

Revised August 2011

HONORS TOPIC 12: Acids and Bases

- **Definitions**

Arrhenius

An Acid is a substance that dissolves in water to produce H^+ (H_3O^+) ions

A Base is a substance that dissolves in water to produce OH^- ions

Brønsted Lowry

An Acid is a proton (hydrogen ion H^+) donor

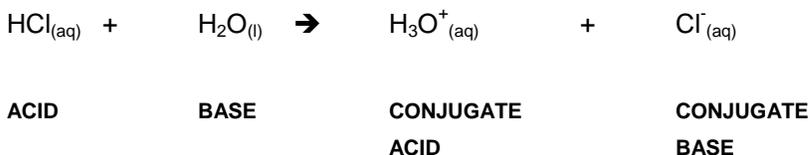
A Base is a proton (hydrogen ion H^+) acceptor

Lewis

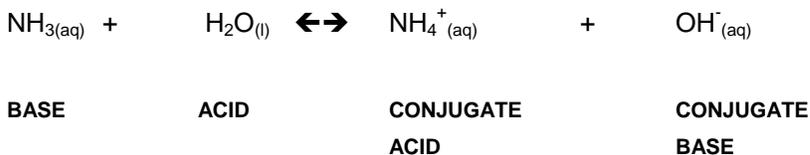
An Acid acts as an electron pair acceptor

A Base acts an electron pair donor

Example #1. Water acting as a base by accepting a proton, HCl acting as an acid by donating a proton. HCl acting as an acid by producing H_3O^+ ions in solution. H^+ ions provided by the HCl act as an acid by accepting a pair of electrons from the oxygen atom in the water (a base) to form a dative bond in H_3O^+ .



Example #2. Water acting as an acid by donating a proton, NH_3 acting as a base by accepting a proton. NH_3 acting as a base by producing OH^- ions in solution.





Example #3. Electron rich ammonia acting as a base by donating a pair of electrons to electron deficient Boron, that in turn acts as an acid by accepting the pair of electrons. A dative bond is formed.



BASE **ACID**

- **Conjugate acid & base pairs**

Conjugate acid and bases pairs are related by a hydrogen ion on either side of the equation.

Consider example #2 above. If the reaction proceeds in the forward direction NH_3 act as a base by accepting a proton to form NH_4^+ . If the reaction were to then go backwards the NH_4^+ would acts as an acid by donating a proton. NH_3 and NH_4^+ are species that can act as a base and as an acid respectively and are related by virtue of the fact they have a difference of only one H^+ . As such they are said to be a conjugate pair.



- **Strength of acids and bases**

Strong acids and strong bases exhibit complete ionization. For a strong acid (HA) and a strong base (B) the following reactions go to completion (completely to the right hand side with no reverse reaction) to produce large numbers of $\text{H}_3\text{O}^+_{(\text{aq})}$ and $\text{OH}^-_{(\text{aq})}$ respectively.



Weak acids and weak bases have very little ionization and equilibria are set up (reverse reactions also possible) with the equilibria lying heavily on the left hand side, i.e. the unionized (undissociated) form.



Examples

Weak acids	Organic acids (e.g. ethanoic acid)
Strong acids	HI, HBr, HCl, HNO_3 , H_2SO_4
Weak bases	Ammonia and organic bases (e.g. methylamine)
Strong bases	LiOH , NaOH , KOH , $\text{Ba}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$, $\text{Ca}(\text{OH})_2$



- **pH scale**

The pH scale is used to indicate the strength of an acid or a base. Traditionally, it ranges from 0-14. Acids have pH's less than 7, bases have pH's greater than 7, and 7 on the scale is considered to be neutral. pH is defined as

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \text{ or } \text{pH} = -\log [\text{H}^+]$$

In the case of strong acids and bases dissociation is complete and therefore the concentration of the H^+ ions or OH^- ions can be determined directly from the stoichiometric ratio in the balanced equation and the concentration of acid or base. Therefore we can simply plug in the numbers to the equation above, or, use the fact that at 298 K

$$14 = \text{pH} + \text{pOH}$$

$$\text{pOH} = -\log [\text{OH}^-]$$

Task 12a

1. **Calculate the pH of a solution of 0.030 M Hydrochloric acid.**
2. **Calculate the pH of a 0.010 M solution of Calcium hydroxide.**
3. **Calculate the H_3O^+ concentration in a solution with a pH of 4.32.**
4. **Calculate the pH of a solution made by dissolving 2.00 g of potassium hydroxide in distilled water to a total volume of 250. mL.**



- **Ka The equilibrium constant for weak acids**

For a weak acid (HA) dissociation is incomplete and the following equilibrium is set up.

	$\text{HA}_{(\text{aq})}$	+	$\text{H}_2\text{O}_{(\text{l})}$	\rightleftharpoons	$\text{H}_3\text{O}^+_{(\text{aq})}$	+	$\text{A}^-_{(\text{aq})}$
Initial	X				0		0
Change	- x				+ x		+ x
Equilibrium	X - x				0 + x		0 + x

In this case, we are not sure of how much of the acid has dissociated, so it is not possible to go directly to $\text{pH} = -\log [\text{H}_3\text{O}^+_{(\text{aq})}]$ from the concentration of the acid. However, we do know that since the acid is weak and there is very little dissociation, x will be small and can be considered negligible in the term X-x. Additionally we know that the concentrations of $\text{H}_3\text{O}^+_{(\text{aq})}$ and $\text{A}^-_{(\text{aq})}$ will be equal. Considering those facts, and knowing that $\text{H}_2\text{O}_{(\text{l})}$ is a pure liquid and will not appear in the Ka expression, the following expressions can be derived.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

Often written as,

$$K_a = \frac{[\text{H}_3\text{O}^+]^2}{[\text{HA}]}$$

Another useful relationship in calculations is

$$\text{pKa} = -\log K_a$$

Task 12b

1. *If ethanoic acid has a pKa of 4.74, what is its Ka?*
2. *If propanoic acid has a Ka of 1.38×10^{-5} , what is its pKa?*
3. *Which is the stronger acid, propanoic or ethanoic?*
4. *What is the pH of a 1.00 M solution of ethanoic acid?*
5. *Calculate the Ka value of a 0.100 M solution of a weak acid with a $[\text{H}^+]$ of 1.75×10^{-3} M.*
6. *Calculate the Ka of a solution of 0.250 M of a weak acid with a pH of 5.11.*



- **Kb The equilibrium constant for weak bases**

For a weak base dissociation is incomplete. For example the following equilibrium is set up for ammonia.

	$\text{NH}_3(\text{aq})$	+	$\text{H}_2\text{O}(\text{l})$	\rightleftharpoons	$\text{NH}_4^+(\text{aq})$	+	$\text{OH}^-(\text{aq})$
Initial	X				0		0
Change	- x				+ x		+ x
Equilibrium	X - x				0 + x		0 + x

In this case, we are not sure of how much of the base has dissociated, so it is not possible to go directly to $\text{pOH} = -\log[\text{OH}^-(\text{aq})]$ from the concentration of the base. However, we do know that since the base is weak and there is very little dissociation, x will be small and can be considered negligible in the term X-x. Additionally we know that the concentrations of $\text{NH}_4^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ will be equal. Considering those facts, and knowing that $\text{H}_2\text{O}(\text{l})$ is a pure liquid and will not appear in the Kb expression, the following expressions can be derived.

$$K_b = \frac{[\text{OH}^-][\text{NH}_4^+]}{[\text{NH}_3]}$$

Often written as,

$$K_b = \frac{[\text{OH}^-]^2}{[\text{NH}_3]}$$

Another useful relationship in calculations is

$$\text{pKb} = -\log K_b$$

Task 12c

1. *If ammonia has a pKb of 4.75, what is its Kb?*
2. *If methylamine has a Kb of 4.38×10^{-4} , what is its pKb?*
3. *Which is the stronger base, ammonia or methylamine?*
4. *What is the pH of a 0.44 M solution of ammonia?*
5. *Calculate the Kb value of a 0.100 M solution of a weak base with $[\text{OH}^-]$ of 1.75×10^{-3} M.*
6. *Calculate the Kb of a solution of 0.250 M of a weak base with a pH of 9.12.*



- **Indicators**

An acid base indicator is a substance that changes color according to the pH of the solution. They are often weak acids where the ionized and the unionized form have different colors.



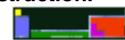
COLOR 1

COLOR 2

This equilibrium will be disturbed by the addition of $\text{H}^+_{(aq)}$ or $\text{OH}^-_{(aq)}$. For example, a high concentration of $\text{H}^+_{(aq)}$ (acid solution) will shift the equilibrium to the LHS giving COLOR 1. A high concentration of $\text{OH}^-_{(aq)}$ (basic solution) uses up $\text{H}^+_{(aq)}$ to form water and will shift the equilibrium to the RHS giving COLOR 2. In practice the indicator will change color over a small, given range of pH.

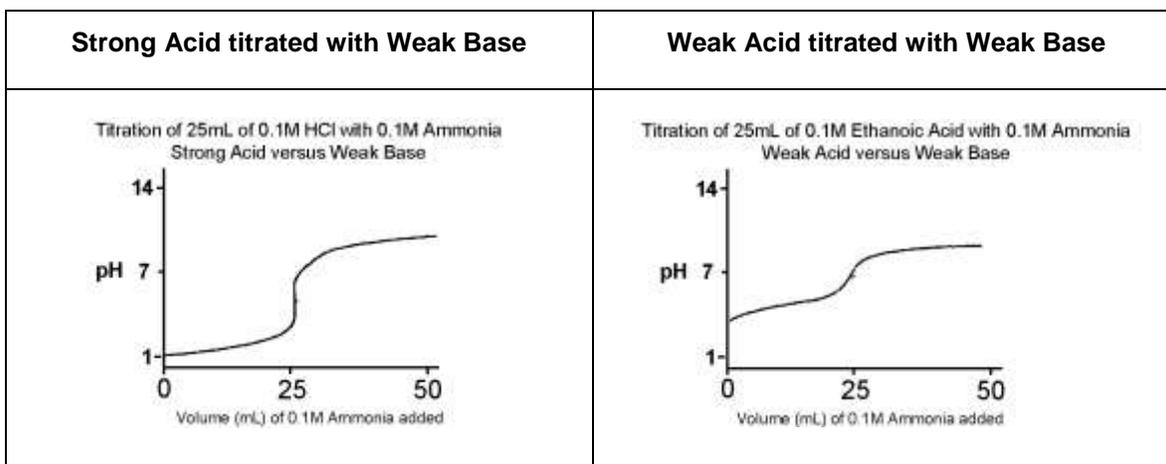
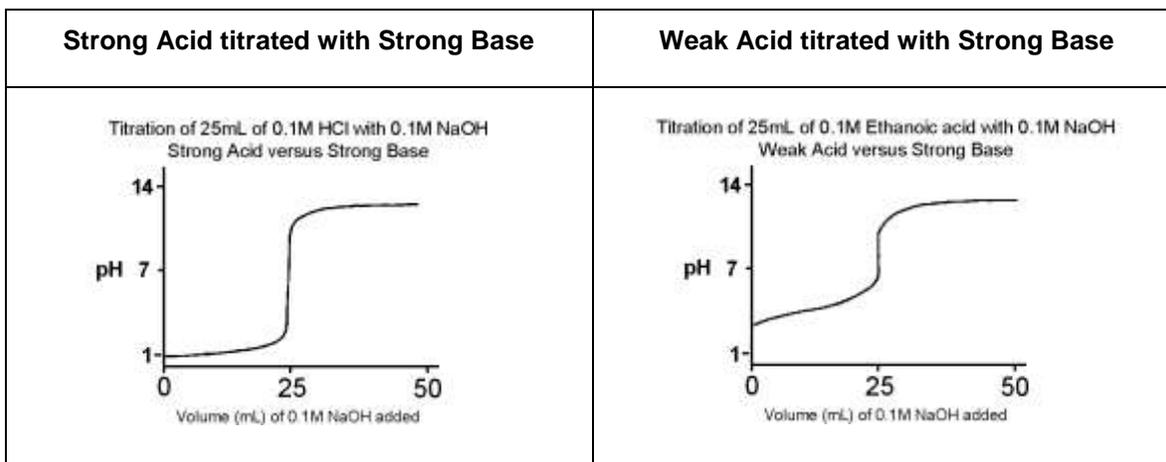
Examples

	Color in acid conditions	Approximate point at which color change takes place	Color in basic conditions
Methyl orange	Red	5-6	Yellow
Methyl red	Red	6-7	Yellow
Litmus	Red	7-8	Blue
Phenolphthalein	Colorless	9-10	Pink



Choice of indicator and titration curves

The following titration curves can be plotted from data.



The equivalence point represents the point at which equimolar amounts of acid and base have reacted and is located in the middle of the vertical portion of the titration curve. Since in an acid/base titration we need to find the equivalence point, we must choose an indicator that changes color (reaches the end point) over the pH range of the sharp vertical step on each graph.

So using the titration plots and the table above, suitable indicators can be chosen;

- | | |
|-------------------------|-----------------|
| Strong acid Strong base | most indicators |
| Weak acid Strong base | phenolphthalein |
| Strong acid Weak base | methyl orange |

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For a weak acid weak base titration there is no sharp change in pH at the equivalence point. Therefore no indicator will change color sharply at the end-point, therefore no indicator is suitable. A pH meter can be used to determine the end-point in such titrations.