Name ____ Date ____________________

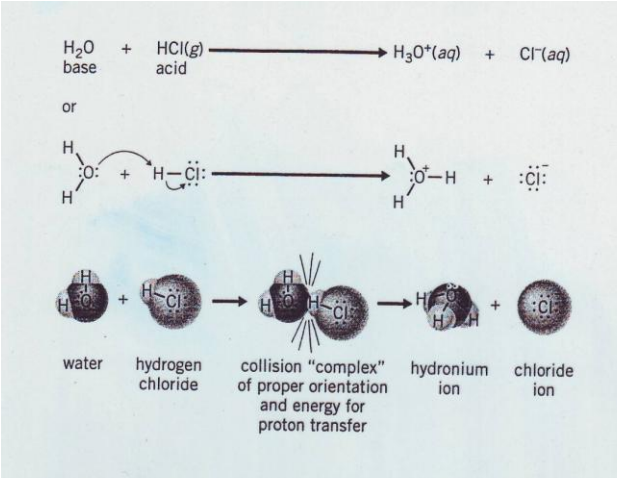
Acids & Bases

Properties

	Acid	Base
Taste	sour/tart	bitter
Feel	burn	slippery
Metal rxn	yes	no
Litmus	red	blue
Conductivity	yes	yes

Definitions

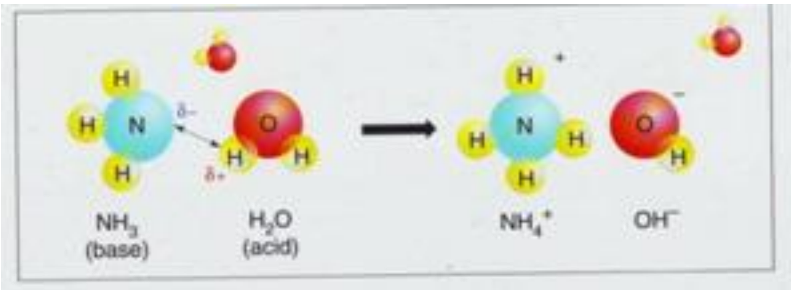
Arrhenius acid (1880's) –



In actuality the reaction occurs by the H^+ attaching itself to H_2O forming H_3O^+ , the hydronium ion. You may write either symbol, but you need to know the true form.

Some common every day acids are vinegar, citric, gastric (HCl 0.1 M)

Arrhenius base -



Where does the OH^- really come from?

Some common every day bases are soap, bleach, ammonia (NH_3), drano, detergent, milk of magnesia, antacids

Problem: The Arrhenius definition couldn't explain neutralization rxns like $NH_3 + HCl \rightarrow NH_4Cl$,
In this rxn, all that happens is that the H^+ from HCl moves to NH_3 & the opposite ions attract. In effect a PROTON is DONATED

Bronsted – Lowry acid

Bronsted – Lowry base

Thus according to Bronsted-Lowry – acid base reactions are proton transfer reactions.
So rxns w/ water are also acid base rxns in which H_2O acts as acid or base. *Acid Pictures powerpoint slide 3*

Eg. 1 Identify the acid and the base in the reaction $HCHO_2 + H_2O \leftrightarrow H_3O^+ + CHO_2^-$

General form of an acid rxn $\mathbf{HA + H_2O \leftrightarrow A^- + H_3O^+}$

General form of an base rxn $\mathbf{B + H_2O \leftrightarrow OH^- + BH^+}$

Eg. 2 Identify the acid & base in the following rxns. All of these are weak acids & bases.

a $HOCl + H_2O \leftrightarrow H_3O^+ + OCl^-$

b $HSO_4^- + PO_4^{3-} \leftrightarrow SO_4^{2-} + HPO_4^{2-}$

c $H_2O + NH_3 \leftrightarrow NH_4^+ + OH^-$

d $NaCO_3 + H_2O \leftrightarrow HCO_3^- + OH^-$

e $CaO(s) + H_2O \leftrightarrow Ca^{2+} + 2OH^-$

The best way to look at acid base rxns is to view them as an equilibrium. Identify the acid & base in each rxn, forward & reverse. You can identify substances that are related to each other by donating or accepting a proton as a Conjugate acid-base pair. Any Bronsted-Lowry equilibium has 2 acid-base pairs. An easy way to identify the acid member of the pair is that it has one more H^+ than the base member.

Eg 3 Write the conjugate base for nitric acid & hydrogen sulfate ion.

Eg. 4 What are the conjugate acids of OH^- & PO_4^{3-}?

Amphoteric, amphiprotic -

Strength

Autoionization of water –

Write K_c for autoionization of H_2O

K_w =

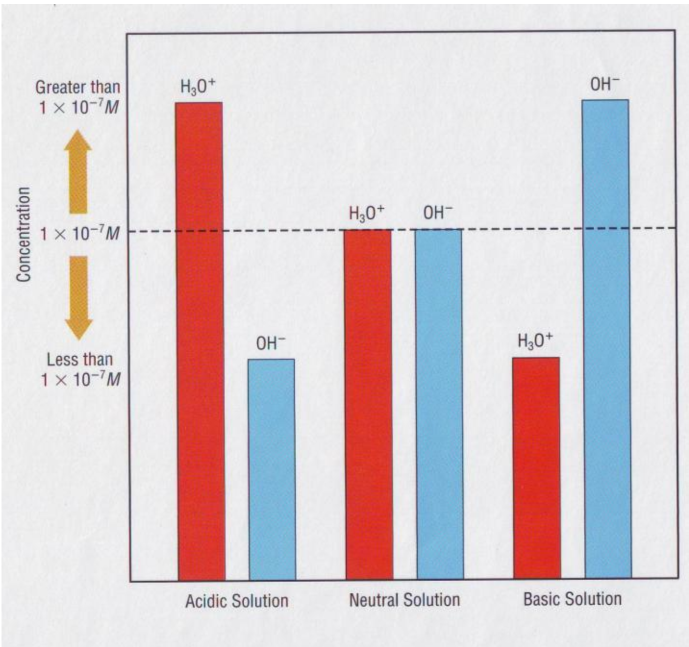
This value holds true for 25°C. Just like other equilibriums K_w is affected by temperature

K_w is also referred to as the ion-product constant of water.

A neutral solution is

An acid solution has

A basic solution has



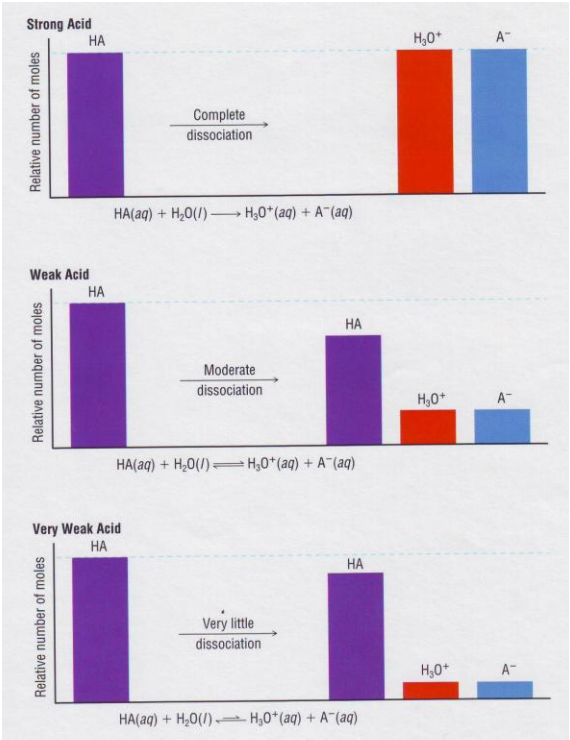
Acid Strength

What determines the strength of an acid or base?

Strong acids & bases -

Weak acids & bases -

TABLE 14.1 in the book provides properties of weak vs strong acids



The value of K gives an indication of the relative strength of an acid. The smaller the K_a value the weaker the acid.

Eg. 7 Place the following in order of acid strength from weakest to strongest:

a) HOCl, HI, NH_4^+, $HC_7H_5O_2$

b) HF, H_2SO_4, H_3BO_3, HSO_4

Using K_w to calculate $[H^+]$ & $[OH^-]$

Eg. 5 In a sample of blood at 25°C the $[H^+] = 4.6 \times 10^{-8}$ M. Find $[OH^-]$ and decide if the sample is acidic, basic, or neutral.

Eg. 6 An aqueous solution of sodium bicarbonate has $[OH^-] = 7.8 \times 10^{-6}$. What is $[H^+]$? Acidic, basic, or neutral?

pH scale

Because you take the negative log, pH decreases as $[H^+]$ increases.
Because it's in a log scale for every 1 unit change in pH the $[H^+]$ increases by 10 times. If the pH changes by 2 then the $[H^+]$ changes by ________.

Write the expression for K_w.

pH + pOH = 14.00

Remember how sig figs work w/ logs.

Eg. 1 Calculate the pOH & $[H^+]$ of a solution w/ a pH of 11.68.

Eg. 1b Calculate the pH & $[OH^-]$ of a solution with a pOH of 3.67

Calculating pH of strong acids

Here we will start to work on a problem solving strategy we will use for the ENTIRE unit.

Eg. 2 Calculate the pH & $[OH^-]$ of a 0.341 M HCl solution.

With all acid base problems the following steps will help the solution process

1) Identify the major species in the beaker & find their K values

2) Decide which species is contributing significantly to the $[H^+]$ of the solution & write the equation for it.

3) Determine $[H^+]$ using the significant species.

4) Calculate pH.

Eg. 3 Calculate the pH & $[OH^-]$ of a 2.6×10^{-3} M HNO_3 solution.

1)

2)

3)

4)

Eg. 4 Calculate the $[H^+]$ of a solution of $HClO_4$ solution that has a pH of 5.78.

Eg. 5 Calculate the pH & $[H^+]$ of a 5.3×10^{-10} M HBr solution.

Eg. 6 Calculate the pH of a solution made of 7.2×10^{-5} M $HClO_4$ and 3.4×10^{-4} HI.

Strong Bases

Strong Bases:

Rxn w/ water

Calculations w/ strong bases

Similar to calculations with a strong acid but w/ an added step.

Eg. 2 Calculate pH of a solution of 0.20M NaOH.

Eg. 3 Determine the pH of a solution of 0.30 M $Ba(OH)_2$

Eg. 4 The pH of a solution of $Mg(OH)_2$ is 13.65. Determine the concentration of $Mg(OH)_2$.

Weak acids

Weak acids do not

Equation for weak acid $HA + H_2O \leftrightarrow H_3O^+ + A^-$

Write the K expression for a weak acid.

K_a = acid ionization constant

Review meaning for K_a values

K_a & pH Calculations for weak acids

A similar process is used to calculate pH of weak acids, but the addition of solving an equilibrium problem is added. *Keep in mind that many weak acids have very small K_a values, so assumptions can be made. The 5% rule for the assumptions made for small K values works because the K_a values themselves are accurate to within 5%.*

Eg. 1 A 1.0 M solution of acetic acid has a K_a = to 1.8×10^{-5}. Calculate the pH.

1) major species (what's in the beaker)

2) significant contributor – largest K is the most significant contributor

3) write the equilibrium expression for the significant contributor

4) using the given info set up the equilibrium problem

5) determine concentrations at equilibrium

6) Use K_a to solve for $[H^+]$. Remember assumptions if they apply

7) check 5% rule

8) calculate pH

Eg. 2 Lactic acid ($HC_3H_5O_3$) which is present in sour milk, also gives sauerkraut its tartness. Calculate the pH of a 0.100 M solution of lactic acid at 25°C.

1) major species (what's in the beaker?)

2) significant contributor: largest K is the most significant contributor

3) write the equilibrium expression for the significant contributor

4) using the given info set up the equilibrium problem

5) determine concentrations at equilibrium

6) Use K_a to solve for $[H^+]$. Remember assumptions if they apply

7) check 5% rule

8) calculate pH

Eg. 3 Calculate the pH of a solution made of 0.60 M acetic acid & 4.3 M boric acid.

Calculate the concentration of the borate ion.

Percent dissociation
Formula

The percent of dissociated acid in a solution increases as the solution [ ] decreases.
Why? LeChatelier:

Kinetics:

Mathematics:

Note: As a solution becomes more dilute, it is less likely that the 5% rule will apply due to the fact that % dissociation increases as [HA] decreases.

Eg. 4 Calculate the % dissociation of acetic acid in eg. 1 & eg. 3. Compare.

Eg. 5 An 0.400 M acid solution is 4.50% dissociated. Calculate the K_a for this acid and the pH.