

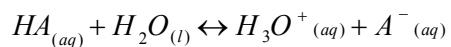
Acids and Bases Pt 1.

The transfer of protons as a driving force

Acids & (Ka)



- ✳ Have Sour taste
- ✳ React with metals to form hydrogen gas
- ✳ Turn Blue Litmus Red (BAR)
- ✳ Have pH<7
- ✳ React bases



Bases & Kb



- ✳ Bitter taste
- ✳ Soapy feel
- ✳ Turn Red Litmus Blue
- ✳ (RBB)
- ✳ Have pH>7
- ✳ React acids



Perspectives on acids vs. bases:

- ☼ Arrhenious: H^+ donors and OH^- donors.
- ☼ Bronsted- Lowry: H^+ donors and H^+ acceptors
- ☼ Lewis: e- pair deficient & e- pr donors.

S.A. Arrhenious, Nobel Prize winning chemist, born 1859.
1923 by Johannes N. Bronsted- (1879-1947) of Denmark
and Thomas M. Lowry(1874-1936) of England
1930's, the most general definition so far for
acids and bases was proposed by Gilbert N.
Lewis(1875-1946) of the United States.

Conjugate Acid-Base Pairs:

- ☼ Differ by one H^+
- ☼ Are formed by acid base reactions.
- ☼ $BA + BB \rightleftharpoons CA + CB$
 - BB is Brönsted-Base
 - BA=Brönsted acid
 - CA=Conjugate Acid
 - CB=Conjugate Base

So, Which Reaction is Significant?

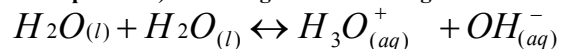
- ❖ Neutral :
 - ❖ Anions of the strong acids (except HSO_4^-)
 - ❖ Most metal cations
- ❖ Acidic:
 - ❖ Molecules containing H^+ attached to an electronegative atom
 - ❖ Cations that are conjugate acids of weak bases
 - ❖ HSO_4^-
 - ❖ Al^{+3} and other polarizable metals
- ❖ Basic:
 - ❖ Anions except the conjugate base of strong acids
 - ❖ Weak bases (covalent substances with lone pairs)
- ❖ Amphoteric:
 - ❖ Anions containing H^+ except HSO_4^- (Must compare K_a with K_b for the ion)
 - ❖ Water

Amphoteric= Amphiprotic:
able to act as either an acid
of/r as a base

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Autoionization of Water (K_w)

- Water ionizes to a very small extent ($K_w=10^{-14}$ at room temperature) according to the following reaction:



- Since water is present in all aqueous solutions, the K_w equilibrium exists in all aqueous solutions.

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pH and K_w

$$p = -\log(x) = \log(1/x)$$

From the K_w equilibrium, we derive the pH equations:

$$\begin{aligned} 2H_2O(l) &\leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)} \\ K_w &= [H_3O^+_{(aq)}][OH^-_{(aq)}] \\ \log(K_w) &= \log[H_3O^+_{(aq)}] + \log[OH^-_{(aq)}] \\ -\log(K_w) &= (-\log[H_3O^+_{(aq)}]) + (-\log[OH^-_{(aq)}]) \\ pK_w &= pH + pOH \end{aligned}$$

Thus, pH is defined for aqueous solutions only, is temperature dependent, because K_w is.

Thus, at room temp ($25^\circ C$,
 $14 = pH + pOH$)

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Aqueous acid solutions:

- Dominant Equilibrium is K_a reaction.
- Write the net ionic equation
- Set up table for starting quantities
- Determine direction of reaction, using Q
- Calculate the concentrations, in terms of x
- Solve for x by substituting into M.A.E.
- Calculate pH

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Aqueous base solutions:

- ✳ Dominant Equilibrium is K_b reaction.
- ✳ Write the net ionic equation
- ✳ Set up table for starting quantities
- ✳ Determine direction of reaction, using Q
- ✳ Calculate the concentrations, in terms of x
- ✳ Solve for x by substituting into M.A.E.
- ✳ Calculate pOH , and convert to pH

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Pt.2. Acid Base Reactions

Titration, common ions, and buffers

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Common Ion Effect

- ✳ The presence of a common ion inhibits the reaction. This is due to LeChâtelier's Principle.
- ✳ To determine the effect, decide the order of the reaction's occurrence by comparing K values.
- ✳ The most efficient reaction occurs first.
- ✳ Consider concentrations from previous rxns in subsequent equilibria.

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Mixtures of Acids

- ☼ Acids will not react with one another, but they each have a K_a reaction.
- ☼ The most efficient reaction (higher K_a) occurs 1st. Carry the common ion (H_3O^+) concentration as a starting quantity in subsequent reactions
- ☼ Determine the pH at the end.

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Rule of 20: If the K_a for the more efficient reaction is $20 > K_a$ for the next reaction, the amount of the common ion is determined by it alone. Can calculate pH from it .

This is because the common ion inhibited its production in the second reaction.

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Mixtures of Bases

- ☼ Bases will not react with one another, but they each have a K_b reaction.
- ☼ The most efficient reaction (higher K_b) occurs 1st. Carry the common ion (OH^-) concentration as a starting quantity in subsequent reactions
- ☼ Determine the pOH at the end, then find pH

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Rule of 20: If the K_b for the more efficient reaction is $20 > K_b$ for the next reaction, the amount of the common ion is determined by it alone. Can calculate pOH from it .

This is because the common ion inhibited its production in the second reaction.

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Buffers

- ☼ Buffers are a special case of the common ion effect. They are a combination of a weak acid and its conjugate base.
- ☼ They resist changes in the pH because any acid added to the mixture is reacted by the conjugate base, and vice-versa.

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Because any acid will produce its conjugate in an acid base reaction, the K_a equilibrium will serve to describe the buffer system. (Since any base will form its conjugate acid in an acid/base reaction, the K_b reaction could also be used.)

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The Henderson-Hasselbalch Equation for Buffers



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

$$\log(K_a) = \log([H_3O^+]) + \log\left(\frac{[A^-]}{[HA]}\right)$$

$$-\log([H_3O^+]) = -\log(K_a) + \log\left(\frac{[A^-]}{[HA]}\right)$$

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

Effective as long as the concentrations of each substance are within an order of magnitude of one another ($.1 < \text{ratio} < 10$)

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Buffer Characteristics

- ☛ Buffer capacity: The ability to compensate for added acid or base. Defined for each component.
 - Acid Buffer capacity = moles of conjugate base present
 - Base Buffer capacity = moles of acid present.
- ☛ To choose buffer system, try to find a system whose pKa is close to desired system pH. Then, calculate necessary ratio of components to obtain desired pH.

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Choose a buffer system to maintain a pH of 5.5. Describe its composition completely.

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Buffer reactions

- ☼ Since we have a conjugate pair, reaction (removal of) one substance results in the production (addition of) the other.
- ☼ Thus If we add an acid:

$$pH = pKa + \log\left(\frac{[base] - [acid_2]}{[acid] + [acid_2]}\right)$$

This is as long as we are within the buffering capacity

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Titration

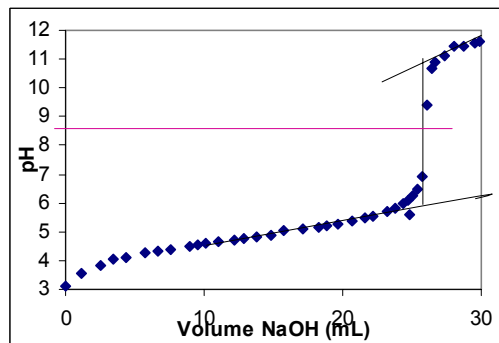
- ☼ The controlled addition of one substance to a known quantity of another substance until the stoichiometric requirements are met.
- ☼ Equivalence point: The volume of titrant needed to achieve an equal molar concentration of titrant and the titrate.
- ☼ Endpoint: The volume of titrant necessary to achieve the stoichiometric ratio of reactants. (This may be the last of several equivalence points)

Titrant: The substance being added (in burette)
Titrate: the substance to which titrant is added (in Erlenmeyer flask)

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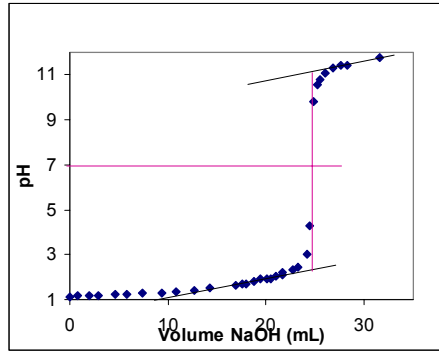
Titration of Acetic acid by NaOH



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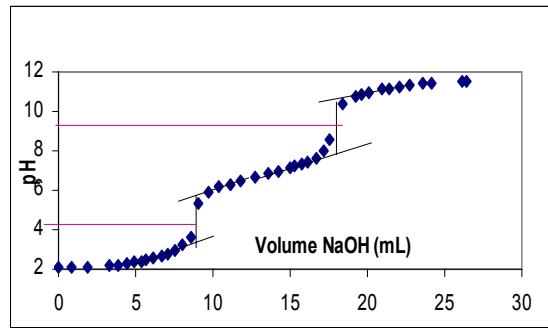
Titration of Nitric Acid by NaOH



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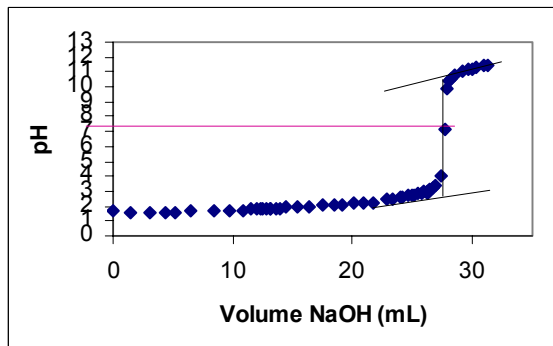
Titration of Phosphoric Acid by NaOH



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Titration of Sulfuric Acid with NaOH



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Reactions of Acids and Bases

- ☼ Produce conjugate acid and bases
- ☼ May not be neutral products:
 - ☉ SA + SB → Neutral
 - ☉ WA + SB → Basic
 - ☉ SA + WB → Acid
 - ☉ WA + WB → amphoteric check K_a vs K_b for dominance. We will not do calculations based on this type of reaction.
- ☼ Occur in order of strongest acid to weakest acid and strongest base to weakest base.

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Titration Problems

- ☼ Write the N.I.E.
- ☼ Dilute the starting quantities
- ☼ React the Ltg Rgt
- ☼ Analyze what remains:
 - ☉ Xs SA? $\text{pH} = -\log[\text{H}_3\text{O}^+]$; Xs OH? $\text{pOH} = -\log[\text{OH}^-]$
 - ☉ Xs WA? Write K_a rxn and solve for $[\text{H}_3\text{O}^+]$
 - ☉ Xs WB? Write K_b rxn and solve for $[\text{OH}^-]$
 - ☉ Buffer? Use Henderson-Hasselbalch Eqn.
- ☼ Endpoint: $V_{\text{titrant}} = V_{\text{titrate}} * C_{\text{titrate}} * (n_{\text{titrant}} / n_{\text{titrate}}) / C_{\text{titrate}}$

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25 mL of 2.5M HCl are titrated with 2.2M NaOH. What is the pH at 0, 15 and 30 mL?

Va	Ca	Vb	Cb	Vt	Cp
25		0		25+0	0
[start]					
Rxt?					

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25 mL of 2.5M HCl are titrated with 2.2M NaOH. What is the pH at 0, 15 and 30 mL?

Va	Ca	Vb	Cb	Vt
25		15		25+15
[start]				
Rxt				
remain				

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25 mL of 2.5M HCl are titrated with 2.2M NaOH. What is the pH at 0, 15 and 30 mL?

Va	Ca	Vb	Cb	Vt	Cp
25		30		25+30	0
[start]					
rxn					
remain					

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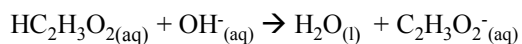
25 mL of 2.5M $\text{HC}_2\text{H}_3\text{O}_2$ are titrated with 3.2M NaOH. What is the pH at 0, 15 and 30 mL?

Va	Ca	Vb	Cb	Vt	Cp ($\text{C}_2\text{H}_3\text{O}_2^-$)
25		0			0
[start]					
rxn					

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25 mL of 2.5M HC₂H₃O₂ are titrated with 3.2M NaOH. What is the pH at 0, 15 and 30 mL?



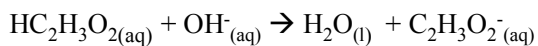
Va	Ca	Vb	Cb	Vt	Cp (C ₂ H ₃ O ₂ ⁻)
25		15		25+15	0
[start]					0
rxn					
remain					

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Recall: Buffer is mix of weak acid and its conjugate base

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25 mL of 2.5M HC₂H₃O₂ are titrated with 3.2M NaOH. What is the pH at 0, 15 and 30 mL?



Va	Ca	Vb	Cb	Vt	Cp (C ₂ H ₃ O ₂ ⁻)
25		30		25+30	0
[start]					0
rxn					
remain					

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Selection of an indicator

- ☼ Indicators change color within specific pH ranges. This is because each indicator is a weak organic acid.
- ☼ At varying concentrations of conjugate base, the color varies.
- ☼ Determine your endpoint pH
- ☼ Select an indicator whose color change is **light to dark** at this pH.

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