₂H₃O₂ / NaC₂H₃O₂ 		$\frac{[A^-]}{[HA]} = 10^{pH-pK_a}$							
$pH > pK_a \text{ when } [A^-] > [HA]$ $pH < pK_a \text{ when } [A^-] < [HA]$		<p>Note: $pH = pK_a$ when $[A^-] = [HA]$</p>							
<p>Dissociation Constants for H₃PO₄</p>		<p>pH of a Polyprotic Acid: (0.25M H₃PO₄)</p>							
<table border="1"> <tr> <td><i>H₃PO₄</i></td> <td>$K_{a1} = 7.5 \times 10^{-3}$</td> </tr> <tr> <td><i>H₂PO₄⁻</i></td> <td>$K_{a2} = 6.2 \times 10^{-8}$</td> </tr> <tr> <td><i>HPO₄²⁻</i></td> <td>$K_{a3} = 4.8 \times 10^{-13}$</td> </tr> </table>	<i>H₃PO₄</i>	$K_{a1} = 7.5 \times 10^{-3}$	<i>H₂PO₄⁻</i>	$K_{a2} = 6.2 \times 10^{-8}$	<i>HPO₄²⁻</i>	$K_{a3} = 4.8 \times 10^{-13}$		$H_3A_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + H_2A^-_{(aq)}$	
<i>H₃PO₄</i>	$K_{a1} = 7.5 \times 10^{-3}$								
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<i>HPO₄²⁻</i>	$K_{a3} = 4.8 \times 10^{-13}$								
<p>Note: The 1st step is most important for calculating the pH of the solution:</p>		$K_{a1} = \frac{[H_3O^+][H_2A^-]}{[H_3A]}$							
$[H_3O^+] = [H_2A^-]$		$pH = -\log[H_3O^+]$							
$K_{a2} = [HA^{2-}]$		<hr/> $H_2A^-_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HA^{2-}_{(aq)}$							
<p>Amphoteric Ion Reactions in Water:</p>		<p>pH of an Amphoteric Salt: (0.4M NaH₂PO₄)</p>							
$H_2A^-_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + HA^{2-}_{(aq)} \quad K_{a2}$		$pH \approx \frac{1}{2}(pK_{a1} + pK_{a2}) \text{ "Isoelectric Point"}$							
$H_2A^-_{(aq)} + H_2O_{(l)} \rightarrow OH^-_{(aq)} + H_3A_{(aq)} \quad K_{b3}$		$[H^+] \approx \sqrt{K_{a1} \cdot K_{a2}} \quad K_{b3} = \frac{K_w}{K_{a1}}$							
$K_{a2} = \frac{[H_3O^+][HA^{2-}]}{[H_2A^-]} \quad K_{b3} = \frac{[H_3A][OH^-]}{[H_2A^-]}$		$[H^+] = \sqrt{K_{a1} \cdot K_{a2} \cdot \frac{[H_3A]}{[H_2A^-]}}$							

Standard Form of a Quadratic Equation:

$$ax^2 + bx + c = 0$$

The Quadratic Formula:

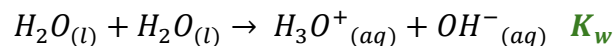
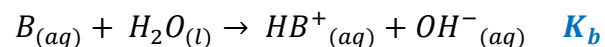
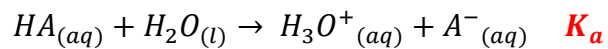
$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Dilution Formula:

$$M_1V_1 = M_2V_2$$

Moles:

$$n = MV$$

pH of a Weak Acid / Weak Base Salt: (0.2M NH₄F)

$$[H^+] \approx \sqrt{\frac{K_a K_w}{K_b}} \quad \text{If } [B] \approx [HA]$$

$$[H^+] = \sqrt{\frac{K_a K_w}{K_b} \cdot \frac{[B]}{[HA]} \cdot \frac{[HB^+]}}{[A^-]}}$$

Titration:	pH at Equiv. point
Strong Acid – Strong Base	pH = 7
Weak Acid – Strong Base	pH > 7
Weak Base – Strong Acid	pH < 7

Acid-Base Titrations:

- ICE Tables – Use Molarity
- BCA Tables – Use Moles

At ½ Veq (Equivalence Volume):

$$pH = pK_a \quad \text{and} \quad [A^-] = [HA]$$

Acid-Base Indicators:

Indicator:	K_a	pK_a	$HIn \text{ to } In^-$
Methyl Orange	3.4×10^{-4}	3.5	Red to Yellow
Methyl Red	7.9×10^{-6}	5.1	Red to Yellow
Bromthymol Blue	1.0×10^{-7}	7.0	Yellow to Blue
Phenolphthalein	5.0×10^{-10}	9.3	Clear to Pink