

IV. Acids and Bases

Properties of Acids and Bases

Acids	Bases
pH < 7.0	pH > 7.0
taste sour	taste bitter
react with metals to produce hydrogen gas	feel slippery
pH paper turns red/orange	pH paper turns blue/green
phenolphthalein → colourless	phenolphthalein → pink
bromothymol blue → yellow	bromothymol blue → blue
cabbage juice → pink	cabbage juice → blue

pH and pOH

pH ("power of hydrogen") is a measure of the **acidity** of a solution: the **lower** the pH, the more acidic the solution is and the **higher** the pH, the more basic the solution is

pOH is a measure of the **basicity** (or alkalinity) of a solution: the **lower** the pOH, the more basic the solution is and the **higher** the pOH, the more acidic the solution is

pH or pOH = 7.0 indicates a **neutral** solution

The pH Scale

1.0	1.0x10 ⁻¹	1.0x10 ⁻²	1.0x10 ⁻³	1.0x10 ⁻⁴	1.0x10 ⁻⁵	1.0x10 ⁻⁶	1.0x10 ⁻⁷	1.0x10 ⁻⁸	1.0x10 ⁻⁹	1.0x10 ⁻¹⁰	1.0x10 ⁻¹¹	1.0x10 ⁻¹²	1.0x10 ⁻¹³	1.0x10 ⁻¹⁴	[H ⁺]
0	1.0	2.0	3.0	4.0	5.0	6.0	7.0	8.0	9.0	10.0	11.0	12.0	13.0	14.0	pH
ACID														BASE	
14.0	13.0	12.0	11.0	10.0	9.0	8.0	7.0	6.0	5.0	4.0	3.0	2.0	1.0	0	pOH
1.0x10 ⁻¹⁴	1.0x10 ⁻¹³	1.0x10 ⁻¹²	1.0x10 ⁻¹¹	1.0x10 ⁻¹⁰	1.0x10 ⁻⁹	1.0x10 ⁻⁸	1.0x10 ⁻⁷	1.0x10 ⁻⁶	1.0x10 ⁻⁵	1.0x10 ⁻⁴	1.0x10 ⁻³	1.0x10 ⁻²	1.0x10 ⁻¹	1.0	[OH ⁻]

Arrhenius Concept

Acid: substance that will dissociate to give H⁺ ions

Base: substance that will dissociate to give OH⁻ ions

The Arrhenius concept provides a good description of **strong** acids and bases

Strength of Acids and Bases

The strength of acids and bases is a measure of the degree of ionization; a strong acid or base will completely dissociate in water.

Strong Acids:

HCl → H ⁺ + Cl ⁻	hydrochloric acid	HNO ₃ → H ⁺ + NO ₃ ⁻	nitric acid
HBr → H ⁺ + Br ⁻	hydrobromic acid	HIO ₄ → H ⁺ + IO ₄ ⁻	periodic acid
HI → H ⁺ + I ⁻	hydroiodic acid	H ₂ SO ₄ → H ⁺ + HSO ₄ ⁻	sulphuric acid
HClO ₄ → H ⁺ + ClO ₄ ⁻	perchloric acid	HClO ₃ → H ⁺ + ClO ₃ ⁻	chloric acid

Strong Bases:

LiOH → Li ⁺ + OH ⁻	lithium hydroxide	Ba(OH) ₂ → Ba ²⁺ + 2OH ⁻	barium hydroxide
NaOH → Na ⁺ + OH ⁻	sodium hydroxide	Sr(OH) ₂ → Sr ²⁺ + 2OH ⁻	strontium hydroxide
KOH → K ⁺ + OH ⁻	potassium hydroxide	Ca(OH) ₂ → Ca ²⁺ + 2OH ⁻	calcium hydroxide
RbOH → Rb ⁺ + OH ⁻	rubidium hydroxide	(note: other group II elements are insoluble in water)	
CsOH → Cs ⁺ + OH ⁻	cesium hydroxide		

Calculations for Strong Acids and Bases

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

$$[\text{H}^+] = 10^{-\text{pH}}$$

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

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

$$\text{pH} + \text{pOH} = 14.00$$

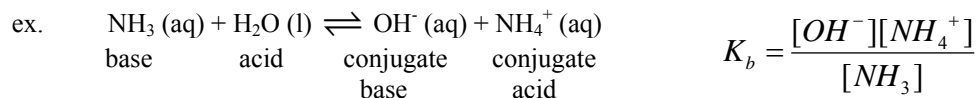
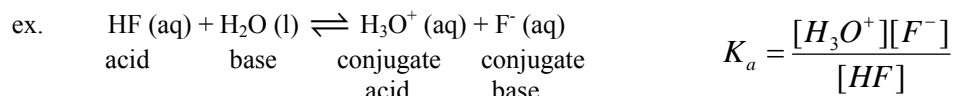
$$1.0 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$$

K_a values for Common Acids. The higher the K_a value, the stronger the acid.

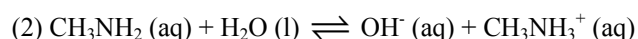
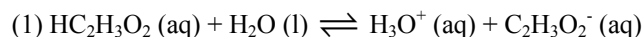
Acid	Formula	K_a
iodic acid	HIO_3	0.17
oxalic acid	$\text{H}_2\text{C}_2\text{O}_4$	5.9×10^{-2}
sulphurous acid	H_2SO_3	1.5×10^{-2}
hydrogen sulphate ion	HSO_4^-	1.2×10^{-2}
chlorous acid	HClO_2	1.2×10^{-2}
phosphoric acid	H_3PO_4	7.5×10^{-3}
hydrofluoric acid	HF	7.2×10^{-4}
nitrous acid	HNO_2	4.0×10^{-4}
hydrogen oxalate ion	HC_2O_4^-	6.1×10^{-5}
acetic acid	$\text{HC}_2\text{H}_3\text{O}_2$ (or CH_3COOH)	1.8×10^{-5}
carbonic acid	H_2CO_3	4.3×10^{-7}
hydrogen sulphite ion	HSO_3^-	1.0×10^{-7}
hydrosulphuric acid	H_2S	1.0×10^{-7}
dihydrogen phosphate ion	H_2PO_4^-	6.2×10^{-8}
hypochlorous acid	HOCl	3.5×10^{-8}
hypobromous acid	HOBr	2.0×10^{-9}
hydrocyanic acid	HCN	6.2×10^{-10}
boric acid	H_3BO_3	5.8×10^{-10}
ammonium ion	NH_4^+	5.6×10^{-10}
hydrogen carbonate ion	HCO_3^-	5.6×10^{-11}
hydrogen phosphate ion	HPO_4^{2-}	4.8×10^{-13}
hydrogen sulphide ion	HS^-	1.3×10^{-13}

K_b values for Common Bases. The higher the K_b value, the stronger the base.

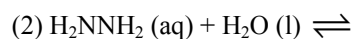
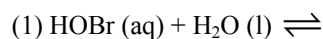
Base	Formula/ Conjugate Acid	K_b
diethylamine	$(\text{C}_2\text{H}_5)_2\text{NH} / (\text{C}_2\text{H}_5)_2\text{NH}_2^+$	1.3×10^{-3}
ethylamine	$\text{C}_2\text{H}_5\text{NH}_2 / \text{C}_2\text{H}_5\text{NH}_3^+$	5.6×10^{-4}
methylamine	$\text{CH}_3\text{NH}_2 / \text{CH}_3\text{NH}_3^+$	4.4×10^{-4}
triethylamine	$(\text{C}_2\text{H}_5)_3\text{N} / (\text{C}_2\text{H}_5)_3\text{NH}^+$	4.0×10^{-4}
ammonia	$\text{NH}_3 / \text{NH}_4^+$	1.8×10^{-5}
hydrazine	H_2NNH_2 (or N_2H_4) / H_2NNH_3^+	3.0×10^{-6}
hydroxylamine	$\text{HONH}_2 / \text{HONH}_3^+$	1.1×10^{-8}
pyridine	$\text{C}_5\text{H}_5\text{N} / \text{C}_5\text{H}_5\text{NH}^+$	1.7×10^{-9}
aniline	$\text{C}_6\text{H}_5\text{NH}_2 / \text{C}_6\text{H}_5\text{NH}_3^+$	3.8×10^{-10}



ex. Write a K_a/K_b expression for the following acid/base equilibria. Identify the conjugate acid-base pairs.



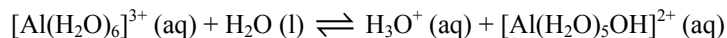
ex. Complete the following acid-base equilibria. Identify the conjugate acid-base pairs and write a K_a/K_b expression.



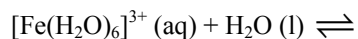
Hydrated Metal Ions

Certain metal ions (particularly transition metal ions and aluminum) when in solution will form hydrated complexes that may act as acid.

ex. When Al^{3+} is in solution, it forms $[\text{Al}(\text{H}_2\text{O})_6]^{3+}$, the acid equilibrium for this complex is as follows: ($K_a = 1.4 \times 10^{-5}$)



ex. When Fe^{3+} is in solution, it forms $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$. What is the acid equilibrium for this complex? ($K_a = 6.0 \times 10^{-3}$)



Calculations for Weak Acids and Bases

The equilibrium concentration of the hydronium ion or hydroxide ion must be determined using an ICE box.

ex. Determine the pH and pOH of 1.0 M solution of HF ($K_a = 7.2 \times 10^{-4}$).

ex. Determine the pOH and pH of a 1.0 M solution of NH_3 ($K_b = 1.8 \times 10^{-5}$).

ex. The pH of a 0.650 M solution of citric acid ($\text{H}_3\text{C}_6\text{H}_5\text{O}_7$) is 1.63. Determine K_a .

ex. The pOH of a 0.020 M solution of toluidine ($\text{CH}_3\text{C}_6\text{H}_4\text{NH}_2$) is 5.35. Determine K_b .

Percent Dissociation (also called percent ionization)

The percent dissociation of a compound is a measure of the degree of dissociation.

The percent dissociation can be calculated according to the following equation:

$$\text{Percent dissociation} = \frac{\text{amount dissociated}}{\text{initial concentration}} \times 100\% \quad \text{amount dissociated} = "x"$$

ex. Calculate the percent dissociation for a 0.22 M solution of $\text{C}_2\text{H}_5\text{NH}_2$ ($K_b = 5.6 \times 10^{-4}$)

ex. A 0.100 M solution of lactic acid ($\text{HC}_3\text{H}_5\text{O}_3$) is 3.7% dissociated at equilibrium. Calculate the K_a value for lactic acid.

Amphoteric Compounds

amphoteric: a compound that can act as either an acid or a base.

ex. Water is amphoteric: $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$

note: the autoionization of water produces both a hydronium ion and a hydroxide ion so the resulting solution is neutral.

The equilibrium expression for this reaction is as follows:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-], \text{ where } K_w = 1.0 \times 10^{-14} \text{ (} K_w \text{ is called the dissociation constant of water)}$$

other amphoteric substances include: the hydrogen sulphate ion (HSO_4^-), the hydrogen sulphite ion (HSO_3^-), the dihydrogen phosphate ion (H_2PO_4^-), the hydrogen phosphate ion (HPO_4^{2-}), the hydrogen sulphide ion (HS^-) etc.

Relationship between K_a and K_b

For a weak acid, if the K_a value is known, then the K_b value of its conjugate base can be calculated. For a weak base, if the K_b value is known, then the K_a value for its conjugate acid can be calculated. This relationship is described according to the following equation:

$$K_w = K_a \times K_b \quad K_w = 1.0 \times 10^{-14}, K_a \text{ is the acid dissociation constant, and } K_b \text{ is the base dissociation constant}$$

Strength of a conjugate acid or base

The stronger the acid, the weaker its conjugate base and the stronger the base, the weaker its conjugate acid.

ex. Which of the following substances is a stronger acid, HF or HCN? Which of the following substances is a stronger base, F^- or CN^- ?

HF has $K_a = 7.2 \times 10^{-4}$ and HCN has $K_a = 6.2 \times 10^{-10}$. Since K_a for HF is greater than K_a for HCN, HF is a stronger acid than HCN. F^- has $K_b = 1.0 \times 10^{-14} / 7.2 \times 10^{-4} = 1.4 \times 10^{-11}$ and CN^- has $K_b = 1.0 \times 10^{-14} / 6.2 \times 10^{-10} = 1.6 \times 10^{-5}$. Since K_b for CN^- is greater than K_b for F^- , CN^- is a stronger base than F^- .

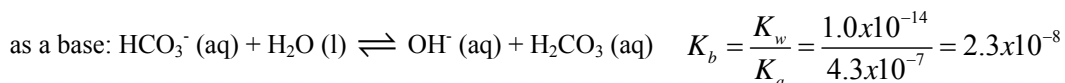
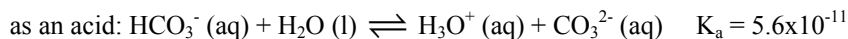
ex. Which of the following substances is a stronger base, NH_3 or $\text{C}_5\text{H}_5\text{N}$? Which of the following substances is a stronger acid, NH_4^+ or $\text{C}_5\text{H}_5\text{NH}^+$?

NH_3 has $K_b = 1.8 \times 10^{-5}$ and $\text{C}_5\text{H}_5\text{N}$ has $K_b = 1.7 \times 10^{-9}$. Since K_b for NH_3 is greater than K_b for $\text{C}_5\text{H}_5\text{N}$, NH_3 is a stronger base than $\text{C}_5\text{H}_5\text{N}$.

NH_4^+ has $K_a = 5.6 \times 10^{-10}$ (given on K_a table) and $\text{C}_5\text{H}_5\text{NH}^+$ has $K_a = 1.0 \times 10^{-14} / 1.7 \times 10^{-9} = 5.9 \times 10^{-9}$. Since K_a for $\text{C}_5\text{H}_5\text{NH}^+$ is greater than K_a for NH_4^+ , $\text{C}_5\text{H}_5\text{NH}^+$ is a stronger acid than NH_4^+ .

Comparing Equilibria for Amphoteric compounds

ex. the hydrogen carbonate ion, HCO_3^- can act as an acid (donating a proton) or a base (accepting a proton)

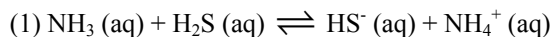


Since $K_b > K_a$, a solution containing hydrogen carbonate would have a *basic* pH.

Acid Base Reactions

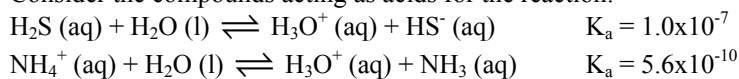
Weak acids can react with weak bases and result in an acid/base equilibrium.

Consider the following reactions and identify the conjugate acid-base pairs.

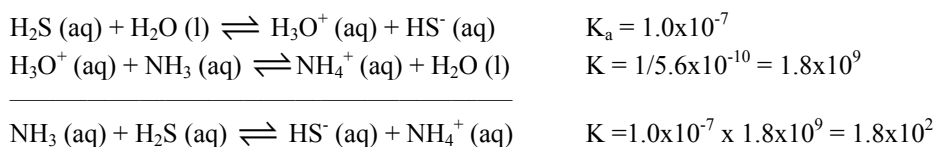


Calculating K for an acid-base reaction:

Consider the compounds acting as acids for the reaction:



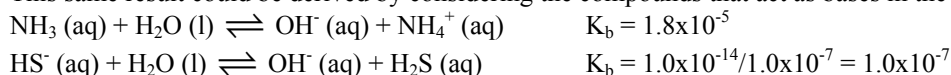
Note, if the second equilibrium is reversed (and the reciprocal of the equilibrium constant is taken), the two equilibria can add up to give the overall equilibrium (and the K for the equilibrium can be obtained by multiplying K for each of the reactions)



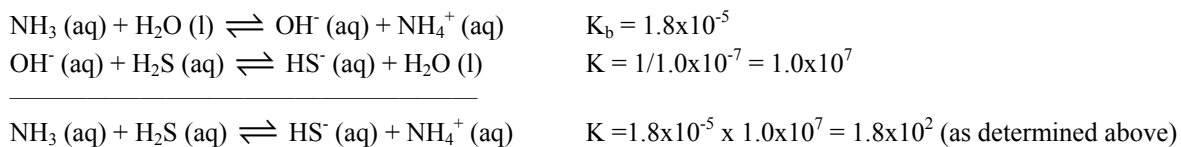
(note: $\text{H}_3\text{O}^+ (\text{aq})$ and $\text{H}_2\text{O} (\text{l})$ are cancelled when the equilibria are added)

An important result is that $K > 1$. This is significant since H_2S is a stronger acid than NH_4^+ , therefore the equilibrium position is driven to the right, favouring the products.

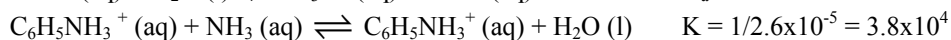
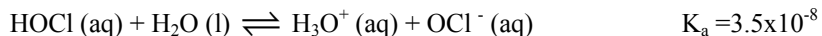
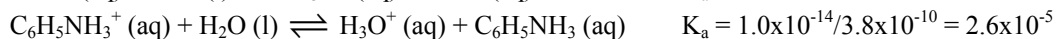
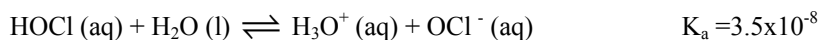
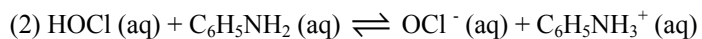
This same result could be derived by considering the compounds that act as bases in the reaction:



Note, if the second equilibrium is reversed (and the reciprocal of the equilibrium constant is taken), the two equilibria can add up to give the overall equilibrium (and the K for the equilibrium can be obtained by multiplying K for each of the reactions)



(note: $\text{OH}^- (\text{aq})$ and $\text{H}_2\text{O} (\text{l})$ are cancelled when the equilibria are added)



An important result is that $K < 1$. This is significant since $\text{C}_6\text{H}_5\text{NH}_3^+$ is a stronger acid than HOCl, therefore the equilibrium position is driven to the left, favouring the reactants.

(again, the K value could also be calculated by considering the compounds acting as bases)

Mixtures of Acids

In general, only the strongest acid will significantly contribute to the concentration of hydronium ions.

ex. Calculate the pH of a mixture containing 0.50 M HCl and 0.50 M HF.

Since HCl is a strong acid and HF is a weak acid, the HF will not significantly contribute to the H^+ concentration, so the pH can be calculated from the HCl concentration.

ex. Calculate the pH of a mixture containing 0.10 M HNO_2 ($K_a = 4.0 \times 10^{-4}$) and 0.10 M HCN ($K_a = 6.2 \times 10^{-10}$)

Since K_a for $\text{HNO}_2 \gg$ than K_a for HCN, only the HNO_2 will significantly contribute to $[\text{H}_3\text{O}^+]$. An ICE table can then be used to determine pH.

Mixtures of Bases

In general, only the strongest base will significantly contribute to the concentration of hydroxide ions.

ex. Calculate the pOH of a mixture containing 0.20 M Sr(OH)_2 and 0.20 M NH_3 .

Since Sr(OH)_2 is a strong base and NH_3 is a weak base, the NH_3 will not significantly contribute to the OH^- concentration, so the pOH can be calculated from the Sr(OH)_2 concentration.

ex. Calculate the pOH of a mixture containing 0.50 M $\text{C}_6\text{H}_5\text{NH}_2$ ($K_b = 3.8 \times 10^{-10}$) and 0.50 M NH_3 ($K_b = 1.8 \times 10^{-5}$)

Since K_b for $\text{NH}_3 \gg$ than K_b for $\text{C}_6\text{H}_5\text{NH}_2$, only the NH_3 will significantly contribute to $[\text{OH}^-]$. An ICE table can then be used to determine pOH.

Solutions of Acids or Bases Containing a Common Ion of the Conjugate base/acid

An ICE table can be used to determine pH/pOH, where the initial concentration of the conjugate ion will not be zero.

ex. A solution contains 1.0 M HF and 1.0 M NaF, Calculate the pH of the resulting solution. (How does this pH compare to a 1.0 M solution of HF?)

ex. A solution contains 1.0 M NH_3 and 1.0 M NH_4Cl , Calculate the pH of the resulting solution. (How does this pH compare to a 1.0 M solution of NH_3 ?)

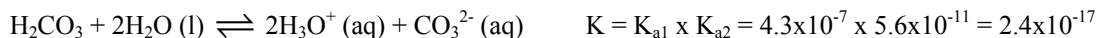
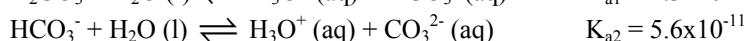
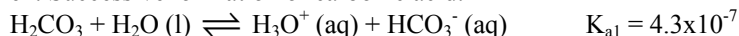
Number of Acidic Protons

monoprotic acids: have only one acidic hydrogen atoms (ex. HCl and HC₂H₃O₂)

diprotic acids: have two acidic hydrogen atoms (ex. H₂SO₄ and H₂C₂O₄)

polyprotic acids: have three or more acidic hydrogen atoms (ex. H₃PO₄)

ex. Successive ionization of carbonic acid.



For a diprotic acid, $K_{a1} > K_{a2}$. (The first proton is more easily lost than the second proton.)

Acid-Base Properties of Salts

The cation or anion of a salt may hydrolyze (react with water) and produce an acidic or basic solution.

Cations

Cations found in strong bases will not hydrolyze.

Cations that are the conjugate acids of weak bases will hydrolyze to produce acidic solutions.

Anions

Anions found in strong acids will not hydrolyze.

Anions that are the conjugate bases of weak acids will hydrolyze to produce basic solutions.

ex. Determine if the pH of the following solutions will be acidic, basic, or neutral.

(1) NaCl

NaOH is a strong base so Na⁺ will not hydrolyze

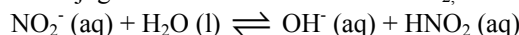
HCl is a strong acid so Cl⁻ will not hydrolyze

Since neither the cation or anion hydrolyze, the solution will be neutral

(2) NaNO₂

NaOH is a strong base so Na⁺ will not hydrolyze

NO₂⁻ is the conjugate base of the weak acid HNO₂, so NO₂⁻ will hydrolyze:

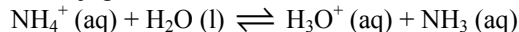


The resulting hydrolysis reaction produces hydroxide ions so the solution is basic.

(3) NH₄Cl

HCl is a strong acid so Cl⁻ will not hydrolyze

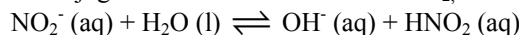
NH₄⁺ is the conjugate acid of the weak base NH₃, so NH₄⁺ will hydrolyze:



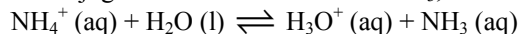
The resulting hydrolysis reaction produces hydronium ions so the solution is acidic.

(4) NH₄NO₂

NO₂⁻ is the conjugate base of the weak acid HNO₂, so NO₂⁻ will hydrolyze:



NH₄⁺ is the conjugate acid of the weak base NH₃, so NH₄⁺ will hydrolyze:



Since both ions hydrolyze, the K_a and K_b values for the resulting equilibria must be compared.

HNO₂ has K_a = 4.0 × 10⁻⁴, therefore, NO₂⁻ has K_b = 2.5 × 10⁻¹¹
NH₄⁺ has a K_a = 5.6 × 10⁻¹⁰

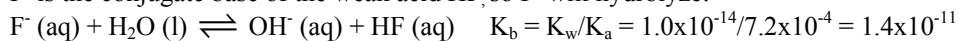
Since K_a > K_b, the hydrolysis of NH₄⁺ will favour the products more than the hydrolysis of NO₂⁻, as a result there will be more hydronium ions than hydroxide ions and the solution will be acidic.

Calculations for Salts

ex. Calculate the pH and pOH of a 0.20 M solution of NaF.

NaOH is a strong base so Na^+ will not hydrolyze

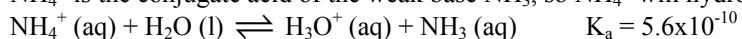
F^- is the conjugate base of the weak acid HF, so F^- will hydrolyze:



ex. Calculate the pH and pOH of a 0.015 M solution of NH_4Br .

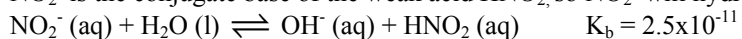
HBr is a strong acid so Br^- will not hydrolyze

NH_4^+ is the conjugate acid of the weak base NH_3 , so NH_4^+ will hydrolyze:

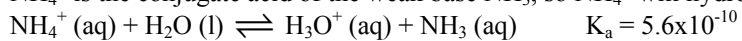


ex. Calculate the pH and pOH of a 0.120 M solution of NH_4NO_2 .

NO_2^- is the conjugate base of the weak acid HNO_2 , so NO_2^- will hydrolyze:



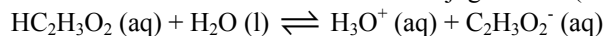
NH_4^+ is the conjugate acid of the weak base NH_3 , so NH_4^+ will hydrolyze:



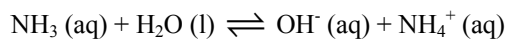
Buffers

buffer: is a solution that resists changes in pH.

A buffer contains a weak acid and a salt of its conjugate base (ie. $\text{HC}_2\text{H}_3\text{O}_2$ and $\text{NaC}_2\text{H}_3\text{O}_2$)



OR a weak base and a salt of its conjugate acid (ie. NH_3 and NH_4Cl).



Buffer Calculations

Buffer calculations are based on calculations for solutions with a common ion.

ex. A solution contains 0.50 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.50 M $\text{NaC}_2\text{H}_3\text{O}_2$. Calculate the pH of the solution.

Determine the resulting pH if 0.010 mol of NaOH are added to 1.0 L of this buffer.

Determine the resulting pH if 0.010 mol of HCl are added to 1.0 L of this buffer.

ex. A solution contains 0.25 M NH_3 and 0.25 M NH_4Cl . Calculate the pH of the solution.

Determine the resulting pH if 0.010 mol of HCl are added to 1.0 L of this buffer.

Determine the resulting pH if 0.010 mol of NaOH are added to 1.0 L of this buffer.

The Henderson-Hasselbalch Equation

Determining the pH of any solution containing a weak acid and its conjugate base can be simplified using the following equation:

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

Determining the pOH of any solution containing a weak base and its conjugate acid can be simplified using the following equation:

$$pOH = pK_b + \log\left(\frac{[BH^+]}{[B]}\right) \quad \text{where: } pK_b = -\log K_b$$

Preparing Buffers

ex. Calculate the pH of a buffer made by adding 0.50 g of sodium fluoride to 200 mL of 0.10 M hydrofluoric acid. Assume the volume of the solution remains constant.

ex. Determine the mass of methylammonium chloride ($\text{CH}_3\text{NH}_3\text{Cl}$) that should be added to 300 mL of 0.20 M methylamine (CH_3NH_2) in order to make a buffer with a pH of 11.40. Assume the volume of the solution remains constant.

Titration

titration: an analytical procedure in which the concentration of a known solution is used to determine the concentration of an unknown solution (used for acids and bases, oxidation and reduction, and solubility, etc)

Acid Base Titrations

Acid base titrations involve neutralization reactions (recall, acids and bases react to form salt and water)

For an experiment, determine the volume of a solution with known concentration that is required to titrate (exactly react with) a given volume of solution with an unknown concentration.

The solution with known concentration is called the **titrant**. The solution with the unknown concentration is called the **analyte**.

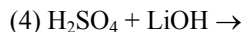
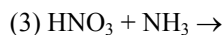
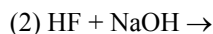
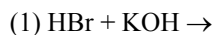
Usually several "trials" are performed and the average volume is used for calculations

Chemical indicators (ie. phenolphthalein or bromothymol blue) can be used to monitor when the reaction is complete and has reached its equivalence point.

Neutralizations Reactions

To determine the net ionic equation for a neutralization reaction, the spectator ions must be decided. Only strong acids, strong bases, and soluble salts will dissociate completely to form ions. Weak acids, weak bases, and insoluble salts will not significantly dissociate.

ex. Write the net ion equation for the following neutralization reactions.



Titration Calculations

ex. What is the volume of 0.50 M HCl required to titrate 150.0 mL of 0.25 M NH_3 ?

ex. What is the concentration if 50.00 mL of HF requires 25.00 mL of 0.10 M NaOH to titrate?

Titration Curve Calculations

Strong Acid/ Strong Base

ex. A 0.100 M solution of NaOH is used to titrate 100.0 mL of a 0.100 M HCl solution.

Determine the pH at the following intervals in the experiment.

(1) No NaOH has been added pH =

(5) 100.0 mL of NaOH has been added pH =

(2) 25.00 mL of NaOH has been added pH =

(6) 125.0 mL of NaOH has been added pH =

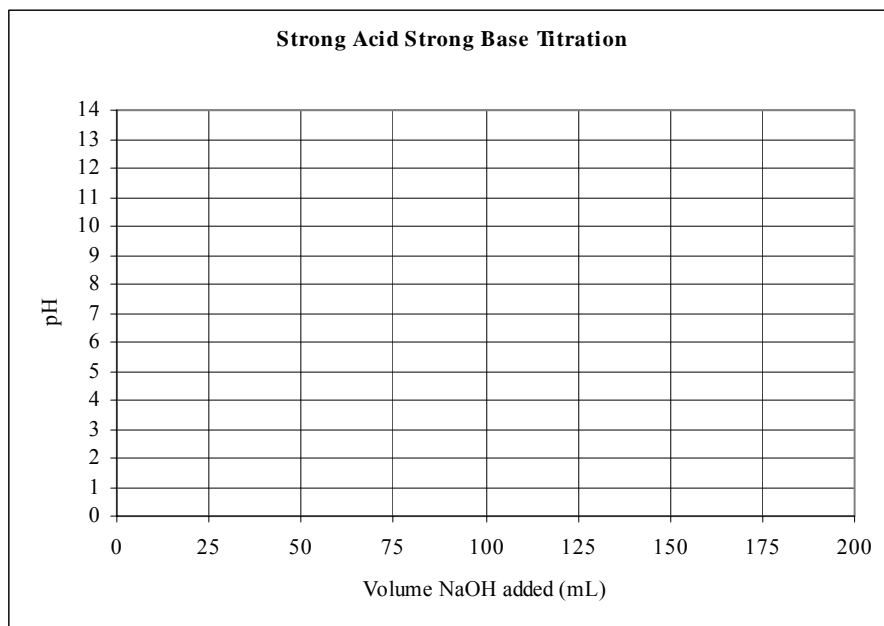
(3) 50.00 mL of NaOH has been added pH = 1.48

(7) 150.00 mL of NaOH has been added pH = 12.30

(4) 75.00 mL of NaOH has been added pH = 1.85

(8) 175.0 mL of NaOH has been added pH = 12.43

(9) 200.0 mL of NaOH has been added. pH = 12.52



Strong Acid/Weak Base

ex. A 0.100 M solution of HCl is used to titrate 100.0 mL of a 0.100 M NH₃ solution.

Determine the pH at the following intervals in the experiment.

(1) No HCl has been added pH =

(5) 100.0 mL of HCl has been added pH =

(2) 25.00 mL of HCl has been added pH =

(6) 125.0 mL of HCl has been added pH =

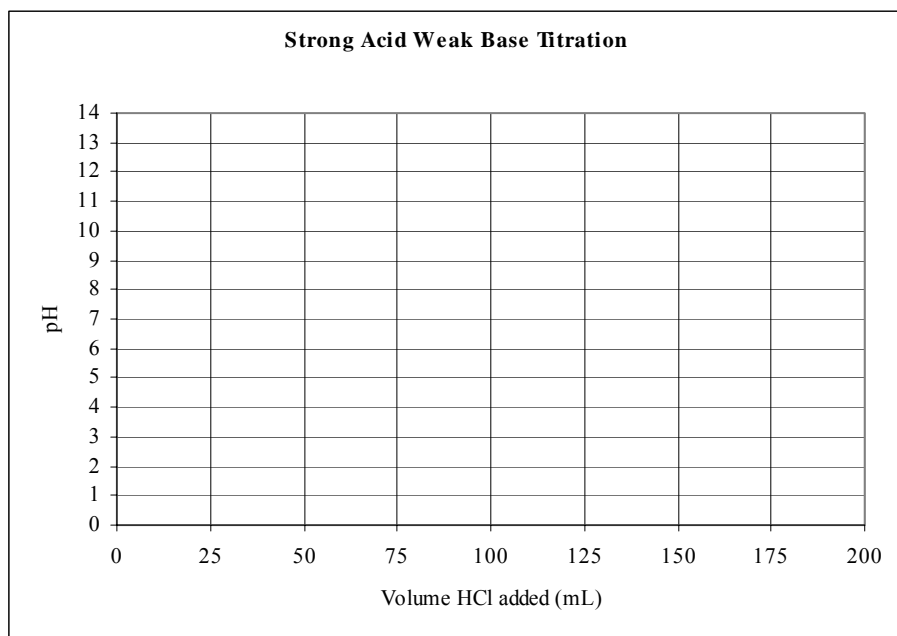
(3) 50.00 mL of HCl has been added pH = 9.26

(7) 150.00 mL of HCl has been added pH = 1.70

(4) 75.00 mL of HCl has been added pH = 8.77

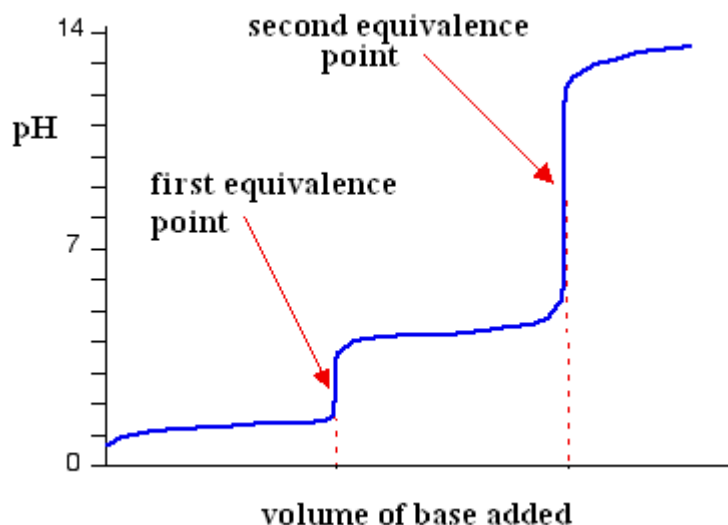
(8) 175.0 mL of HCl has been added pH = 1.56

(9) 200.0 mL of HCl has been added. pH = 1.30



Diprotic Titration

If a strong base is titrated with a weak diprotic acid, there will be two equivalence points as shown in the following titration curve.



Lewis Acid Base Model

The Lewis acid base model provides the most general description of an acid or base.

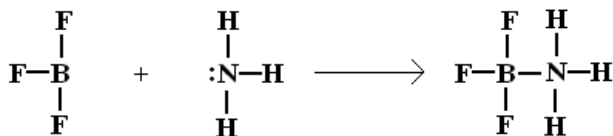
Lewis Acid: electron acceptor

Lewis Base: electron donator

Some common Lewis Acid-Base reactions:

(1) Boron compounds (such as BH_3 and BF_3) with group five element compounds (usually nitrogen or phosphorus, such as NH_3 or PH_3)

ex. $\text{BF}_3 + \text{NH}_3 \rightarrow \text{BF}_3\text{NH}_3$



The electrons are accepted by the lewis acid, BF_3 and donated by the lewis base, NH_3 .

(Generally, the boron compound acts like the lewis acid and the group five element compound acts like the lewis base.)

(2) metallic oxide + water \rightarrow base

ex. $\text{BaO} + \text{H}_2\text{O} \rightarrow \text{Ba}(\text{OH})_2$

The metallic oxide is acting like a lewis base and water is acting like a lewis acid.

(3) non-metallic oxide + water \rightarrow acid

ex. $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$

The non-metallic oxide is acting like a lewis acid and water is acting like a lewis base.

(4) metallic oxide + non-metallic oxide \rightarrow salt

ex. $\text{BaO} + \text{CO}_2 \rightarrow \text{BaCO}_3$

The metallic oxide is acting like a lewis base and the non-metallic oxide is acting like a lewis acid

(5) formation of a complex ion.

ex. $\text{Al}^{3+} + 6\text{Cl}^- \rightarrow [\text{AlCl}_6]^{3-}$ ex. $\text{Cu}^{2+} + 4\text{NH}_3 \rightarrow [\text{Cu}(\text{NH}_3)_4]^{2+}$

The metals are acting like lewis acids and the ligands are acting like lewis bases

A comparison of the three Acid-Base Models		Acids	Bases
Arrhenius	Definition:	Dissociate to produce H ⁺ ions	Dissociate to produce OH ⁻ ions
	Example:	HCl	NaOH
Bronsted-Lowry	Definition:	proton donor	proton acceptor
	Example:	HC ₂ H ₃ O ₂	NH ₃
Lewis	Definition:	electron acceptor	electron donor
	Example:	non-metallic oxide	metallic oxide