AP Chemistry – Acid-Base Equilibrium and Buffers

Defining Acids/Bases

1. Arrhenius definition:
   Acids – donate H^+
   Bases – donate OH^-
2. Bronstead-Lowry definition
   Acids – donate H^+
   Bases – accept (gain) H^+

General pH Calculations

$[H^+]$ ⟷ $[H^+] \times [OH^-] = 1\text{E-}14$ ⟷ $[OH^-]$

$[H^+]$ ⟷ $-\log[H^+] = pH$ ⟷ pH

$[OH^-]$ ⟷ $-\log[OH^-] = pOH$ ⟷ pOH

pH ⟷ pH + pOH = 14 ⟷ pOH

Strong Acids

1. **HCl**
2. **HBr**
3. **HI**
4. **H_2SO_4**
5. **$HClO_3$**
6. **HNO_3**

Not one of these….
Then it is WEAK!

Conjugates

Conjugates are ions (usually) that are left after the acid donates or base accepts protons (they are the "anti")

Conjugate Examples

$HA \rightleftharpoons H^+ + A^-$
(HA = Acid; A^- = Conjugate Base)

$B + H_2O \rightleftharpoons HB^+ + OH^-$
(B = Base; HB^+ = Conjugate Acid)

Strong Bases

Group 1 hydroxides
Ex: NaOH or LiOH

Group 2 hydroxides
Ex: $Ca(OH)_2$

Trick alert – not all G2 hydroxides are soluble

Not one of these….
Then it is WEAK!

pH Calculations in Your Head

$[H^+]$	pH estimate	pH actual
1×10^{-5}	5	5
2×10^{-5}	4.75	4.7
3×10^{-5}	4.5	4.5
6×10^{-5}	4.25	4.2
10×10^{-5}	4	4

General formula

1×10^{-n}	$pH = n$
2×10^{-n}	$pH = n - 0.25$
3×10^{-n}	$pH = n - 0.5$
6×10^{-n}	$pH = n - 0.75$
10×10^{-n}	$pH = n - 1$

Calculating pH of WEAK ACIDS

To get pH, need equil $[H^+]$ (use QCK to get x!)

Example: Solve for pH of weak acid, HA ($K_a = 1.3\text{E-}9$)

	HA ⇌	$A^-_{(aq)}$ +	$H^+_{(aq)}$
Q	1.0 M	0	0
C	-x	x	x
K	1 - x	x	x

$K_a = \frac{[A^-][H^+]}{[HA]}$ ➡ $1.3\text{E-}9 = \frac{(x)(x)}{(1-x)}$

Solve for x (ignore minus x!)

$1.3\text{E-}9 = \frac{(x)(x)}{(1)}$

$X = 3.6\text{E-}5 = [H^+]$

$-\log[H^+] = pH$
$-\log 3.6\text{E-}5 = 4.4$

pH = 4.4

Calculating pH of WEAK BASES

Use QCK to get X (equil $[OH^-]$), then solve for pH

Example: Solve for pH of weak base, B ($K_b = 2.1\text{E-}10$)

	$B_{(aq)}$ +	$H_2O_{(l)}$ ⇌	$HB^+_{(aq)}$ +	$OH^-_{(aq)}$
Q	1.0 M		0	0
C	-x		x	x
K	1 - x		x	x

$K_b = \frac{[HB^+][OH^-]}{[B]}$ ➡ $2.1\text{E-}10 = \frac{(x)(x)}{(1-x)}$

Solve for x (ignore minus x!)

$2.1\text{E-}10 = \frac{(x)(x)}{(1)}$

$X = 1.5\text{E-}5 = [OH^-]$

$-\log[OH^-] = pOH$
$-\log 1.5\text{E-}5 = 4.8 = pOH$

CAREFUL… this is pOH!!!

pH + pOH = 14
pH = 14 – 4.8
pH = 9.2

Confusion Alert!

We do not use QCK tables for SA/SB calculations b/c they completely dissociate, no equilibrium so… $[ACID] = [H^+]$

Hydrolysis of Salts

Conjugates of weak acids/bases react with water to form acidic/basic solutions

Predicting Acidic/Basic

STRONG WINS!

Determine if ions produced from strong or weak

WB/SA = Acidic (strong wins!)

SB/WA = Basic (strong wins!)

Writing hydrolysis reactions:

Ignore the strong ion! The weak ion reacts with water…

Water is always a reactant!
- Bases have OH as a product
- Acids have H as a product

☐ + $H_2O \rightleftharpoons H_3O^+$ + ☐ (Acidic salt, produces H^+)

☐ + $H_2O \rightleftharpoons OH^-$ + ☐ (Basic salt, produces OH^-)

Hint: Ions and conjugates fill into boxes

Getting Moles From Molarity

$M = \frac{moles}{Liters}$

M x L = moles

See it another way:

$\frac{moles}{Liters} \times Liters = moles$

Buffers – Buffers are solutions that contain both an acid and its conjugate base (or a base and conjugate acid). They are unique because they can neutralize BOTH acids and bases added to the solution (so they resist changes in pH)

Making Buffers

- To make a buffer mix weak acid and its conjugate salt
  Ex: Acetic acid / Sodium acetate
- 1:1 molar ratios offer greatest buffering capacity (pH=pK_a)

$$[H^+] = K_a \left(\frac{[\text{Acid}]}{[\text{Base}]}\right)$$

When [acid]=[base] then $[H^+]=K_a$ (1/2 equiv. pt. ; pH=pK_a)

Calculating pH of Buffer after Acid is added

1 mole of base neutralizes 1 mole of acid!

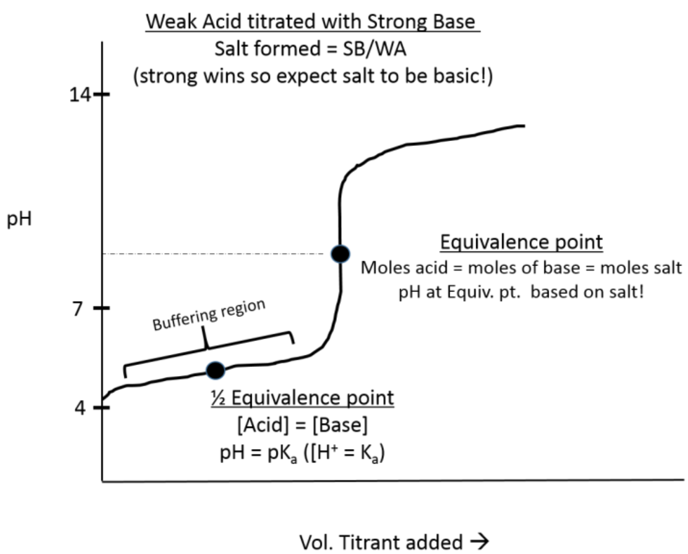
General equation here: $A^- + H^+ \rightarrow HA$

(ACID ADDED…. And subtract from base!)

$$[H^+] = K_a \left(\frac{[\text{moles of Acid} + \text{moles acid added}]}{[\text{moles of Base} - \text{moles acid added}]}\right)$$

Trick alert: Convert everything to moles!!!!

Titration Curve – Weak/Strong



- Use equiv. pt. to determine [initial] of acid
- Know what species most abundant before/after equiv. pt.
- Flip curve if you have weak base titrated with strong acid!

Calculating pH of Buffer after BASE is added

1 mole of base neutralizes 1 mole of acid!

General equation here: $HA + OH^- \rightarrow A^- + H_2O$

(BASE ADDED…. And subtract from acid!)

$$[H^+] = K_a \left(\frac{[\text{moles of Acid} - \text{moles base added}]}{[\text{moles of Base} + \text{moles base added}]}\right)$$

Trick alert: Convert everything to moles!!!!

Titration Curves and Buffering Capacity

- Buffering capacity highest when moles acid/base high

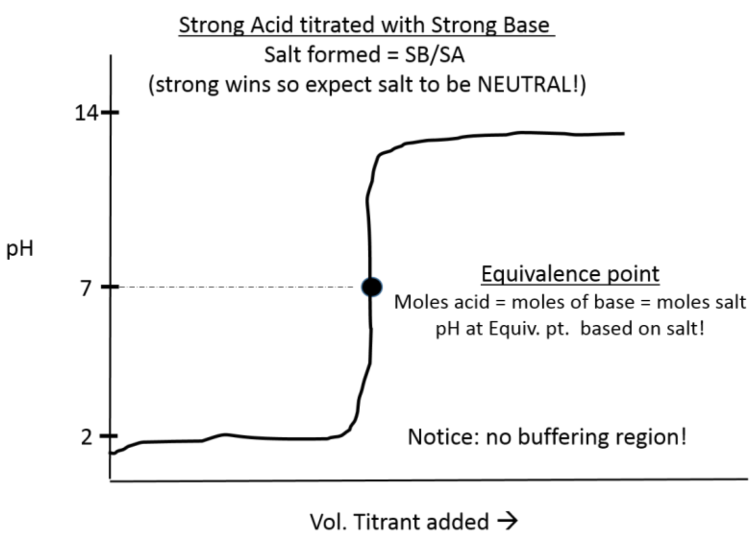
  (higher concentrations = more buffering capacity)

- Also, highest buffering capacity when [Acid]=[Base]

  (this occurs at ½ equivalence point!)

- Highest Buffering capacity when pH = pK_a of buffer

Titration Curve – Strong/Strong



Choosing Indicators (Thymol Blue example)

$HTB \rightleftharpoons TB^- + H^+$

HTB = yellow

TB^- = blue

Therefore:

- pH ABOVE pK_a mean more TB^- and solution would be blue
- pH BELOW pK_a mean more HTB and solution would be yellow

Indicator should have pK_a that is close to Equivalence point pH for titration

Common Acid–Base Indicators

Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.2–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8.2–10	colorless to pink
litmus	5.5–8.2	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

- pK_a of indicators in chart approximately in middle of color change range