Objectives:

Upon completion of this unit I am pretty sure that the student will be able to :

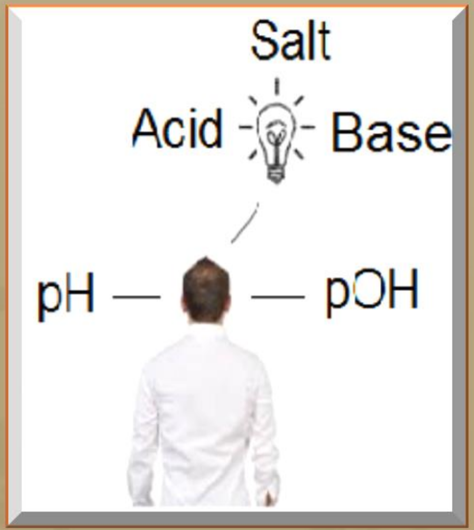
1-Realize the relationship between pH and pOH in the aqueous solution .

2- Distinguish between strong acids and weak acids and between strong base and weak base .

3- Calculate the pH of the solution of all sorts of acids , bases and salts .

4- Understand the concept of the hydrolysis of salts .

5- Have some ideas on the acidity of some common foods , drinks , and various other stuffs that you use in your daily life .



Introdution:

What exactly is pH, why is it so important, How it can be calculated ? This unit has all the answers. Acidic and basic are two extremes that describe chemicals, just like hot and cold are two extremes that describe temperature. Mixing acids and bases can cancel out their extreme effects; much like mixing hot and cold water . A substance that is neither acidic nor basic is neutral.

The pH scale measures how acidic or basic a substance is. It ranges from 0 to 14. A pH of 7 is neutral. A pH less than 7 is acidic, and a pH greater than 7 is basic. Pure water is neutral, with a pH of 7.0. When chemicals are mixed with water, the mixture can become either acidic or basic. Vinegar and lemon juice are acidic substances, while laundry detergents and ammonia are basic.

pH is a vital component of a plant's surroundings. As the pH changes, the ability of a plant to absorb nutrients changes also . Many biological reactions that occurs in the human body are pH dependent .

Ion- product Constant for Water(K_w)

Water partially dissociate according to the equation :

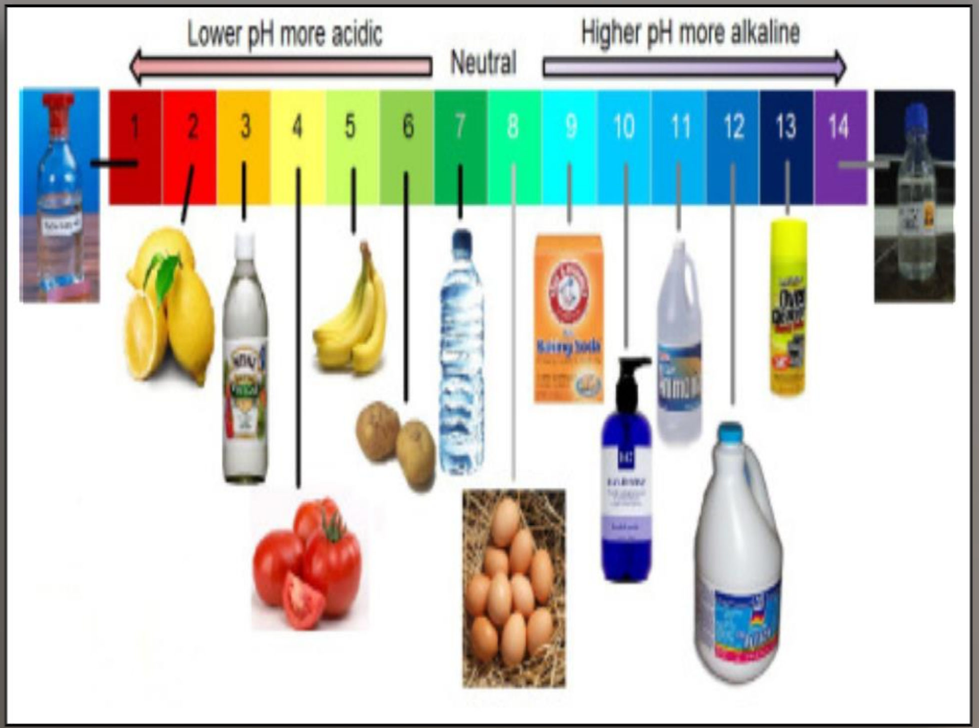
$$H_2O \leftrightarrow H^+ + OH^-$$

At room temperature , Water Dissociation constant (K_{eq}) is equal to :

$$K_{eq} = 1.82 X 10^{-16} = \frac{[H^+][OH^-]}{[H_2O]}$$

$$[H^+][OH^-] = 1.82X10^{-16} \quad X \; [H_2O] = K_w$$

Where K_w is the ion product constant for water . Since the density of water is 1 g / mL at room temperature , the molar concentration of water can be easily calculated :



Ion- product Constant for Water(K_w)

$$[\,H^+\,][\,OH^-\,] = 1.82\times10^{-16} \times [H_2O] = K_w$$

Where K_w is the ion product constant for water . Since the density of water is 1 g / mL at room temperature , the molar concentration of water can be easily calculated :

$$[H_2O] = \frac{(g\ H_2O\ /\ L)}{mw_{\ H_2O}} = \frac{1000\ \ (g\,/\,L)}{18} = 55.6\ M$$

VIDEO
Water Ionization

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Ion- product Constant for Water(K_w)

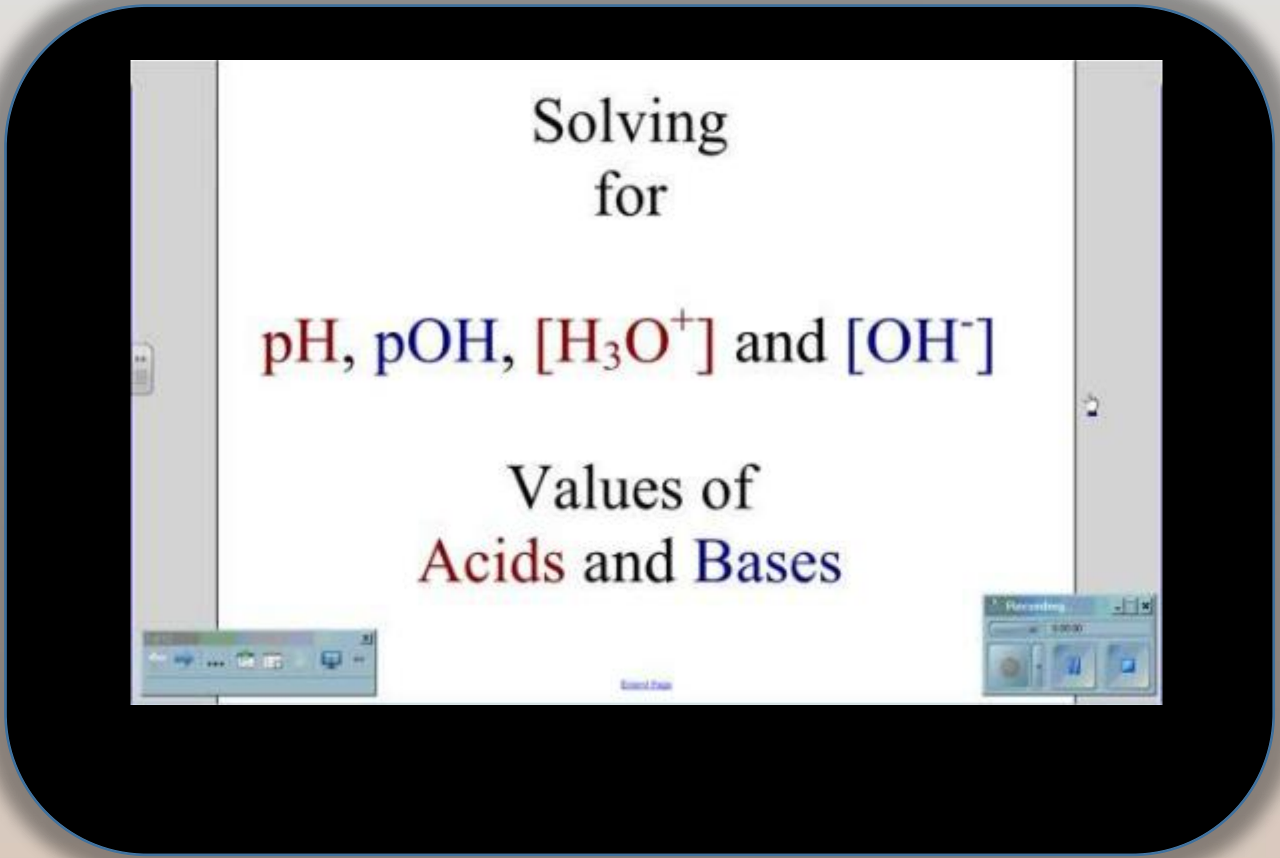
Substituting for $[H_2O]$:

$$[H^+][OH^-] = 1.82X10^{-16} \quad X \quad 55.6$$
$$= 1X10^{-14} = K_w$$

$$[H^+][OH^-] = 1X10^{-14} = K_w$$

Multiplying both sides of this equation by - log :

$$pH + pOH = 14 = pK_w$$



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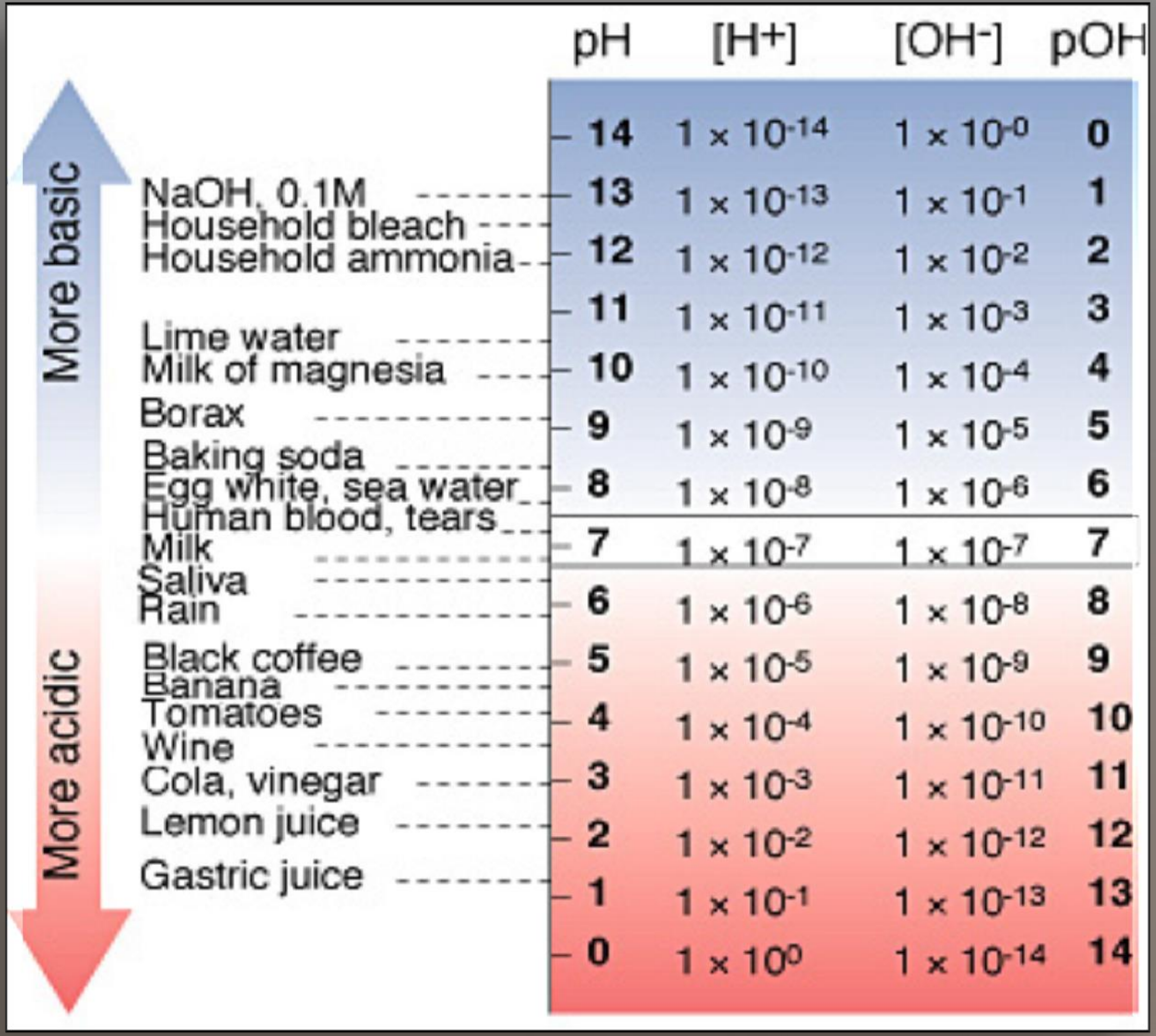
Ion- product Constant for Water(K_w)

If we add acid to the water then $[H^+] > [OH^-]$ but the product of $[H^+]$ $[OH^-]$ will stay constant $1x10^{-14}$. So from the last two equations we can calculate $[H^+]$, pH , $[OH^-]$ and pOH by knowing any one of these as shown in the table on your right .

The Table on your right shows The relationship between $[H^+]$, pH , $[OH^-]$ and pOH in the aqueous solution .

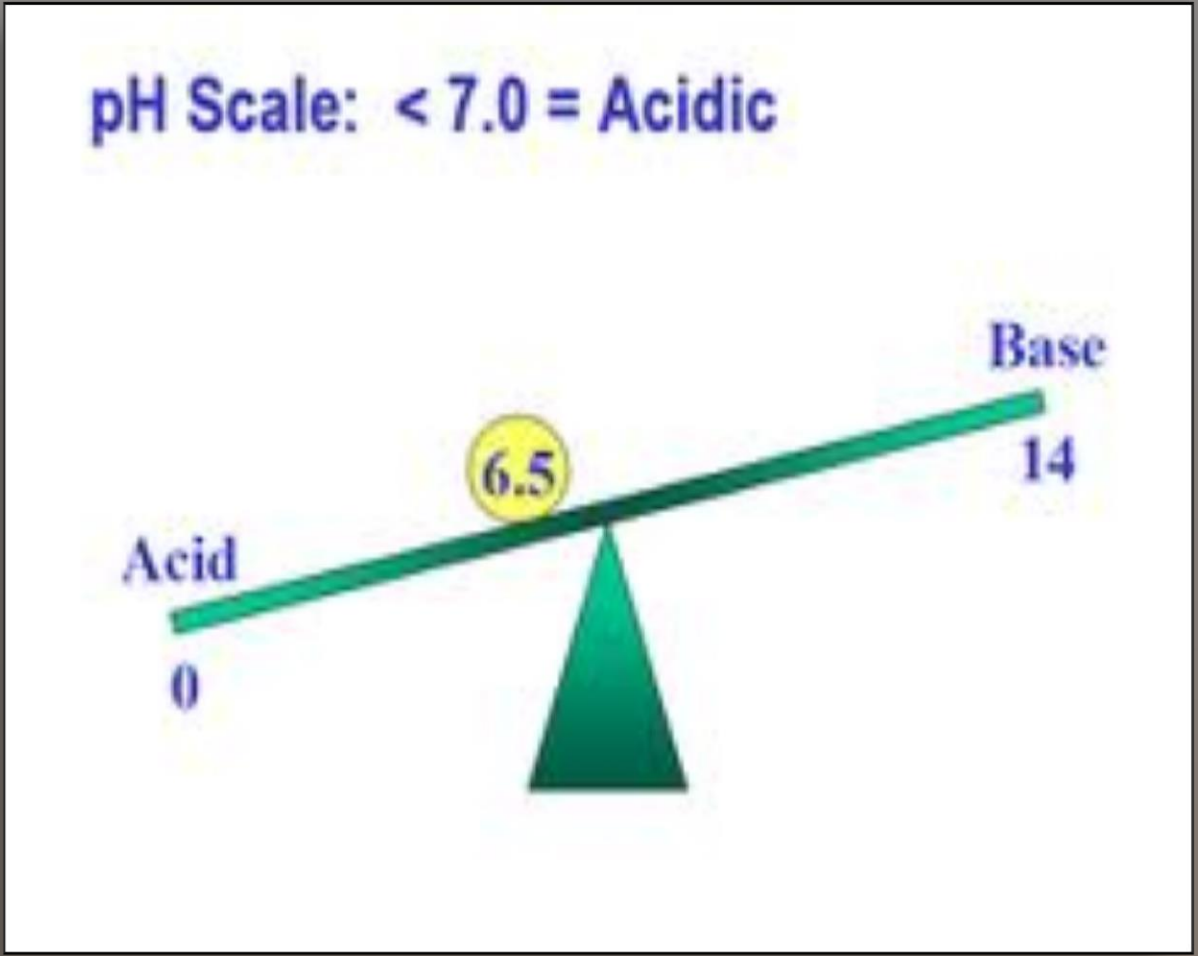


pH	[H+]	[OH-]	pOH
14	1×10^{-14}	1×10^{0}	0
13	1×10^{-13}	1×10^{-1}	1
12	1×10^{-12}	1×10^{-2}	2
11	1×10^{-11}	1×10^{-3}	3
10	1×10^{-10}	1×10^{-4}	4
9	1×10^{-9}	1×10^{-5}	5
8	1×10^{-8}	1×10^{-6}	6
7	1×10^{-7}	1×10^{-7}	7
6	1×10^{-6}	1×10^{-8}	8
5	1×10^{-5}	1×10^{-9}	9
4	1×10^{-4}	1×10^{-10}	10
3	1×10^{-3}	1×10^{-11}	11
2	1×10^{-2}	1×10^{-12}	12
1	1×10^{-1}	1×10^{-13}	13
0	1×10^{0}	1×10^{-14}	14

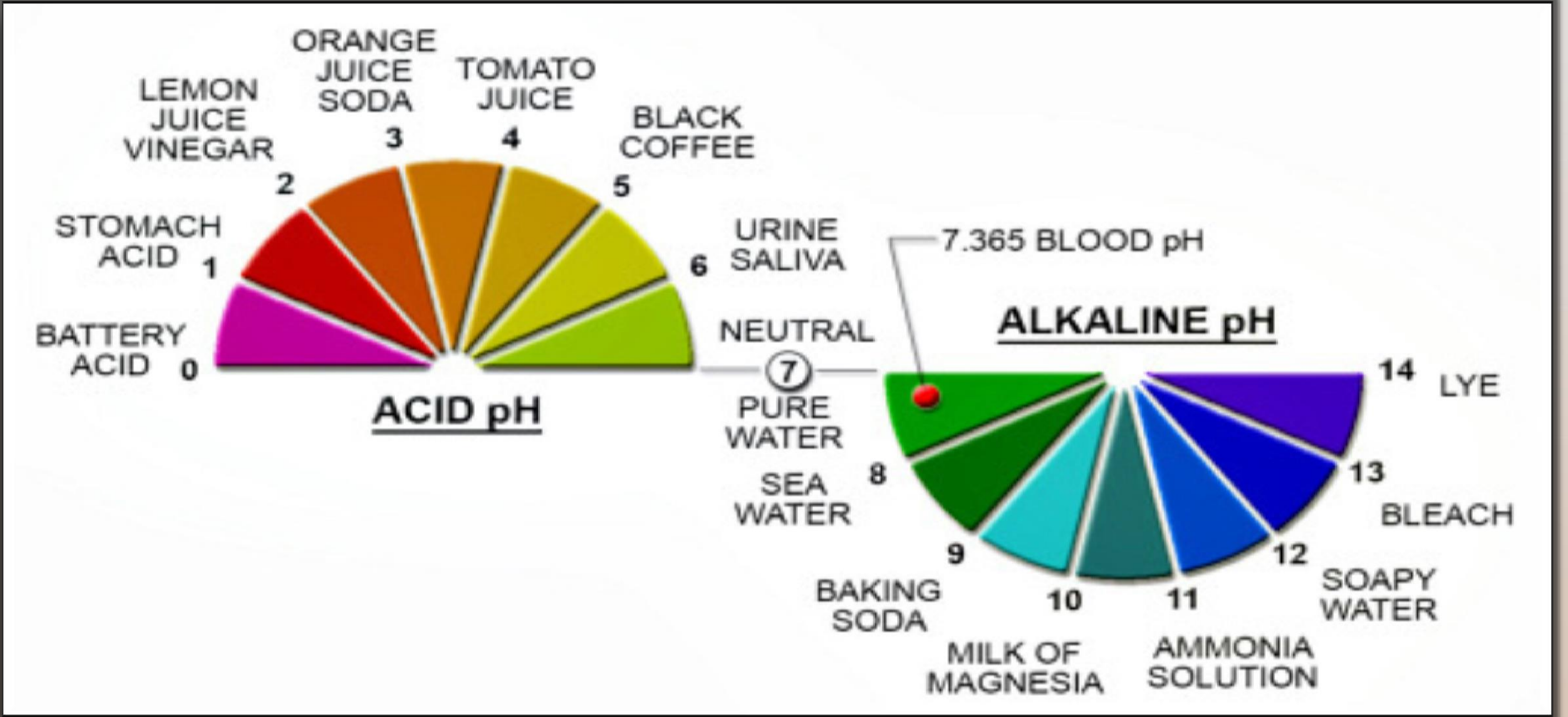
Strong Acid Solution

The strength of the acid is determined by how far the equilibrium lies to the right. Qualitatively, this may be judged by the K_a of the acid. A large K_a indicates a strong acid; a small K_a indicates a weak acid. Strong acids, such as HCl, have K_a values in the vicinity of infinity.

$$K_a = \frac{[H^+][A^-]}{0} = undefined$$



Strong Acid Solution



Strong Acid Solution

This implies that the dissociation of HCl is virtually complete,

HCl	→ H^+	+ Cl^-	
C_a	0	0	Before dissociation
0	C_a	C_a	After complete dissociation

and the equilibrium lies completely to the right, therefore, the concentration of the acid equals the concentration of H^+ ions produced. For instance, a 0.01 M HCl solution will completely dissociate into 0.01 M H^+ and 0.01 M Cl^-. The concentration of HCl after "equilibrium" will be zero! Analogously, strong bases, such as NaOH, will dissociate completely. The concentration of OH^- in solution will be equal to the concentration of the strong base.

Strong Acid Solution

Strong Acid Solution

A typical strong acid problem might be: What is the pH of a C_a M HCl solution? Since HCl is a strong acid, the H^+ ion concentration will be equal to the HCl concentration:

$$[H^+] = C_a \ (M)$$

The pH can be found by taking the negative log of the H^+ ion concentration :

$$pH = -\log[H^+] = -\log C_a$$

VIDEO
pH and pOH

VIDEO
My Channel

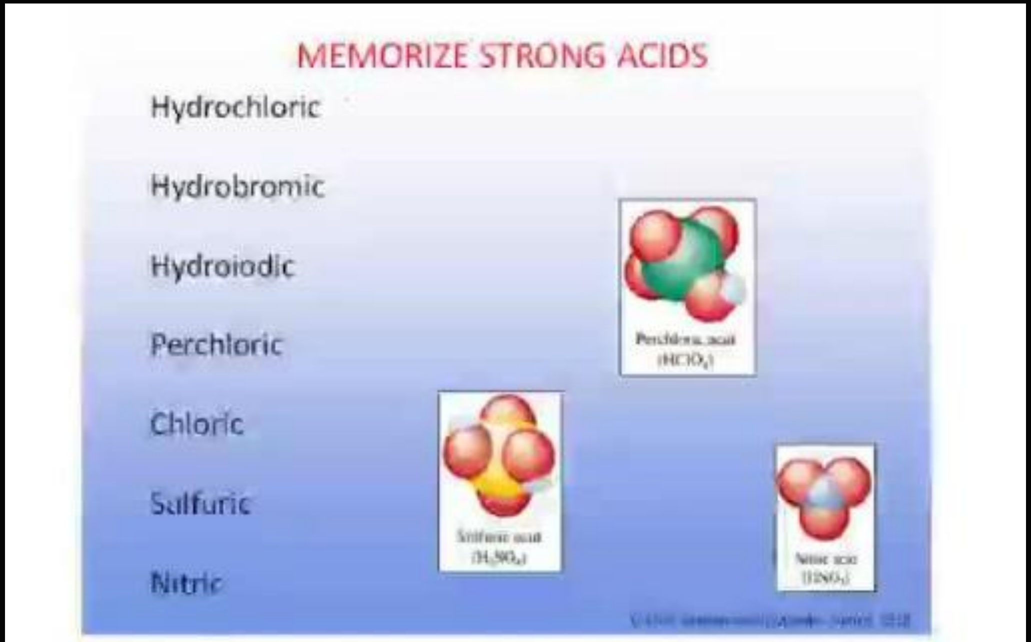
Strong Base Solution

A typical strong base problem might be: What is the pH of a C_b M NaOH solution ? Since NaOH is a strong base, the hydroxide ion concentration will be equal to the NaOH concentration:

$$[OH^-] = C_b \text{ M}$$

The pH can be found by first finding the pOH by taking the negative log of the hydroxide ion concentration, and then converting the pH to pOH. To find the pOH:

$$pOH = -\log [OH^-] = -\log C_b$$



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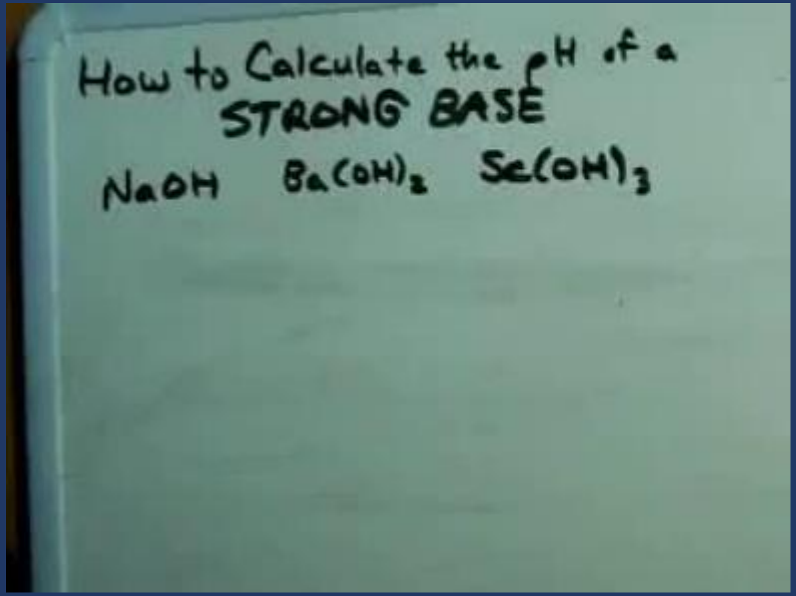
Strong Base Solution

$$pOH = -\log[OH^-] = -\log(0.010) = 2.00$$

The pH can then be calculated from the equation :

$$pH + pOH = 14:$$

The above rules apply to most strong acids and strong bases exceptions are for example $Ca(OH)_2$ where one mole of $Ca(OH)_2$ will produce two moles of OH^- , therefore , instead of C_b we use $2C_b$. Also H_2SO_4 is a special case (see right video)



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Examples for the calculatins of pH of strong acids and bases

Example : Calculate $[OH^-]$ in a 0.01 M solution of HCl ?

Solution :

$pH = -\log C_a = -\log 0.01 = 2$

$pOH = 14 - 2 = 12$

$[OH^-] = 1X10^{-12}$

Or

$[H^+][OH^-] = 0.01\ [OH^-] = 1X10^{-14}$

$[OH^-] = 1X10^{-12}$

Example : Calculate $[H^+]$ in a 0.005 M solution of $Ba(OH)_2$?

Examples for the calculatins of pH of strong acids and bases

Solution :

pOH = - log 2 X 0.005 = 2

pH = 14 – 2 = 12

$[H^+]$ = 1 X 10^{-12} M

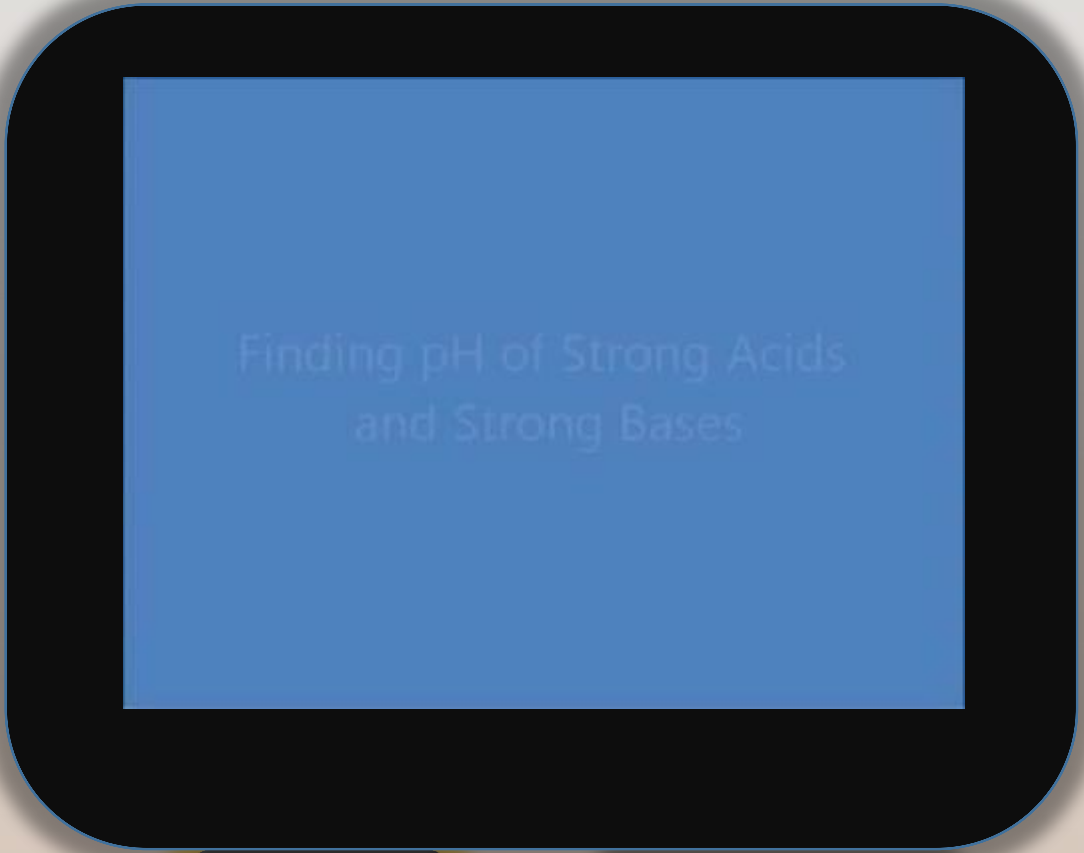
Or

$[H^+][OH^-] = [H^+] \times 2 \times 0.005 = 1 \times 10^{-14}$

$[H^+]$ = 1X10^{-12} M

Example : Calculate the pH for the following mixture :

100 ml 0.02 M HCl + 100 ml 0.03 M NaOH



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Examples for the calculatins of pH of strong acids and bases

Solution :

	$NaOH$ +	HCl →	$NaCl$ +	H_2O
Before reaction (mmoles)	0.03X100=3	0.02X100=2	0	0
After reaction (mmoles)	1	0	2	2

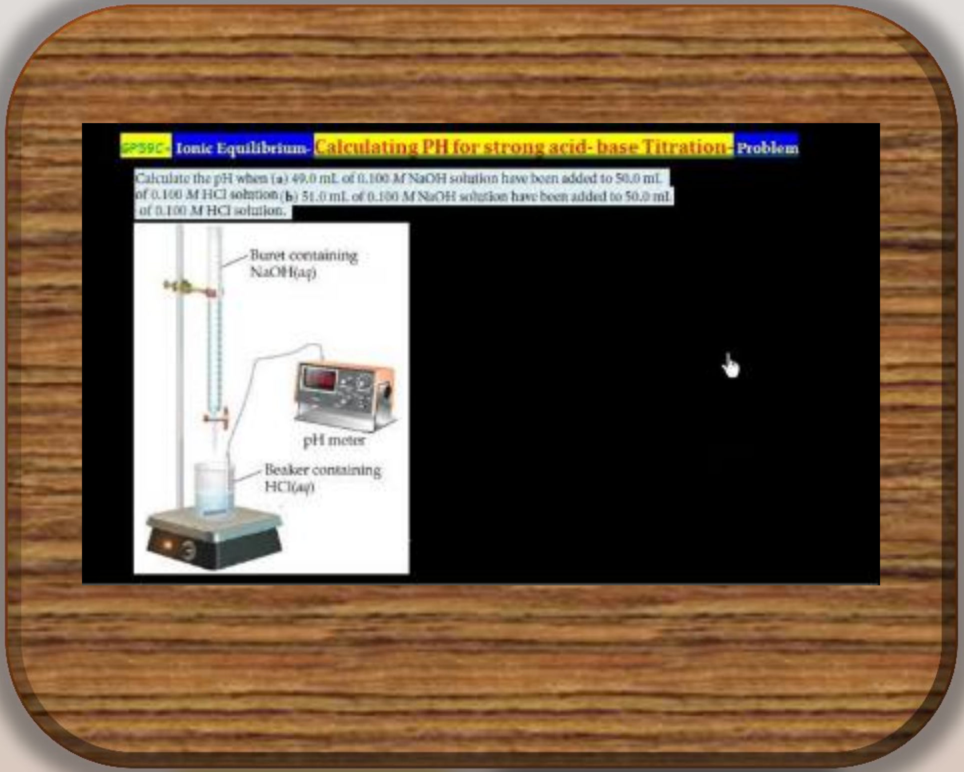
The limiting reactant is HCl and the excess reactant is NaOH , Therefore , the solution is basic :

$$pOH = -\log \frac{1}{200} = 2.3$$

$$pH = 14 - 2.3 = 11.7$$

Example : Calculate the pH of the following mixture :

2 ml HCl (pH = 3) + 3 ml NaOH (pH = 10)



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Examples for the calculatins of pH of strong acids and bases

Solution :

For the HCl : pH = 3

$[H^+] = 1X10^{-3}$ M

no. of mmoles of H^+ = $1X10^{-3}$ M X 2 (ml)

= $2X10^{-3}$

for the NaOH :

pH = 10 ∴ pOH = 4

$[OH^-] = 1X10^{-4}$ M

no. of mmoles of OH^- = $1X10^{-4}$ M X 3 (ml)

= $3X10^{-4}$ = $0.3 X10^{-3}$

Examples for the calculatins of pH of strong acids and bases

$$NaOH + HCl \rightarrow NaCl + H_2O$$

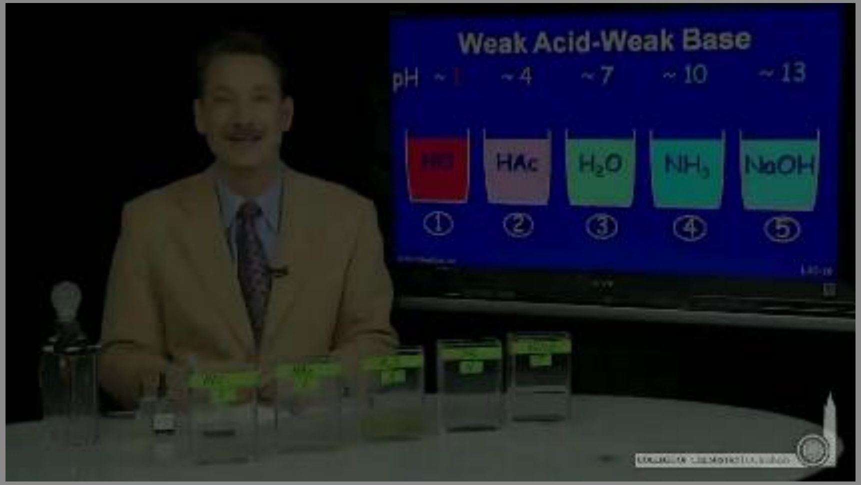
	NaOH	HCl	NaCl	H_2O
Before reaction (mmoles)	$0.3 X10^{-3}$	$2X10^{-3}$	0	0
After reaction (mmoles)	0	$1.7 X 10^{-3}$	$0.3 X10^{-3}$	$0.3 X10^{-3}$

Note that : no. mmoles H^+ (excess) $= 2x10^{-3} - 0.3x10^{-3} = 1.7x10^{-3}$

$$pH = -\log \frac{1.7x10^{-3}}{5} = 3.5$$

Weak Acids and Weak Bases

Weak acids and weak bases do not dissociate completely. An equilibrium exists between the weak acid, water, H^+, and the anion of the weak acid. The equilibrium lies to the left hand side of the equation, indicating that not much H^+ is being produced. The fact that very little H^+ is being produced is the definition of a weak acid. The K_a for a weak acid is small, usually a number less than 1.The less value of K_a or K_b the weaker the acid or the base respectively .



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Weak Acids and Weak Bases

There are three types of problems encountered with weak acids or bases: dissociation, buffers or hydrolysis. We'll look at each type in detail.

In this type of problem, you will be asked to find the H^+ ion concentration and/or the pH of a weak acid whose initial concentration is known. A typical problem may be: what is the pH of C_a M solution of HA (K_a = ……) ?

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pH of weak acid

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Weak Acids and Weak Bases

	HA	$\leftrightarrow$	H^+	+	A^-
before dissociation	C_a		0		0
after dissociation	$C_a - x$		x		x

The equilibrium may be expressed mathematically by setting the K_a equal to the mass action expression:

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(x)^2}{C_a - (x)} \approx \frac{[H^+]^2}{C_a}$$

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Weak Acids and Weak Bases

The solution for x becomes simplified because the x can be neglected as we discussed in the equilibrium unit . This x can be neglected because it will be negligibly small compared to the concentration, C_a M. To determine whether x is negligible C_a must be more or equal to K_a X100 .This simplifies the equation as we saw above . Multiplying both sides by C_a yields:

$$[H^+]^2 = K_a \; X \; C_a \quad \therefore [H^+] = \sqrt{K_a \, X \, C_a}$$

$$Multipling\ \ both\, sides\ \ by\ -\log\ \ we\, get\ :$$

$$pH = -\log \sqrt{K_a \; X \; C_a}$$

Weak Acids and Weak Bases

Weak Acids and Weak Bases

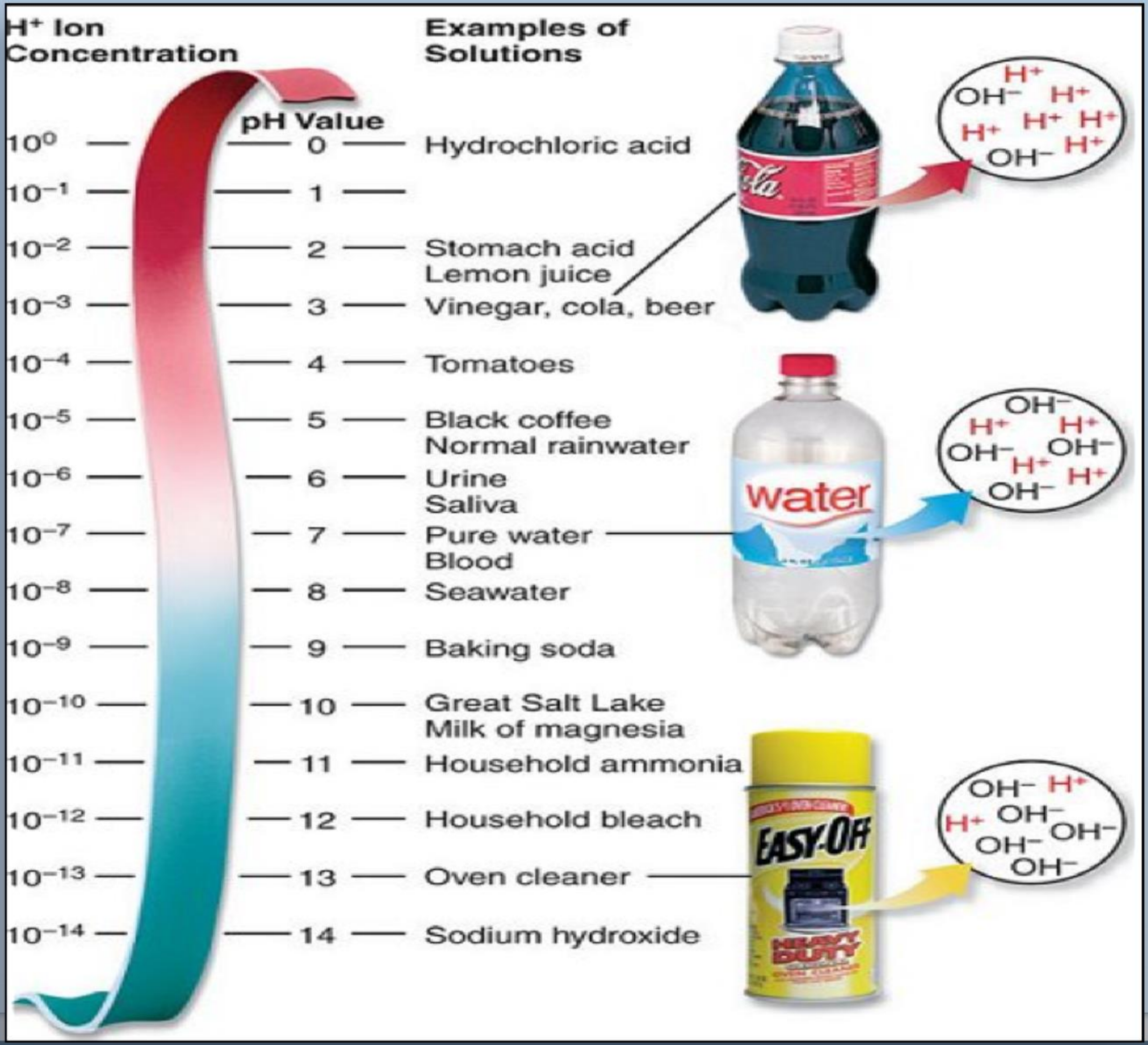
A typical weak base problem may read: What is the hydroxide ion concentration and pH of a C_b M solution of NH_3, $K_b = 1.8 \times 10^{-5}$?

Again, note that K_b is small. We will follow the same format as we used for weak acid solutions ,bearing in mind that :

$$C_b \geq K_b \times 100$$

and we will end up with the following general equation :

$$pOH = -\log\sqrt{K_b.C_b}$$



Weak Acids and Weak Bases

Then subtract the pOH from 14 to find the pH:

$$pH = 14.00 - pOH$$

A typical weak base problem may read: What is the hydroxide ion concentration and pH of a C_b M solution of NH_3, $K_b = 1.8 \times 10^{-5}$?

The following tables show some common weak acids and weak bases .

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Weak Acids and Weak Bases

Some common weak acids .

Name	Formula
Acetic acid	$HC_2H_3O_2$
Phosphoric acid	H_3PO_4
Carbonic acid	H_2CO_3
Hydrofluoric acid	HF
Fluosilicic acid	H_2SiF_6

Weak Acids and Weak Bases

Some common weak bases .

Name	Formula
Ammonia	NH_3
Magnesium hydroxide	$Mg(OH)_2$
Aluminum hydroxide	$Al(OH)_3$
Lime	CaO
Sodium silicate	Na_2SiO_2
Soda ash	Na_2CO_3

Hydrolysis (Salts of Weak Acids or Weak Bases)

What are salts ? Salts are the product of an acid base neutralization. There are four possible acid base reactions that produce salts. They are the reaction of :

1) A strong acid with a strong base.

$HCl + NaOH \rightarrow Na^+ + Cl^- + H_2O$

2) A weak acid with a strong base.

$CH_3COOH + NaOH \rightarrow Na^+ + CH_3COOH^- + H_2O$

Hydrolysis (Salts of Weak Acids or Weak Bases)

3) A weak base with a strong acid.

$NH_3 \quad + \quad HCl \quad \text{-->} \quad NH_4^+ \quad + \quad Cl^-$

4) A weak acid with a weak base.

$CH_3COOH + NH_3 \text{ --> } NH_4^+ + CH_3COOH$

The salts produced in the above four types have a characteristic pH range in water solution by themselves:

VIDEO
Salt Hydrolysis
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Hydrolysis (Salts of Weak Acids or Weak Bases)

1) A salt of a strong acid and a strong base will produce a solution with pH = 7 e.g NaCl salt because neither Na^+ nor Cl^- affect the equilibrium of water :

$$H_2O \leftrightarrow H^+ + OH^-$$

That means , Na^+ will not react with OH^- to form NaOH because NaOH is a strong base which is completely dissociate in aqueous solution and can not exist in the form of NaOH .

Hydrolysis (Salts of Weak Acids or Weak Bases)

Likewise Cl^- , will not react with H^+ to form the strong acid HCl ,therefore , when we dissolve NaCl in water the solution will stay neutral and one can say neither Na^+ nor Cl^- hydrolyze .

Ions of Neutral Salts			
Cations			
Na^+	K^+	Rb^+	Cs^+
Mg^{2+}	Ca^{2+}	Sr^{2+}	Ba^{2+}
Anions			
Cl^-	Br^-	I^-,	
ClO_4^-	BrO_4^-	ClO_3^-	NO_3^-

2) A salt of a weak acid and a strong base will produce a basic solution e.g CH_3COONa because CH_3COO^- will react with H^+ to form the weak acid CH_3COOH which can exist in water in the molecular formula :

$$CH_3COO^- + H_2O \leftrightarrow CH_3COOH + OH^-$$

So , $[OH^-]$ will be more than $[H^+]$ in the aqueous solution of the salt .

Hydrolysis (Salts of Weak Acids or Weak Bases)

The dissociation constant of the above reaction is :

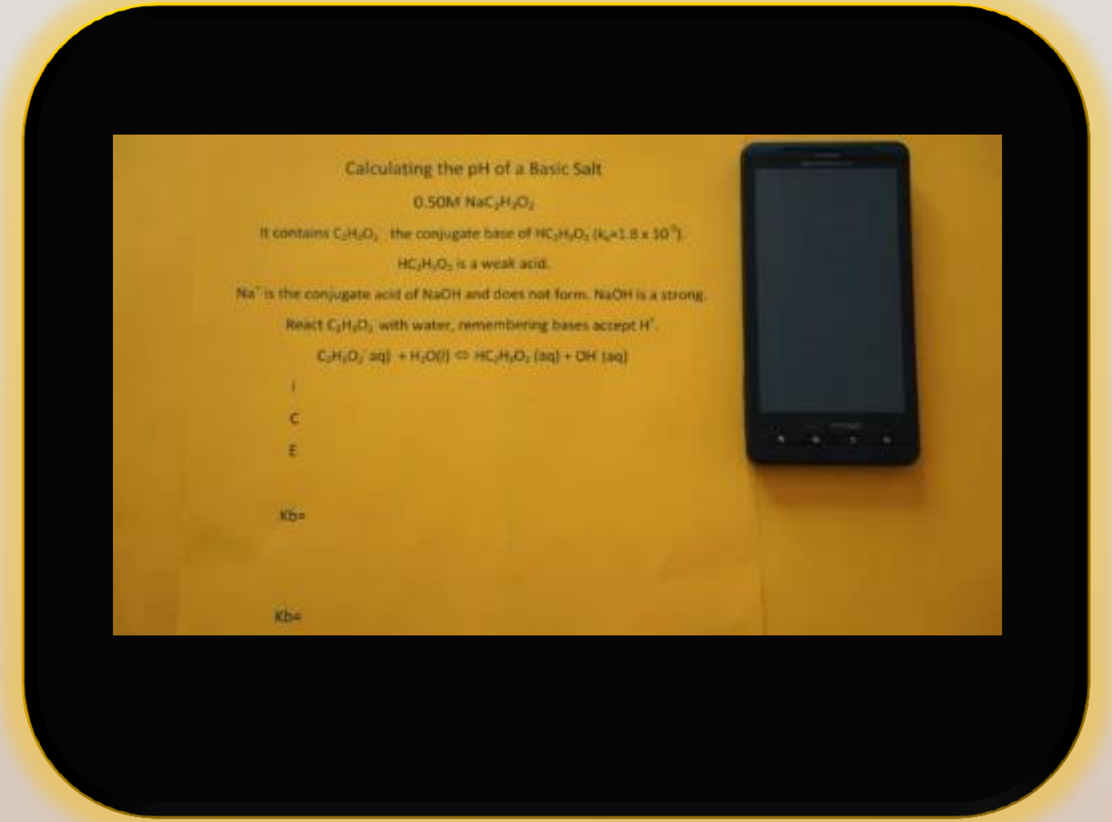
$$K_{b^-} = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]} = \frac{[OH^-]^2}{[CH_3COO^-]} = \frac{[OH^-]^2}{C_s}$$

Note that the salt dissociate completely, so :

Salt concentration (C_s) = $[CH_3COO^-]$

We can treat this kind of salt as a weak base and calculate its pH from the following equation :

$$pOH = -\log \sqrt{K_{b^-} . C_s}$$



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Hydrolysis (Salts of Weak Acids or Weak Bases)

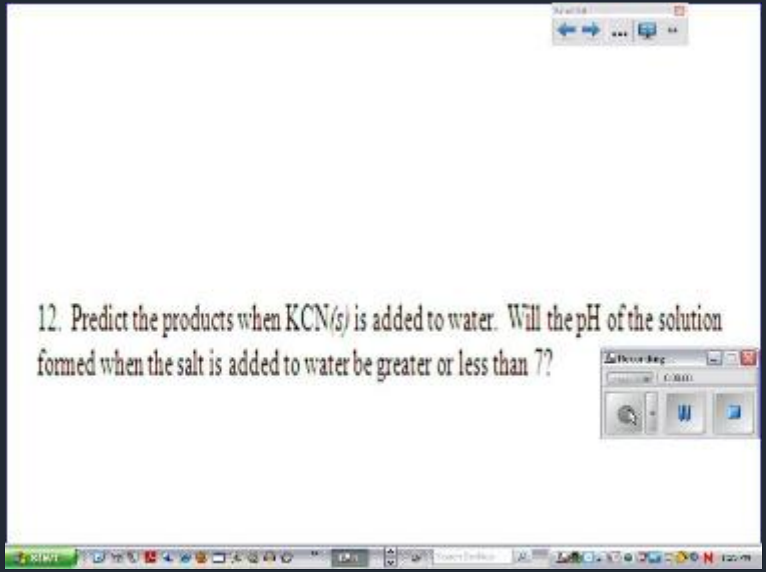
The value of K_{b^-} is not available but we can get it from the dissociation constant of the mother weak acid thus :

$$CH_3COOH \leftrightarrow CH_3COO^- + H^+$$

$$K_a = \frac{[CH_3COO^-][H^+]}{[CH_3COOH]}$$

$$Ka \times K_{b^-} = Kw$$

$$K_{b^-} = \frac{K_w}{K_a}$$



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Hydrolysis (Salts of Weak Acids or Weak Bases)

Substituting for K_{b^-} :

$$pOH = -\log \sqrt{K_{b^-} . C_s}$$

$$\boxed{pOH = -\log \sqrt{\frac{K_w . C_s}{K_a}} \quad \therefore \quad pH = 14 - pOH}$$

From this equation we can tell that the weaker the acid (the smaller K_a) the more basic will be its salt solution (pOH is directly proportional to $1/K_a$) .

Hydrolysis (Salts of Weak Acids or Weak Bases)

The last equation is used to calculate the pH of the solution of a salt of weak acid and strong base .Note that Na^+ from the salt does not hydrolyze as we mentioned above . The table on the right shows some ions that behave exactly like CH_3COO^-

3) A salt of a weak base and a strong acid e.g NH_4Cl will produce an acidic solution with pH less than 7. We can repeat the discussion above concerning the salt of a weak acid and strong base with the salt of a weak base and strong acid and we will end up with following equation which is used to calculate the pH of a solution of weak base :

Basic Ions			
F^-	CH_3COO^-	NO_2^-	PO_4^{3-}
CN^-	CO_3^{2-}	S^{2-}	

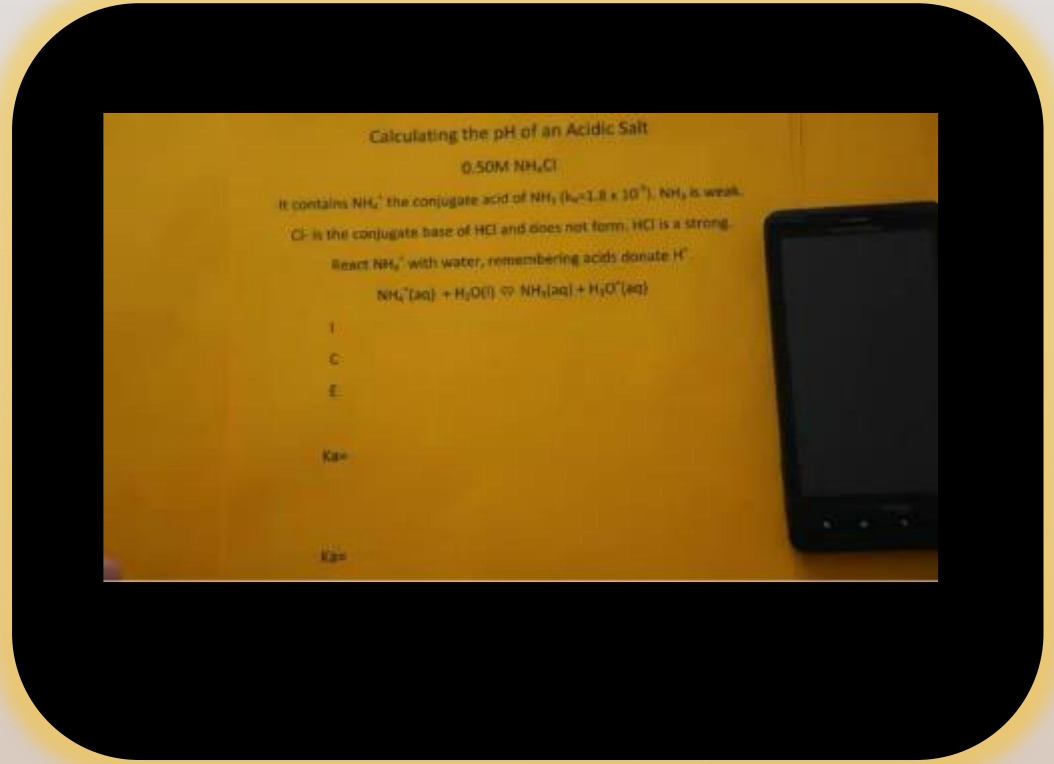
Acidic Ions			
NH_4^+	Al^{3+}	Pb^{2+}	Sn^{2+}
Transition metal ions			
HSO_4^-	$H_2PO_4^-$		

Hydrolysis (Salts of Weak Acids or Weak Bases)

$$pH = -\log \sqrt{\frac{K_w \,.\, C_s}{K_b}}$$

Where C_s is the concentration of the salt and K_b is the dissociation constant of the mother weak base. Also from this equation we can tell that the weaker the base (the smaller K_b) the more acidic will be its salt solution (pH is directly proportional to $1/K_b$) .

Notice, the mention of a strong acid or strong base will be usually omitted and the phrases "salt of a weak base" or " salt of a weak acid" will be used.



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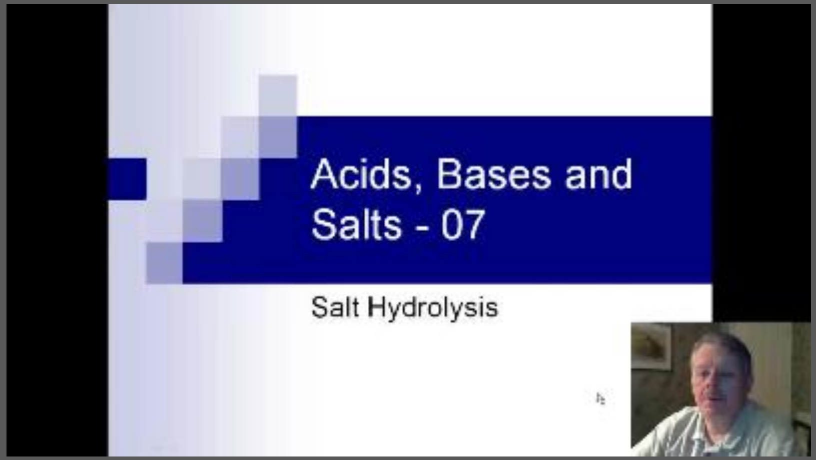
Hydrolysis (Salts of Weak Acids or Weak Bases)

Notice, the mention of a strong acid or strong base will be usually omitted and the phrases "salt of a weak base" or " salt of a weak acid" will be used.

4) A salt of a weak acid and a weak base produces a solution whose pH depends on the strengths of the acid and base which made the salt. Since we are not going to deal with this type of salt in this course (we will know why later) we will not discuss it farther . The second and third types of salts mentioned just above are very important in chemistry .



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Summary

In this unit we discussed the neutralization reaction and the relationship between pH , pOH , $[H^+]$ and $[OH^-]$ in the aqueous solution . The behavior of strong and weak acids and bases and their salts is explained . The calculations of the pH of all sorts of solutions , acids , bases and salts are investigated . The concept of the hydrolysis of salts has been cleared . We briefly pointed out about the acidity of some common foods , drinks , and various other stuffs that we use in our daily life.

Tutorial

Exercise 1 : Calculate the pH of 0.2 M solution of Na_2CO_3 ?

H_2CO_3 : $K_{a1} = 4.3 \times 10^{-7}$, $K_{a2} = 4.7 \times 10^{-11}$

Your answer :

Our answer next slide

Tutorial

Answer 1 :

$$pOH\ (Na_2CO_3) = -\log\sqrt{\frac{K_w\ x\ C_s}{K_{a_2}}} = -\log\sqrt{\frac{1X10^{-14}\ 0.2}{4.7\ X\,10^{-11}}} = 2.2$$

$$pH\ (Na_2CO_3) = 14 - 2.2 = 11.8$$

Tutorial

Exercise 2 : Calculate the pH of 0.5 M solution of NH_4Cl ?

$K_b\ (NH_3) = 1.75 \times 10^{-5}$

Your answer :

Our answer next slide

Tutorial

Answer 2 :

$$pH\ (NH_4Cl) = -\log \sqrt{\frac{K_w\ X\ C_s}{K_b}} = -\log \sqrt{\frac{1x10^{-14}\ X\ 0.5}{1.75\,X\,10^{-5}}} = 4.8$$

Tutorial

Exercise 3 : Calculate the pH of 0.1 M solution of Na_2HPO_4 ?

.

H_3PO_4 : $K_{a1} = 7.11 X 10^{-3}$, $K_{a2} = 6.34 X 10^{-8}$, $K_{a3} = 4.2 X 10^{-13}$

Your answer :

Our answer next slide

Tutorial

Answer 3 :

$$pH\ (Na_2HPO_4) = -\log \sqrt{K_{a_2} \cdot K_{a_3}} = -\log \sqrt{6.34\,X\,10^{-8}\ x\ 4.2\,X\ 10^{-13}}$$

$$= 9.8$$

Tutorial

Exercise 4 : Calculate the pH of 0.2 M solution of $Ba(OH)_2$?

Your answer :

Our answer next slide

Tutorial

Answer 4 :

$$pOH = -\log[Ba(OH)_2]\ X\ 2 \quad = -\ \log\ 0.2\ X\ 2 = 0.4$$

$$pH = 14 - 0.4 = 13.6$$

Tutorial

Exercise 5 : Calculate the pH of the solution resulting from adding 30 mL of 0.1 M HCl solution to 20 mL of 0.2 M $Ca(OH)_2$ solution ?

Your answer :

Our answer next slide

Tutorial

Answer 5 :

2 HCl +	$Ca(OH)_2 \rightarrow$	$CaCl_2$ +	$2H_2O$	
30x0.1 = 3.0	20x0.2= 4.0	0	0	(mmole)(I)
0	2.5	1.5	3.00	(mmoles) (C)

Suppose the reaction is complete . The final solution is basic because HCl is the limiting reactant and $Ca(OH)_2$ is the excess reactant .

$$pOH = -\log \frac{2.5\ X\ 2}{20+30} = 1.0 \quad \therefore \quad pH = 14 - 1.0 = 13$$

Note that we multiplied the concentration of $Ca(OH)_2$ by 2 because one mole of the later gives 2 moles OH^- .

Tutorial

Exercise 6 : Calculate the pH of the solution resulting from adding 400 mg of NaOH (mw = 40) to 50 mL of 0.1 M NH_4Cl and the volume is completed to 500 mL with distilled water ?

Your answer :

Our answer next slide

Tutorial

answer 6 :

NH_4Cl +	NaOH $\rightarrow$	NH_3 +	NaCl +	H_2O		
50X0.1=5	400/40=10	0	0	0	(mmole)	(I)
0	5	5	5	5	(mmole)	(C)

After the reaction is complete , the final solution contains H_2O , NaCl (neutral salt) , NH_3 and excess NaOH . To simplify the calculations we will ignore NH_3 (weak base) and calculate the pH of only a strong NaOH solution :

$$pOH = -\log \frac{5}{500} = 2 \quad \therefore \; pH = 14 - 2 = 12$$

Tutorial

Exercise 7 : Calculate the pH of a solution resulting from adding 25 mL of 0.1 M of NaOH to 25 mL of 0.15 M HCl solution ?

Your answer :

Our answer next slide

Tutorial

Answer 7 :

NaOH	+ HCl	$\rightarrow$	NaCl	+	H_2O		
25X0.1=2.5	25X0.15 =3.75		0		0	(mmoles)	(I)
0	1.25		2.5		2.5	(mmoles)	(C)

The resulting solution after the reaction is acidic because NaOH is the limiting reactant and HCl is the excess reactant (1.25 mmoles of HCl is remaining) therefore the pH can be calculated as such :

$$pH = -\log \frac{1.25}{25 + 25} = 1.6$$

Tutorial

Exercise 8 : Calculate the pH of a solution resulting from adding 25 mL of 0.1 M of NaOH to 25 mL of 0.05 M HCl solution ?

Your answer :

Our answer next slide

Tutorial

Answer 8 :

NaOH	+	HCl	$\rightarrow$	NaCl	+	H_2O		
25X0.1=2.5		25X0.05 =1.25		0		0	(mmoles)	(I)
1.25		0		1.25		1.25	(mmoles)	(C)

The solution after complete reaction is basic because HCl is the limiting reactant and the remaining mmoles of NaOH is 1.25 mmoles as shown in the above reaction .

$$pOH = -\log\frac{1.25}{25+25} = 1.6 \quad \therefore \quad pH = 12.4$$

Tutorial

Exercise 9 : Calculate the pH of a solution resulting from adding 25 mL of 0.1 M of NaOH to 25 mL of H_2O ?

Your answer :

Our answer next slide

Tutorial

Answer 9 :

There is no reaction but the solution of NaOH will be diluted :

$$25 \times 0.1 = 50 \times M \quad \therefore \quad M = 0.05$$

$$pOH = -\log 0.05 = 1.3 \quad \therefore \quad pH = 12.7$$

Tutorial

Exercise 10 : Calculate the pH of a solutions resulting from adding 50 ml 0.08 M HCl to 40 ml 0.1 NH_3 ?

Your answer :

Our answer next slide

Tutorial

Answer 10 :

HCl	+	NH_3	$\rightarrow$	NH_4Cl	
0		0		4	(mmoles) (C)

The result is a solution of a weak base salt (NH_4Cl) so we will apply the equation for the calculation of pH of this kind of salt :

$$pH = -\log \sqrt{\frac{1 X 10^{-14} X \dfrac{4}{90}}{1.75 X 10^{-5}}} \approx 5.3$$

Tutorial

Exercise 11 : Calculate the pH of a solution resulting from adding 200 ml 0.1 M HCl to 50 ml 0.1 M NH_3 ?

Your answer :

Our answer next slide

Tutorial

Answer 11 :

HCl	+	NH_3	$\rightarrow$	NH_4Cl	
15		0		5	(mmoles) (C)

The final solution after complete reaction composes of weak base salt NH_4Cl (acidic) and the remaining of HCl while NH_3 is finished . To facilitate the calculations we will ignore the effect of NH_4Cl and calculate the pH of the remaining HCl

$$pH = -\log \frac{15(mmoles)}{250\,(ml)} = 1.2$$

Tutorial

Exercise 12 : Calculate the pH of a solution resulting from adding 100 ml 0.025 M NaOH to 100 ml 0.025 HA ? K_a for HA = $1.75X10^{-5}$

Your answer :

Our answer next slide

Tutorial

Answer 12 :

HA	+	NaOH	$\rightarrow$	NaA	+	NaCl	+	H_2O	
0		0		2.5		2.5		2.5	(mmoles) (C)

All of acetic acid HA and NaOH have been converted to products . The final solution consists from H_2O , NaCl (neutral) and weak acid salt (NaA) . We will apply the equation that used for the calculation of such salt :

$$pOH = -\log \sqrt{\frac{1X10^{-14}\ X\ \frac{2.5}{200}}{1.75X10^{-5}}} = 5.6 \quad \therefore \quad pH = 8.4$$

Tutorial

Exercise 13 : Calculate the pH of a solution resulting from adding 100 ml 0.2 M NaOH to 100 ml 0.1 HA ?

Your answer :

Our answer next slide

Tutorial

Answer 13 :

$$\begin{array}{ccccccccc} HA & + & NaOH & \rightarrow & NaA & + & NaCl & + & H_2O \\ 0 & & 10 & & 10 & & 10 & & 10 \end{array} \quad \text{(mmoles) (C)}$$

After the reaction is completed , the solution is consists of H_2O , neutral salt (NaCl) , weak acid salt (NaA) and the remaining of NaOH . We will ignor the effect of the weak acid salt and calculate the pH from the remaining NaOH thus :

$$pOH = -\log \frac{10\,(mmoles)}{200\,(ml)} = 1.3 \quad \therefore\ pH = 12.7$$

Tutorial

Exercise 14 : Calculate the pH of a solution resulting from adding 10 mL of $Ba(OH)_2$ solution (pH = 10) to 15 mL of HCl solution (pH = 5) ?

Your answer :

Our answer next slide

Tutorial

Answer 14 :

pH = 10 ∴ pOH = 4 ∴ $[OH^-] = 1 \times 10^{-4}$ ∴ $[Ba(OH)_2] = 1 \times 10^{-4}/2$

mmoles $Ba(OH)_2$ = $10 \times 10^{-4}/2 = 5 \times 10^{-4}$

pH = 5 ∴ $[H^+] = 1 \times 10^{-5}$

mmoles HCl = $15 \times 1 \times 10^{-5} = 1.5 \times 10^{-4}$

$2\,HCl$	+	$Ba(OH)_2$	$\rightarrow$	$BaCl_2$	+	$2\,H_2O$	
1.5×10^{-4}		5×10^{-4}		0		0	(mmoles)(I)
0		4.25×10^{-4}		0.75×10^{-4}		1.5×10^{-4}	(mmoles) (C)

$$pOH = -\log \frac{2 \times 4.25 \times 10^{-4}}{25\,(ml)} \approx 4.5 \quad \therefore \quad pH = 9.5$$

Tutorial

Exercise 15 : A solution of a weak acid HA has a concentration of 0.5 M and pH = 4 , calculate the dissociation constant of this acid ?

Your answer :

Our answer next slide

Tutorial

Answer 15 :

$$pH = -\log\sqrt{K_a X C_a}$$

$$4 \quad = -\log\sqrt{K_a X 0.5} \quad \therefore \; K_a = 2 X 10^{-8}$$

VIDEO

K_a of an acid

VIDEO Library

Tutorial

Look at the following videos which will clear some of the subjects that have been previously discussed .

VIDEO
Salt hydrolysis ?

VIDEO
Library

VIDEO
Salts

على الراغبين الاستماع الى محاضرات الاستاذ الدكتور/ ابراهيم زامل الزامل باللغة العربية عن هذا الموضوع الرجوع الى الروابط التالية :

الاتزان في تفاعلات الحموض و القواعد (حسابات الـرقم الهيدروجيني)

الاتزان في تفاعلات الحموض و القواعد (حسابات الـرقم الهيدروجيني) ٢

الاتزان في تفاعلات الحموض و القواعد (حسابات الـرقم الهيدروجيني) ٣