Acids and Bases – Formula Sheet:

Arrhenius Definition:

- Acids produce H^+ ions in solutions.
- Bases produce OH^- ions in solutions.

Bronsted-Lowry Definition:

- Acids are proton donors.
- Bases are proton acceptors.

Lewis Definition:

- Acids are electron pair acceptors.
- Bases are electron pair donors.

Acid-Base Equations: (0.1 M HCl or 0.15M KOH)

$$pH = -\log[H^+] \qquad pOH = -\log[OH^-]$$

$$pH + pOH = 14$$

$$[H^+] = 10^{-pH} \qquad [OH^-] = 10^{-pOH}$$

$$[H^+][OH^-] = 1 \, x \, 10^{-14}$$

$$[H^+] = [H_3O^+]$$

Strong Acids:
HCl / HBr / HI / HNO_3 / $HClO_4$ / H_2SO_4

Weak Acids:
HF / HNO_2 / $HClO$ / HCN / $HC_2H_3O_2$

Strong Bases: $NaOH$ / KOH **Weak Bases:** NH_3

Autoionization of Water:

$$H_2O_{(l)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + OH^-{}_{(aq)}$$

$$\boldsymbol{K_w} = [H^+][OH^-]$$

$$\boldsymbol{K_w} = 1 \, x \, 10^{-14} \; at \; 25\,°C$$

pH of a Weak Base: (0.25 M NH_3)

$$B_{(aq)} + H_2O_{(l)} \rightarrow HB^+{}_{(aq)} + OH^-{}_{(aq)}$$

$$\boldsymbol{K_b} = \frac{[HB^+]\,[OH^-]}{[B]}$$

If $K_b < 1 \times 10^{-4}$, then → $[OH^-] \approx \sqrt{[B] \cdot K_b}$

$$pOH = -\log[OH^-] \qquad pH = 14 - pOH$$

$$pH = \frac{1}{2}(14 + pK_a + \log[B])$$

pH of a Weak Acid: (0.5M $HC_2H_3O_2$)

$$HA_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + A^-{}_{(aq)}$$

$$\boldsymbol{K_a} = \frac{[H_3O^+]\,[A^-]}{[HA]}$$

If $K_a < 1 \times 10^{-4}$, then → $[H^+] \approx \sqrt{[HA] \cdot K_a}$

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

$$pH = \frac{1}{2}(pK_a - \log[HA])$$

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^-
	$HSO_4{}^-$	$CO_3{}^{2-}$

Strong Base Ions: OH^- O^{2-} H^- $NH_2{}^-$

Percent Ionization for Acids:

$$\%\ Ionization = \frac{[H^+]}{[HA]} \times 100\%$$

Percent Ionization for Bases:

$$\%\ Ionization = \frac{[OH^-]}{[B]} \times 100\%$$

Acid:	K_a Value:
HNO_2	4.0×10^{-4}
HF	7.2×10^{-4}
$HC_2H_3O_2$	1.8×10^{-5}
$HClO$	3.5×10^{-8}
$H_2PO_4^{\ -}$	6.2×10^{-8}
$NH_4^{\ +}$	5.6×10^{-10}
HCN	6.2×10^{-10}

Acid-Base Dissociation Constant Equations:

$$pK_a = -\log K_a \qquad pK_b = -\log K_b$$

$$pK_a + pK_b = 14$$

$$K_a = 10^{-pK_a} \qquad K_b = 10^{-pK_b}$$

$$K_a \times K_b = 1 \times 10^{-14}$$

Note: A buffer solution is made up of a weak acid and its conjugate weak base. Buffer solutions resist changes to its pH.

Examples of Buffer Solutions:

1. HF / NaF
2. NH_4Cl / NH_3
3. $HC_2H_3O_2$ / $NaC_2H_3O_2$

$$pH > pK_a \text{ when } [A^-] > [HA]$$
$$pH < pK_a \text{ when } [A^-] < [HA]$$

pH - Buffer Solution: (0.5M NH_4Cl / 0.4M NH_3)

Henderson-Hasselbalch Equation:

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

$$\frac{[A^-]}{[HA]} = 10^{pH-pKa}$$

Note: $pH = pK_a \text{ when } [A^-] = [HA]$

Dissociation Constants for H_3PO_4

H_3PO_4	$K_{a1} = 7.5 \times 10^{-3}$
$H_2PO_4^{\ -}$	$K_{a2} = 6.2 \times 10^{-8}$
$HPO_4^{\ 2-}$	$K_{a3} = 4.8 \times 10^{-13}$

Note: The 1st step is most important for calculating the pH of the solution:

$$[H_3O^+] = [H_2A^-]$$

$$K_{a2} = [HA^{2-}]$$

pH of a Polyprotic Acid: (0.25M H_3PO_4)

$$H_3A_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + H_2A^-{}_{(aq)}$$

$$K_{a1} = \frac{[H_3O^+]\,[H_2A^-]}{[H_3A]}$$

$$pH = -\log[H_3O^+]$$

$$H_2A^-{}_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + HA^{2-}{}_{(aq)}$$

$$K_{a2} = \frac{[H_3O^+]\,[HA^{2-}]}{[H_2A^-]}$$

Amphoteric Ion Reactions in Water:

$$H_2A^-{}_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + HA^{2-}{}_{(aq)} \quad K_{a2}$$

$$H_2A^-{}_{(aq)} + H_2O_{(l)} \rightarrow OH^-{}_{(aq)} + H_3A_{(aq)} \quad K_{b3}$$

$$K_{a2} = \frac{[H_3O^+]\,[HA^{2-}]}{[H_2A^-]} \qquad K_{b3} = \frac{[H_3A]\,[OH^-]}{[H_2A^-]}$$

pH of an Amphoteric Salt: (0.4M NaH_2PO_4)

$$pH \approx \frac{1}{2}(pK_{a1} + pK_{a2}) \quad \text{"Isoelectric Point"}$$

$$[H^+] \approx \sqrt{K_{a1} \cdot K_{a2}} \qquad K_{b3} = \frac{K_w}{K_{a1}}$$

$$[H^+] = \sqrt{K_{a1} \cdot K_{a2} \cdot \frac{[H_3A]}{[H_2A^-]}}$$

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 NH_4F)

$$HA_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + A^-{}_{(aq)} \quad \boldsymbol{K_a}$$

$$B_{(aq)} + H_2O_{(l)} \rightarrow HB^+{}_{(aq)} + OH^-{}_{(aq)} \quad \boldsymbol{K_b}$$

$$H_2O_{(l)} + H_2O_{(l)} \rightarrow H_3O^+{}_{(aq)} + OH^-{}_{(aq)} \quad \boldsymbol{K_w}$$

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

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

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):

$$pH = pKa \quad and \quad [A^-] = [HA]$$

Acid-Base Indicators:

Indicator:	$\boldsymbol{K_a}$	$\boldsymbol{pK_a}$	$\boldsymbol{HIn\ to\ In^-}$
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***