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CHAPTER 16 ACIDS AND BASES

What you will learn:-

- Typical behavior of acid and bases; neutralization reactions
- Strong and weak acids and bases
- Ionization of water and the pH scale
- Indicators and using acid base titrations to determine an unknown concentration

GRADE 11 REVISION

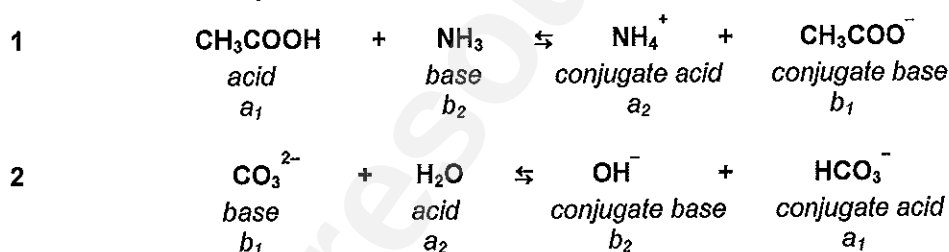
	Properties	Common Examples
ACIDS	Corrosive; sour taste; change the colour of blue litmus to red; neutralize bases to form a salt + water; react with carbonates to form a salt, H ₂ O and CO ₂ ; provide H ⁺ ions in solution.	HCl hydrochloric acid HNO ₃ nitric acid CH ₃ COOH ethanoic acid H ₂ SO ₄ sulfuric acid (COOH) ₂ oxalic acid H ₃ PO ₄ phosphoric acid
BASES	Solutions feel soapy; change the colour of red litmus to blue; neutralize acids to form a salt + water; provide OH ⁻ ions in solution.	NaOH sodium hydroxide KOH potassium hydroxide LiOH lithium hydroxide NH ₄ OH ammonium hydroxide Ca(OH) ₂ calcium hydroxide Mg(OH) ₂ magnesium hydroxide

BRØNSTED-LOWRY DEFINITION

Acids are proton donors. Bases are proton acceptors.

The proton in this definition is an H⁺ ion. An H⁺ ion is a hydrogen atom with no electron therefore it is simply a proton. The H⁺ ion combines with a water molecule to form an H₃O⁺ ion.

Consider the following examples:-



In the above examples CH₃COOH and H₂O **donate protons** and behave as **acids**. NH₃ and CO₃²⁻ **accept protons** and behave as **bases**.

These are **acid-base** or **protolytic** reactions because there is a **transfer of protons**.

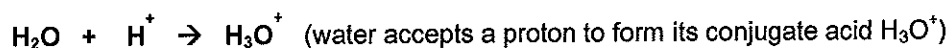
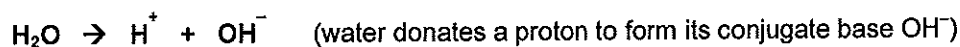
In an acid-base reaction there are always two conjugate acid-base pairs.

The **conjugate base** of an acid is the particle formed when the acid has donated a proton.

The **conjugate acid** of a base is the particle formed when the base has accepted a proton.

Acids can be **monoprotic**, **diprotic** or **triprotic** – they can donate one, two or three H⁺ ions. (E.g. HCl is monoprotic and H₃PO₄ is triprotic)

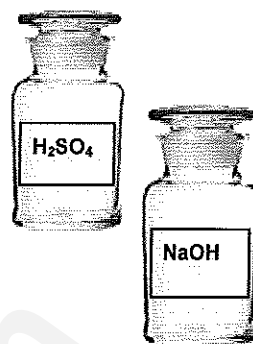
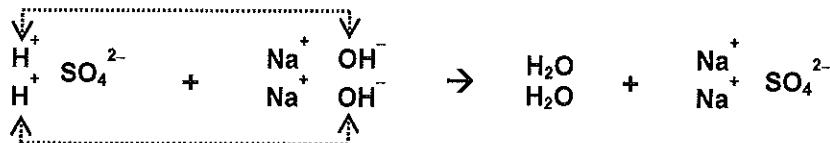
Substances, such as H₂O, which can donate or accept protons (can act as acids or bases) are said to be **amphiprotic** and are called **ampholytes**.



Other examples of ampholytes are HSO₄⁻ and HCO₃⁻.

NEUTRALIZATION

When an acid reacts with a base the H^+ ions provided by the acid combine with the OH^- ions provided by the base to form water molecules. The positive ion from the base and the negative ion from the acid remain in solution as a salt.



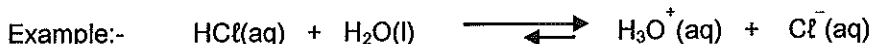
EXERCISE 16.1

- 1 Give one word or term for each of the following:-
 - 1.1 A solution which turns red litmus blue.
 - 1.2 A reaction in which H^+ ions are transferred.
 - 1.3 A reaction which forms a salt and water only.
 - 1.4 The H_3O^+ ion.
 - 1.5 The particle formed when a base accepts a proton.
 - 1.6 A substance which can give and take protons.
 - 1.7 An acid which can donate three protons.
- 2 Use Lewis structures and show how a water molecule can form a hydronium ion.
- 3 Some formulae are given for common laboratory substances. Name each substance and say whether it is an acid, base or salt.
 - (i) K_2SO_4 (ii) H_2SO_4 (iii) NH_4OH (iv) CaCO_3 (v) HCl(aq) (vi) NaOH (vii) HNO_3 (viii) MgCl_2
- 4 Say which of the above substances you would use to make the following and in each case write a balanced equation for the reaction.
 - (i) ammonium nitrate (ii) sodium sulfate (iii) calcium chloride.
- 5 Give the conjugate bases for the following:- H_2SO_4 ; NH_4^+ ; H_2O ; HCO_3^- ; H_3PO_4
- 6 Give the conjugate acids for the following:- NO_3^- ; HS^- ; SO_4^{2-} ; HPO_4^{2-} ; H_2O
- 7.1 Why are some acids said to be monoprotic while others are diprotic?
- 7.2 From the following list pick the acids which are monoprotic.
 - (i) H_2CO_3 (ii) CH_3COOH (iii) $(\text{COOH})_2$ (iv) HNO_3 (v) H_3PO_4 .
- 8.1 What is an ampholyte?
- 8.2 Write equations to show why the HSO_4^- ion is regarded as an ampholyte.
- 9 Label the conjugate acid/base pairs in the following reactions.
 - 9.1 $\text{H}_2\text{O} + \text{HCl} \rightarrow \text{Cl}^- + \text{H}_3\text{O}^+$
 - 9.2 $\text{H}_2\text{CO}_3 + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{HCO}_3^-$
 - 9.3 $\text{S}^{2-} + \text{H}_3\text{O}^+ \rightarrow \text{HS}^- + \text{H}_2\text{O}$
 - 9.4 $\text{H}_2\text{O} + \text{HCOO}^- \rightarrow \text{HCOOH} + \text{OH}^-$
 - 9.5 $\text{BrO}^- + \text{H}_2\text{O} \rightarrow \text{HBrO} + \text{OH}^-$
 - 9.6 $\text{H}_2\text{SO}_4 + \text{NO}_3^- \rightarrow \text{HNO}_3 + \text{HSO}_4^-$
 - 9.7 $\text{H}_2\text{O} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{OH}^-$
 - 9.8 $\text{NH}_4^+ + \text{Cl}^- \rightarrow \text{NH}_3 + \text{HCl}$
 - 9.9 $\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{HPO}_4^{2-}$

STRENGTH OF ACIDS AND BASES

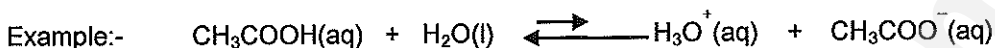
The strength of an acid or base is an indication of the extent to which it is ionized in solution.

- Strong acids and bases are fully ionized or dissociated in solution.



The solution contains a large number of H_3O^+ and Cl^- ions. There are very few un-ionized HCl molecules. Hydrochloric acid is a **strong acid**. Because of the large number of ions the solution is a good conductor of electricity and is a **strong electrolyte**.

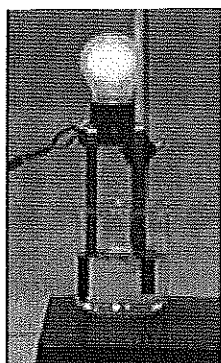
- Weak acids and bases are only partially ionized or dissociated in solution.



There are very few H_3O^+ and CH_3COO^- ions in solution and a large number of un-ionized CH_3COOH molecules. Ethanoic acid is a **weak acid** and the solution is a **weak electrolyte**.

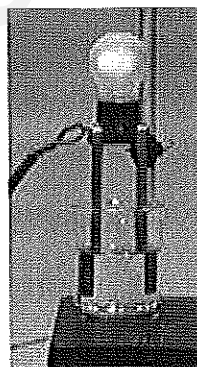
(i)
The bright light shows that the solution of HCl is a good electrolyte.

The concentration of ions is high.



(ii)
The dim light shows that the solution of CH_3COOH is a poor electrolyte.

The concentration of ions is low.



STRONG ACIDS: H_2SO_4 ; HCl ; HNO_3

STRONG BASES: group I oxides and hydroxides

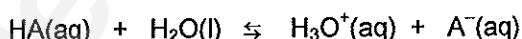
WEAK ACIDS: H_2CO_3 ; CH_3COOH

WEAK BASES: NH_3 ; HCO_3^- ion

- Concentrated** and **dilute** refer to the number of moles of acid or base dissolved in 1 dm^3 of water. **Strong** and **weak** refer to the extent of ionization. It is possible to have a **CONCENTRATED, WEAK acid**.

IONIZATION CONSTANTS FOR ACIDS AND BASES

When acids dissolve in water they ionize according to the following general equation



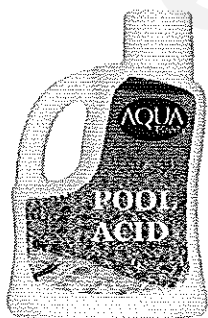
The equilibrium constant for this process is written as:

This is called the ionization constant for the acid.

The value of K_a indicates whether the acid is strong or weak i.e how readily the acid will donate protons.

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

A **strong** acid ionizes to a large extent and has a large K_a value.
eg pool acid HCl .



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{Cl}^-]}{[\text{HCl}]} > 10^2 \text{ at } 25^\circ\text{C}$$

A **weak** acid has a low K_a value as the result of its low ionization
eg vinegar CH_3COOH .



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1,8 \times 10^{-5} \text{ at } 25^\circ\text{C}$$

We can also write an expression for the equilibrium constant (K_b) for a base which ionizes as follows :

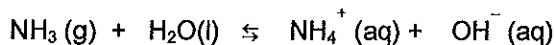


The equilibrium constant for this reaction will be:-

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$



Ammonia is a typical weak base. It ionizes as follows.



$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1,8 \times 10^{-5} \text{ at } 25^\circ \text{C}$$

The value for the equilibrium constant shows that there are not many NH_4^+ and OH^- in ammonia solution. This means that ammonia solution can be used as a mild cleaning agent in the home.



By contrast caustic soda (NaOH) dissociates completely in solution to give a large number of OH^- ions. For this reason caustic soda is used in more powerful cleaning agents such as oven cleaner or drain cleaner.

Some typical K_a values are shown in the table given below.
You do NOT need to remember these values.

A strong acid has a weak conjugate base and a weak acid has a strong conjugate base.

Acid	Ionization in water	Base	K_a
Hydronium ion	$H_3O^+ \rightleftharpoons H^+ + H_2O$	Water	$>10^2$
Sulfuric acid	$H_2SO_4 \rightleftharpoons H^+ + HSO_4^-$	Hydrogen sulfate ion	$>10^2$
Hydrochloric acid	$HCl \rightleftharpoons H^+ + Cl^-$	Chloride ion	$>10^2$
Nitric acid	$HNO_3 \rightleftharpoons H^+ + NO_3^-$	Nitrate ion	$>10^2$
Sulfurous acid	$H_2SO_3 \rightleftharpoons H^+ + HSO_3^-$	Hydrogen sulfite ion	$1,7 \times 10^{-2}$
Hydrogen sulfate ion	$HSO_4^- \rightleftharpoons H^+ + SO_4^{2-}$	Sulfate ion	$1,2 \times 10^{-2}$
Phosphoric acid	$H_3PO_4 \rightleftharpoons H^+ + H_2PO_4^-$	Dihydrogen phosphate ion	$8,5 \times 10^{-3}$
Methanoic acid	$HCOOH \rightleftharpoons H^+ + HCOO^-$	Methanoate ion	$2,0 \times 10^{-4}$
Ethanoic acid	$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$	Ethanoate ion	$1,8 \times 10^{-5}$
Carbonic acid	$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$	Hydrogen carbonate ion	$4,2 \times 10^{-6}$
Ammonium ion	$NH_4^+ \rightleftharpoons H^+ + NH_3$	ammonia	$5,6 \times 10^{-10}$
Hydrogen carbonate ion	$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$	Carbonate ion	$4,8 \times 10^{-11}$
Water	$H_2O \rightleftharpoons H^+ + OH^-$	Hydroxide ion	$1,8 \times 10^{-16}$

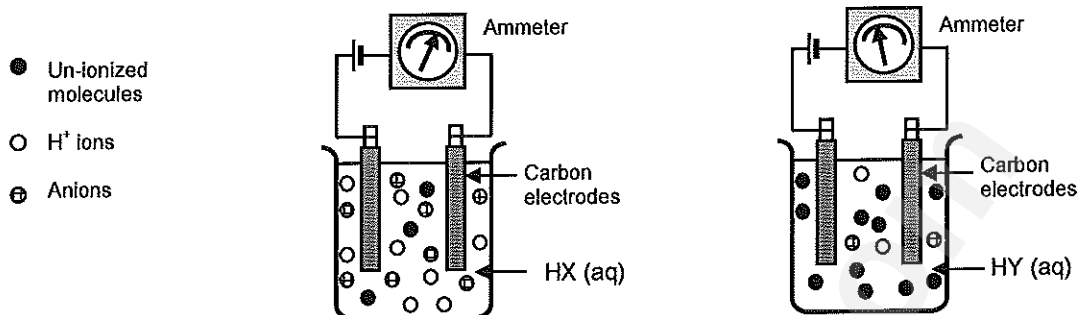
Very weak to very strong

Very weak to very strong

THINK TANK 16.1

1 What is the difference between a weak acid and a dilute acid?

2 The diagrams show two conductivity experiments using solutions of two acids HX and HY.



2.1 What information about the two acids is given by the ammeter reading?

2.2 Compare the concentration of ions in HX(aq) with that in HY(aq).

2.3 The K_a values for these acids are $4,2 \times 10^{-8}$ and $1,7 \times 10^{-2}$. Which value belongs with which acid?

3 Write two equations to show the ionization of sulfuric acid, H_2SO_4 , and mark the conjugate acids and bases.

4 A student measures the H_3O^+ ion concentration in samples of HCl ; H_2SO_4 and H_3PO_4 . Each sample has a concentration of $0,2 \text{ mol.dm}^{-3}$. Explain why:

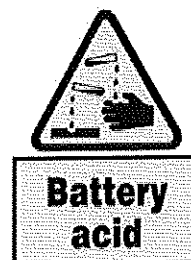
4.1 The concentration of H_3O^+ ions in HCl is $0,2 \text{ mol.dm}^{-3}$ and in H_2SO_4 it is $0,4 \text{ mol.dm}^{-3}$

4.2 The concentration of H_3O^+ ions in H_3PO_4 is $0,042 \text{ mol.dm}^{-3}$ which is much smaller than in the other acids.

5 Sulfuric acid is the acid used in car batteries and it carries the warning sign shown in the picture. Carbonic acid, H_2CO_3 , is formed in fizzy drinks when carbon dioxide dissolves in water.

Look back at the K_a values for these two acids and explain in terms of concentration and strength:-

5.1 Why sulfuric acid carries a warning sign but we can drink carbonic acid.

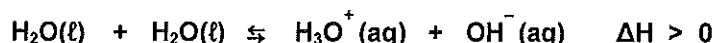


5.2 Give the conjugate base of (i) sulfuric acid (ii) carbonic acid

5.3 Which is the stronger base, the hydrogen sulfate ion or the hydrogen carbonate ion?

IONIZATION OF WATER

Pure water ionizes to a very small extent as follows:-



Water behaves as an ampholyte : one water molecule behaves as a very weak acid and the other water molecule behaves as a very weak base.

In pure water at 25° C the concentration of the hydronium ion is very small.

$$[\text{H}_3\text{O}^+] = 0,0000001 \text{ mol.dm}^{-3} = 10^{-7} \text{ mol.dm}^{-3}$$

For this reason pure water is not an electrolyte.

Tap water and sea water conduct only because of dissolved ionic compounds such as salt.

No current flows through the lamp because water is a non-conductor.



IONIZATION CONSTANT FOR WATER, K_w

We can write an expression for the equilibrium constant for water from the equation given above.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14} \text{ at } 25^\circ \text{C}$$

K_w is called the ionization constant for water. It measures the extent to which water is ionized. Like all equilibrium constants it is temperature dependent so the value 1×10^{-14} is only true at 25° C. Because the ionization of water is endothermic, K_w increases with increasing temperature.

The ionization of water produces equal moles of H_3O^+ and OH^- so from the value for K_w we can determine the concentration of these ions.

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 10^{-7} \text{ mol.dm}^{-3} \text{ at } 25^\circ \text{C}$$

As the acidity or alkalinity of a solution changes the $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ will change from $10^{-7} \text{ mol.dm}^{-3}$ but the product of these two values does not change because a change in concentration does NOT alter the equilibrium constant.

If an acid is added to pure water the concentration of the hydronium ion $[\text{H}_3\text{O}^+]$ will be greater than $10^{-7} \text{ mol.dm}^{-3}$. If hydroxide ions are added to pure water, the concentration of OH^- ions will increase and therefore the concentration of H_3O^+ ions will decrease.

THE pH SCALE

This scale uses numbers from 0 to 14 to indicate whether a solution is acidic or basic. The numbers indicate the concentration of the hydronium ion $[\text{H}_3\text{O}^+]$. As values are often small the pH scale gives a simpler way of writing these numbers.

pH stands for the negative log of the hydronium ion concentration.

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

In pure water or **NEUTRAL** solution $[\text{H}_3\text{O}^+] = 10^{-7} \text{ mol.dm}^{-3}$ $\text{pH} = -(-7) = 7$

In an **ACIDIC** solution the $[\text{H}_3\text{O}^+]$ is high $[\text{H}_3\text{O}^+] > 10^{-7} \text{ mol.dm}^{-3}$ $\text{pH} < 7$
Eg If $[\text{H}_3\text{O}^+] = 10^{-4} \text{ mol.dm}^{-3}$ then $\text{pH} = -(-4) = 4$

In a **BASIC** solution the $[\text{H}_3\text{O}^+]$ is low. $[\text{H}_3\text{O}^+] < 10^{-7} \text{ mol.dm}^{-3}$ $\text{pH} > 7$
Eg If $[\text{H}_3\text{O}^+] = 10^{-9} \text{ mol.dm}^{-3}$ then $\text{pH} = -(-9) = 9$

As the concentration of the H_3O^+ **INCREASES** the pH value **DECREASES**

pH		ACIDIC				NEUTRAL				BASIC				
0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
10^0	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}
$[\text{H}_3\text{O}^+]$		high								low				

WORKED EXAMPLES

CALCULATIONS ON PH

pH stands for the negative log of the hydronium ion concentration.

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

- Calculate the pH of a solution where $[\text{H}_3\text{O}^+] = 2,5 \times 10^{-4} \text{ mol.dm}^{-3}$
 $\text{pH} = -\log 2,5 \times 10^{-4} = 3,6$ On your calculator push $+/- \log 2,5 \exp +/- 4$
- Calculate the $[\text{H}_3\text{O}^+]$ of a solution of $\text{pH} = 5,8$.
 $5,8 = -\log [\text{H}_3\text{O}^+]$ On your calculator push $2^{\text{nd}} \text{ func } \log -5,8$
 $[\text{H}_3\text{O}^+] = 1,58 \times 10^{-6} \text{ mol.dm}^{-3}$
- Calculate the pH of a solution of H_2SO_4 which has a concentration of $0,004 \text{ mol.dm}^{-3}$.
 H_2SO_4 is diprotic. Therefore $[\text{H}_3\text{O}^+] = 2 \times 0,004 = 0,008 \text{ mol.dm}^{-3}$
 $\text{pH} = -\log 0,008 = 2,1$
- Calculate the pH of a solution with $[\text{OH}^-] = 8 \times 10^{-4} \text{ mol.dm}^{-3}$.
 We must first calculate the $[\text{H}_3\text{O}^+]$ using $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$
 $[\text{H}_3\text{O}^+] \times 8 \times 10^{-4} = 10^{-14}$ Therefore $[\text{H}_3\text{O}^+] = 1,25 \times 10^{-11}$
 $\text{pH} = -\log 1,25 \times 10^{-11} = 10,9$
- Calculate the $[\text{OH}^-]$ of a solution of $\text{pH} = 13,4$.
 $13,4 = -\log [\text{H}_3\text{O}^+]$ Therefore $[\text{H}_3\text{O}^+] = 3,98 \times 10^{-14} \text{ mol.dm}^{-3}$
 $[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$ $3,98 \times 10^{-14} [\text{OH}^-] = 10^{-14}$
 $[\text{OH}^-] = 0,25 \text{ mol.dm}^{-3}$

THINK TANK 16.2

Universal indicator can be used to show the approximate pH of a solution. The following diagram shows the colours of universal indicator at various pH values. The pH values for some common substances are marked. Study the diagram and then answer the questions that follow.

Gastric juice			oranges		bread		blood		toothpaste		milk of magnesia		oven cleaner		
Universal Indicator red			orange		yellow		green		blue		dark blue		purple		
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
10^0				10^{-3}				10^{-7}		10^{-9}			10^{-12}		10^{-14}
H_3O^+															
10^{-14}				10^{-11}		10^{-9}		10^{-7}					10^{-2}		OH^- 10^0

- 1.1 Given below is a list of common household substances and their pH values. Fill in these substances on the diagram above. In the space given below write in the colour you would expect to see if you tested these substances with universal indicator?

eggs 7,8 soap 12,3 tomatoes 4,2
 bleach 11 potatoes 5,8 fizzy drinks 3,0
 milk 6,5 seawater 8,5 lemons 2,3

- 1.2 Complete the diagram given above to show how $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ change with changing pH value.
- 1.3 Gastric juice contains dilute hydrochloric acid. What is the approximate concentration of H_3O^+ in gastric juice?

- 1.4 When Uncle Ben has acid indigestion he takes a little milk of magnesia. Why does this help?

- 1.5 Milk of magnesia is a suspension of $\text{Mg}(\text{OH})_2$. Write a balanced equation to show how milk of magnesia reacts with stomach acid (HCl).

2 Explain why:-

- 2.1 $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ in pure water.

- 2.2 $[\text{H}_3\text{O}^+] = [\text{HCl}]$ in a solution of hydrochloric acid.

- 2.3 $[\text{H}_3\text{O}^+] = 2 \times [\text{H}_2\text{SO}_4]$ in a solution of sulfuric acid.

- 3 Complete the following table by calculating either pH or the $[\text{H}_3\text{O}^+]$ of the given solutions. Assume complete ionization.

Acid concentration mols.dm^{-3}	$[\text{H}_3\text{O}^+]$ mols.dm^{-3}	Calculator work	pH
3.1 $[\text{HCl}] = 7 \times 10^{-5}$	7×10^{-5}	$-\log 7 \times 10^{-5}$	
3.2 $[\text{HNO}_3] = 0,32$			
3.3 $[\text{HCl}] = 0,0048$			
3.4 $[\text{H}_2\text{SO}_4] = 0,32 \text{ mol.dm}^{-3}$			
3.5 $[\text{HCl}] =$		2 nd function log -4,7	4,7
3.6 $[\text{HNO}_3] =$			1,8
3.7 $[\text{H}_2\text{SO}_4] =$			0,6

- 4 The concentration of a solution of KOH is $0,04 \text{ mol.dm}^{-3}$. Calculate:-

- 4.1 The concentration of the H_3O^+ ion.

- 4.2 The pH

- 5 Complete the following table by calculating the $[\text{H}_3\text{O}^+]$ of the given solutions and then the pH.

Base concentration mols.dm^{-3}	$[\text{OH}^-]$ mols.dm^{-3}	Calculator work	$[\text{H}_3\text{O}^+]$ mols.dm^{-3}	pH
5.1 $[\text{NaOH}] =$	7×10^{-5}	$[\text{H}_3\text{O}^+].(7 \times 10^{-5}) = 1 \times 10^{-14}$		
5.2 $[\text{Ba}(\text{OH})_2] =$	$2,4 \times 10^{-3}$			
5.3 $[\text{NaOH}] = 0,32$				
5.4 $\text{Ba}(\text{OH})_2 = 0,32$				

HYDROLYSIS OF SALTS

A salt is an ionic compound consisting of positive ions (cations) other than H^+ and negative ions (anions) other than OH^- . Salts are formed when an acid reacts with a base.

When salts are dissolved in water they dissociate completely. The cations may react with water molecules forming H_3O^+ ions and the anions may react with water molecules forming OH^- ions. This reaction of the ions with water is called **HYDROLYSIS**.

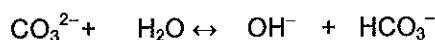
Since hydrolysis produces either extra H_3O^+ ions or extra OH^- ions the pH of the salt solution depends on the type of salt dissolved.

- 1 **Salt of a strong acid + strong base:** E.g. $NaCl$ neutral solution pH = 7

Anions of strong acids (Cl^-) and the cations of strong bases (Na^+) do not undergo hydrolysis. A solution of this type of salt will be neutral.

- 2 **Salt of a weak acid + strong base:** E.g. Na_2CO_3 basic solution pH > 7

Anions of weak acids (CO_3^{2-}) undergo hydrolysis but the cations of strong bases (Na^+) do not.



The CO_3^{2-} ion is the conjugate base of a weak acid HCO_3^- . It is able to remove a proton from a water molecule creating an excess of OH^- ions. The solution is basic.

- 3 **Salt of a strong acid + weak base:** E.g. NH_4Cl acidic solution pH < 7

Cations of weak bases (NH_4^+) undergo hydrolysis but the anions of strong acids (Cl^-) do not.



The NH_4^+ ion is the conjugate acid of a weak base NH_3 . It is able to donate a proton to a water molecule creating an excess of H_3O^+ ions. The solution is acidic.

- 4 **Salt of a weak acid + weak base:** E.g. $(NH_4)_2CO_3$ can be acidic, basic, or neutral

Both cations and anions undergo hydrolysis. The pH of this type of solution depends on the relative strengths of the acids and bases.

THINK TANK 16.3

- What is meant by hydrolysis?
- Which cations undergo hydrolysis?
- Which anions undergo hydrolysis?
- The table shows the pH of some salt solutions. Complete the table to show the strength of the acid and the base from which the salt was formed.

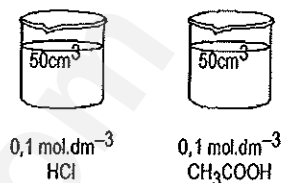
Type of salt	Example	pH of aqueous solution
4.1	KNO_3	neutral
4.2	K_2CO_3	>7 basic
4.3	NH_4Cl	<7 acidic
4.4 Weak acid – weak base	$(NH_4)_2CO_3$	depends on the salt

EXERCISE 16.2

- 1.1 What is the difference between a strong acid and a concentrated acid?
 1.2 Write the following in order of conductivity from highest to lowest:-
 (i) pure water (ii) a solution of sodium hydroxide (iii) a solution of ammonium hydroxide.

- 2.1 Write the expressions for K_a for HCO_3^- and HNO_3 .
 2.2 Which of these particles will donate a proton more easily? Give a reason for your answer.

- 3.1 The diagram shows two beakers which both contain acids.
 3.2 Define an acid according to the Brønsted Lowry theory.
 3.3 Give the conjugate bases of these two acids.
 3.4 Which of these bases will be the stronger?
 3.5 Which one of these solutions will have the higher concentration of hydronium ions. Give a reason for your answer.



- 4.1 What is the value of K_w at 25°C?
 4.2 What does this imply about how well water ionizes?
 4.3 How will the value of K_w change if (i) a strong acid is added to water (ii) the temperature rises to 30°C?
 4.4 Why is the pH of a neutral solution 7?

5 A solution of HCl has a concentration of 0,015 mol.dm⁻³ at 25°C.

- 5.1 What is the concentration of the hydronium ions?
 5.2 Calculate the pH of the solution.

6 A solution of NaOH has a concentration of 0,002 mol.dm⁻³ at 25°C.

- 6.1 What is the concentration of (i) the hydroxide ions (ii) the hydronium ions?
 6.2 Calculate the pH of the solution.

7 Calculate the pH of a solution which has

- 7.1 $[\text{H}_3\text{O}^+] = 0,0065 \text{ mol.dm}^{-3}$
 7.2 $[\text{OH}^-] = 8,6 \times 10^{-2} \text{ mol.dm}^{-3}$

8 Solutions of two acids HX and HY are prepared. Each solution has a concentration of 0,2 mol.dm⁻³. The pH of acid HX solution is 2,7 and the pH of acid HY solution is 0,7.

- 8.1 Calculate the $[\text{H}_3\text{O}^+]$ for each acid sample.
 8.2 Which acid is stronger? Give a reason for your answer

9 A sample of acid has $[\text{H}_3\text{O}^+] = 0,1 \text{ mol}$ in 1 dm³ of water. What happens to the pH if it is diluted to 10 dm³ of water?

10 A learner obtained the following readings for the pH of some solutions .

solution	Hydrochloric acid	Grape juice	Sodium hydroxide.
pH	1	3,1	11

- 10.1 Which solution contains the highest concentration of hydronium ions
 10.2 Determine the $[\text{H}_3\text{O}^+]$ in grape juice.
 10.3 How will the pH of grape juice change if (i) distilled water is added (ii) some of the sodium hydroxide solution is added (iii) some of the hydrochloric acid solution is added?
 10.4 Determine the OH^- concentration in sodium hydroxide.
 11 What are the concentrations of the following solutions?
 11.1 H_2SO_4 solution pH = 0,9
 11.2 $\text{Ba}(\text{OH})_2$ solution pH = 12,6

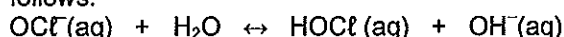
- 12 When 0,05 moles of $\text{HBr}(\ell)$ are added to 0,005 moles of $\text{KOH}(\text{aq})$ in a beaker the final volume of the solution is 75 cm^3 . HBr ionizes completely when dissolved in water.
- 12.1 Write a balanced equation for the reaction between HBr and KOH .
- 12.2 Which reagent is in excess?
- 12.3 Calculate the pH of the final solution.

- 13 You are provided with four dilute solutions of the same concentration.

A NaOH B H_2SO_4 C. NH_4OH D. CH_3COOH

- 13.1 Which one of these (i) Has the lowest pH? (ii) Is a monoprotic acid? (iii) Is a weak base?
- 13.2 Which two solutions will react to form a salt which has a pH of greater than 7 in solution? Explain your answer by referring to the strengths of acids and bases.

- 14 "Pool Chlorine" is calcium hypochlorite $\text{Ca}(\text{OCl})_2$. This dissolves in water to release the hypochlorite ion $\text{OCl}^- (\text{aq})$ which combines with water molecules to form hypochlorous acid as follows.



- 14.1 Give the conjugate acid-base pairs in this reaction.
- 14.2 Give one term for the interaction between water molecules and ions in solution.
- 14.3 The pH in swimming pools has to be maintained at 7,2 to 7,6. How is the pH of a swimming pool affected by adding calcium hypochlorite?
- 14.4 An owner measures the pH of a swimming pool as 7,8.
- 14.4.1 Calculate the $[\text{H}_3\text{O}^+]$ in the pool.
- 14.4.2 Should the owner add pool acid (HCl) or sodium carbonate (Na_2CO_3) to change the pH of the pool? Explain your answer.



- 15 For each of the following (i) give the colour of litmus paper in the solution (ii) Write a balanced equation and clearly show hydrolysis of the relevant ions:-
- 15.1 Ammonium chloride solution
- 15.2 Sodium carbonate solution

- 16 Given the following salt solutions $\text{CH}_3\text{COONa}(\text{aq})$ and $(\text{NH}_4)_2\text{SO}_4(\text{aq})$

- 16.1 Say which ion would undergo hydrolysis in each salt solution.
- 16.2 Will these solutions be acidic, basic or neutral?

- 17 Exactly 0,48g of the strong base $\text{M}(\text{OH})_x$, relative formula mass 120, is present in 160 cm^3 of solution.

- 17.1 What is meant by the concept "strong base"?
- 17.2 If the pH of the solution is 12,7 calculate the number of mols of OH^- which it contains.
- 17.3 Use your answer to 17.2 to calculate the value of x in $\text{M}(\text{OH})_x$.

- 18 A $0,1 \text{ mol.dm}^{-3}$ solution of KOH conducts electricity better than a $0,1 \text{ mol.dm}^{-3}$ solution of NH_3 in water because the KOH solution:-

- A. Contains fewer ions than the NH_3 solution.
- B. Is a stronger base than the NH_3 solution.
- C. Contains more H_3O^+ ions than the NH_3 solution.
- D. Is a weaker base than the NH_3 solution.

- 19 X is a solution of HCl and Y is a solution of HCOOH . The solutions have equal concentrations. How do the concentration of H_3O^+ and the pH of X and Y compare?

- | H_3O^+ | pH |
|----------------------------|-------------------------|
| A. higher for X than for Y | higher for X than for Y |
| B. higher for X than for Y | lower for X than for Y |
| C. lower for X than for Y | higher for X than for Y |
| D. lower for X than for Y | lower for X than for Y |

ACID-BASE TITRATIONS (volumetric analysis)

A **titration** is the addition of a measured volume of one solution, from a burette, to a known volume of another solution until the reaction between them is complete (an **end-point** is reached).
An acid-base titration is carried out to determine an **unknown** concentration by using a solution of **known** concentration.

A solution of known concentration is called a **standard solution**.

An **indicator** is used to show the end-point of a titration by a sharp colour change.

PREPARING A STANDARD SOLUTION

- The substance used for the standard solution must have a high degree of purity; preferably be a solid, because this is easier to measure, and it must not be hygroscopic (absorb water) or efflorescent (give off water).
- We use oxalic acid crystals $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$ and anhydrous sodium carbonate Na_2CO_3 to make standard solutions.
- Weigh the solid accurately and transfer it to a volumetric flask. Add water up to the mark. This gives a known mass in a known volume.
- Calculate the concentration of the solution from

$$c = \frac{m}{MV}$$

c = concentration in $\text{mol} \cdot \text{dm}^{-3}$; m = mass of sample
 M = molar mass; V = volume of solution in dm^3



volumetric flask

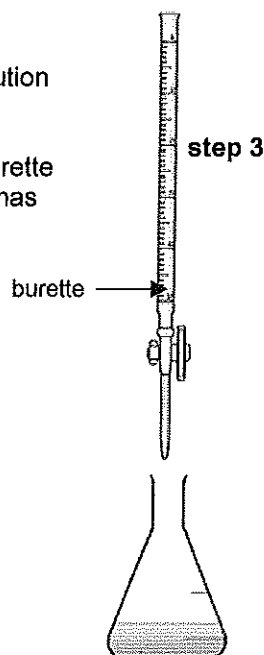
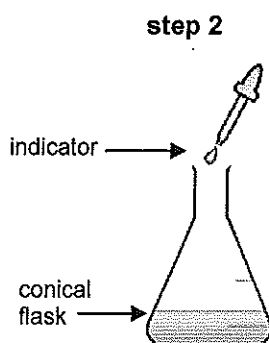
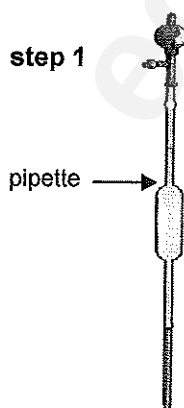
CHOOSING AN INDICATOR

Acid-base titrations are based on neutralization reactions that produce a salt and water. The pH of the salt solution at the end point depends on the strength of the acid and base that react. Indicators are dyes that change colour at a specific pH value so we select the indicator according to the pH of the solution at the end-point.

Acid-base strength	pH of solution	Indicator	Colour change Acid to Base	pH range
Strong acid-weak base	<7 acidic salt	methyl orange	red to yellow	3,1 – 4,4
Strong acid-strong base	= 7 neutral salt	bromothymol blue	yellow to blue	6,0 – 7,6
Weak acid-strong base	>7 basic salt	phenolphthalein	colourless to pink	8,3 – 10,0

DOING A TITRATION

- Use a pipette to measure out a known volume of the standard solution and place this in a conical flask.
- Add a drop of indicator
- Add the solution of unknown concentration drop by drop from a burette until the indicator changes colour. This shows that the end point has been reached.



Write down the balanced equation for the reaction.



When the end-point is reached the acid and base have reacted in the same molar ratio as found in the balanced equation.

We can write:-

$$\frac{\text{number of mols of acid}}{\text{number of mols of base}} = \frac{a}{b}$$

But $n = cV$ therefore mols of acid $= n_A = c_A V_A$ and mols of base $= n_B = c_B V_B$

Substituting in the above equation we have

$$\frac{c_A V_A}{c_B V_B} = \frac{a}{b}$$

If we know both volume values and one concentration value we can find the unknown concentration. (Note : there is no need to change the volume units in this calculation because they cancel.)

E.g. In a titration 6,25 cm³ of dilute H₂SO₄ neutralizes 25 cm³ of NaOH. If the concentration of the H₂SO₄ solution is 0,2 mol.dm⁻³ calculate the concentration of the NaOH.



$$\frac{c_B V_B}{c_A V_A} = \frac{b}{a} \quad \text{Substituting} \quad \frac{c_B \times 25}{0,2 \times 6,25} = \frac{2}{1} \quad c_B = 0,1 \text{ mol.dm}^{-3}$$

THINK TANK 16.4

- 1 What is meant by a standard solution?.....
 - 2 Why is sodium carbonate suitable for use as a standard solution
-
- 3 0,53g of sodium carbonate is dissolved in water to a final volume of 500 cm³.
 - 3.1 Calculate the concentration of the solution of sodium carbonate in mol.dm⁻³.
- 20 cm³ of the sodium carbonate solution is used to neutralize 40 cm³ of H₂SO₄ solution.
- 3.2 Write a balanced equation for the reaction that takes place.
 - 3.3 Calculate the concentration of the acid.
-
- 3.4 What will the pH be at the end-point?.....
 - 3.5 Which indicator should be used in this titration?.....
 - 3.6 Why is universal indicator not used to show the end-point in a titration?

- 4 If 12,6g of oxalic acid crystals $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$ are dissolved in 500 cm^3 of water, calculate the concentration of the oxalic acid solution in mol.dm^{-3} .
- 5 The standard solution of oxalic acid made in question 4 is used in a titration to standardize a solution of sodium hydroxide.
- 5.1 What does the term **titration** mean?.....
- 5.2 What does the term **standardize** mean?.....
- 5.3 Which indicator should be used in this titration?.....
- 5.4 Explain why this indicator is chosen.
- 5.5 The diagrams given below show steps in the titration. The burette is filled with acid.
- 5.5.1 Write labels to name the apparatus used and show which holds the acid and which holds the base.
- 5.5.2 Take the reading shown in diagram (ii).....
- 5.5.3 Give the colour of the contents of the flask in diagram (iii) when the indicator is added.

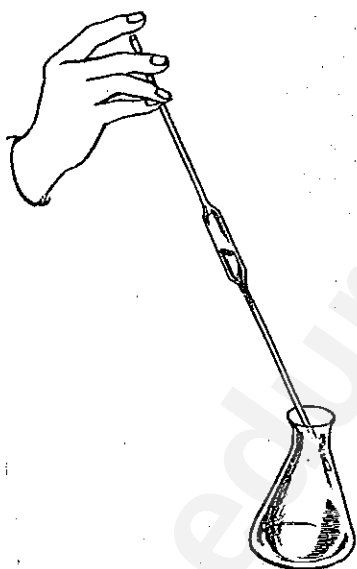


diagram (i)

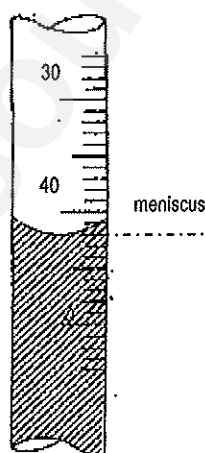


diagram (ii)

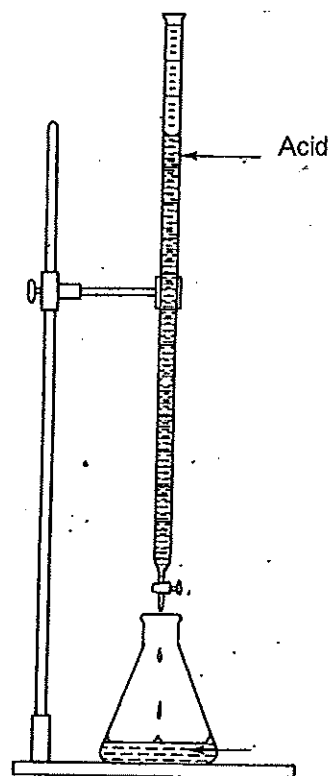


diagram (iii)

5.5.4 Write down a brief description of the process shown in diagram (i) and (iii).

.....

.....

.....

.....

.....

5.6 How does the pH of the mixture change while the acid is being added from the burette to the base in the flask?

.....

5.7 Give the colour change of the indicator at the end-point of the titration?

.....

5.8 During the titration some sodium hydroxide is spilled against the side of the flask and it is washed down with distilled water. Why will this not affect the results of the titration?

.....

5.9 Balance the equation for the chemical reaction that takes place in the flask.



5.10 The table shows the results of the titration. Complete the table and decide which result should be discarded. Concentration of oxalic acid = $0,2 \text{ mol.dm}^{-3}$

Burette Reading		Volume of Acid cm^3	Volume of Base cm^3
Reading 1	Reading 2		
5,5	21,2		25
6,3	21,4		25
4,5	19,4		25
6,8	21,9		25
Average volume			25

5.11 Use the remaining results to calculate the concentration of the sodium hydroxide solution.

- 6 You now use the standardized sodium hydroxide solution to determine the concentration of a sample of dilute H_2SO_4 . Before you start you dilute 25 cm^3 of the acid by adding sufficient water to make a volume of 250 cm^3 .
- 6.1 Give a balanced equation for the reaction.
- 6.2 Give the indicator you would use and the colour change you would see.
- 6.3 Suppose that $7,5 \text{ cm}^3$ of acid react with 25 cm^3 of sodium hydroxide solution. Calculate the concentration of the diluted acid.
- 6.4 Use the fact that $c_1V_1 = c_2V_2$ to calculate the concentration of the original acid sample.

ACTIVITY 16.1 to prepare a standard solution of oxalic acid
of concentration $0,05 \text{ mol.dm}^{-3}$

You will need:- oxalic acid crystals; massmeter; small beaker; distilled water; funnel; volumetric flask.

The formula of oxalic acid crystals is $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$. It has a molar mass of 126 g.mol^{-1} . If you dissolve about 7g in 1 dm^3 of water the concentration will be about $0,05 \text{ mol.dm}^{-3}$. If you need to make a smaller volume, adjust the mass accordingly.

- 1 Measure the mass of the beaker accurately
- 2 Add about 7g of oxalic acid crystals and measure the mass of the beaker and contents.
- 3 Rinse the 1 dm^3 volumetric flask with distilled water and place a funnel in the neck of the flask.
- 4 Add a little distilled water to the oxalic acid in the beaker and stir to dissolve.
- 5 Transfer the solution to the volumetric flask and rinse the beaker with distilled water several times. Add the rinsings to the flask and also rinse the funnel.
- 6 Half fill the flask with distilled water and shake it to dissolve the solid.
- 7 Fill the flask up to the mark with distilled water.
- 8 Calculate the concentration of the solution as shown in the following example.

$$\text{Mass of beaker} = 16,65 \text{ g}$$

$$\text{Mass of beaker} + \text{oxalic acid} = 23,46 \text{ g}$$

$$\text{Mass of oxalic acid} = 6,81 \text{ g}$$

$$\text{Volume} = 1 \text{ dm}^3$$

$$n = m/M \quad n = 6,81/126 = 0,054 \text{ mol}$$

$$c = n/V \quad c = 0,054/1 = 0,054 \text{ mol.dm}^{-3}$$

ACTIVITY 16.2 to standardize a solution of sodium hydroxide using a standard solution of oxalic acid.

You will need:- standard solution of oxalic acid; pipette; burette; burette stand; conical flask; phenolphthalein indicator; white tile or piece of white paper.

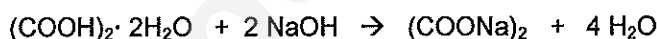
- 1 Rinse a conical flask with distilled water and use a pipette to add a measured volume of the standard solution of oxalic acid to the flask (usually 20 or 25 cm³).
- 2 Note the concentration and volume of the acid used.
- 3 Add about 5 drops of phenolphthalein indicator to the conical flask. The solution should remain colourless.
- 4 Rinse a burette with a little distilled water and then with a small amount of sodium hydroxide solution.
- 5 Fill the burette with NaOH and allow a little to run through the tap into a beaker so that the tip of the burette is also filled.
- 6 Clamp the burette in a stand and take the reading.
- 7 Place the conical flask on a white tile beneath the outlet of the burette.
- 8 Run the NaOH solution into the conical flask a little at a time while swirling the contents of the flask. When the indicator starts to turn pink add the base drop by drop.
- 9 As soon as the indicator changes colour permanently stop titrating and take the reading on the burette.
- 10 Repeat the titration at least 3 times.
- 11 Calculate the concentration of the sodium hydroxide as shown below.

Concentration of acid = 0,054 mol.dm⁻³

Volume of acid = 25 cm³

Trial	1	2	3	4
Volume of base used (cm ³)	22,8	22,2	22,4	22,6

Average volume of base = 22,5 cm³

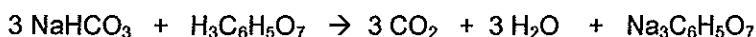


$$\frac{c_B V_B}{c_A V_A} = \frac{b}{a} \quad \text{Substituting} \quad \frac{c_B \times 22,5}{0,054 \times 25} = \frac{2}{1} \quad c_B = 0,12 \text{ mol.dm}^{-3}$$

EXERCISE 16.3

- 1 For each of the following write a balanced equation and then do the calculation.
 - 1.1 25 cm³ of HCl reacts completely with 6,25 cm³ of NaOH. If the concentration of the acid is 0,2 mol.dm⁻³ calculate the concentration of the base.
 - 1.2 25 cm³ of a solution of Ca(OH)₂ of concentration 0,3 mol.dm⁻³ reacts completely with 40 cm³ of a HCl solution. Calculate the concentration of the HCl solution.
 - 1.3 22,5 cm³ of sulfuric acid solution with a concentration of 0,15 mol.dm⁻³ are used to neutralize 30cm³ of potassium hydroxide. Calculate the concentration of the potassium hydroxide solution.
- 2 25 cm³ of Ba(OH)₂ solution is standardized by a solution of HNO₃ of concentration 0,1 mol.dm⁻³.
 - 2.1 The indicator added to the base in the conical flask colours the solution blue. Which indicator was used?
 - 2.2 Why is this indicator used?
 - 2.3 If 20 cm³ of the HNO₃ solution is required to reach the end-point, calculate the concentration of the Ba(OH)₂ solution. $\text{Ba(OH)}_2 + 2\text{HNO}_3 \rightarrow \text{Ba(NO}_3)_2(\text{aq}) + 2\text{H}_2\text{O}$

- 3 What volume of an ammonium hydroxide solution of concentration $0,2 \text{ mol.dm}^{-3}$ is required to neutralize 50 cm^3 of $0,09 \text{ mol.dm}^{-3}$ sulfuric acid. The reaction is as follows:-
 $2\text{NH}_4\text{OH} + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4 + 2\text{H}_2\text{O}$
- 4 Three reagent bottles are labeled:
 X : $0,1 \text{ mol.dm}^{-3} \text{ NaOH}$ Y : $0,05 \text{ mol.dm}^{-3} \text{ H}_2\text{SO}_4$ Z : $0,05 \text{ mol.dm}^{-3} \text{ CH}_3\text{COOH}$
- 4.1 Ethanoic acid is a weak acid. Explain what is meant by "weak acid".
 4.2 Which of these reagents would have the lowest pH?
 4.3 Name the indicator you would use if 'X' is titrated with "Y".
 4.4 What volume of "Z" is required to neutralize 20 cm^3 of "X" according to the following equation?
 $\text{CH}_3\text{COOH(aq)} + \text{NaOH(aq)} \rightarrow \text{CH}_3\text{COONa(aq)} + \text{H}_2\text{O(l)}$
- 5 Antacids are used to neutralize stomach acid that causes indigestion. Some antacids contain weak acids such as citric acid together with sodium bicarbonate. When these are added to water they "fizz" because carbon dioxide gas is produced, as well as a salt and water. Sodium hydrogen carbonate reacts with citric acid as follows.



- 5.1 Why is citric acid triprotic?
 5.2 Sodium citrate is the salt of a weak acid. Would the pH of the solution be 7, greater than 7 or less than 7? Give a reason for your answer.
 5.3 Calculate the mass of solid NaHCO_3 required to make 250 cm^3 of a standard solution of concentration $0,5 \text{ mol.dm}^{-3}$.
 5.4 What mass of citric acid would react completely with the volume of solution given in 5.3?



- 6 $1,95 \text{ g}$ of aluminium hydroxide Al(OH)_3 are required to neutralize 25 cm^3 of hydrochloric acid according to the following equation. $\text{Al(OH)}_3 + 3\text{HCl} \rightarrow \text{AlCl}_3 + 3\text{H}_2\text{O}$
- 6.1 How many moles of aluminium hydroxide react?
 6.2 How many moles of acid are dissolved in the 25 cm^3 ?
 6.3 Calculate the concentration of the acid.

- 7 $0,72 \text{ g}$ of oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$) is dissolved in water and titrated with sodium hydroxide. The end-point is reached with 50 cm^3 of NaOH solution. The following equation gives the acid/base reaction.



- 7.1 Is oxalic acid a strong or weak acid?
 7.2 Name a suitable indicator for this titration.
 Calculate the following:-
 7.3 The number of moles of oxalic acid used in the reaction.
 7.4 The number of moles of sodium hydroxide present in the 50 cm^3 of solution.
 7.5 The concentration of the sodium hydroxide solution.

- 8 A gardener grows a water plant in a pot that holds 100 dm^3 of water. The water from his pump is alkaline so he decides to add a little HCl to keep the water neutral. Unfortunately he adds too much acid and the pH of the water drops to 2 this means the concentration of the acid is $1 \times 10^{-2} \text{ mols.dm}^{-3}$.

- 8.1 Calculate the number of moles of HCl in the water.
 8.2 A friend advises him to add 212 g of sodium carbonate (Na_2CO_3) to restore neutrality
 8.2.1 Write a balanced equation for the reaction.
 8.2.2 Calculate the number of moles of sodium carbonate used.
 8.2.3 Will the water be neutral, acidic or basic after the sodium carbonate is dissolved?



EXERCISE 16.4

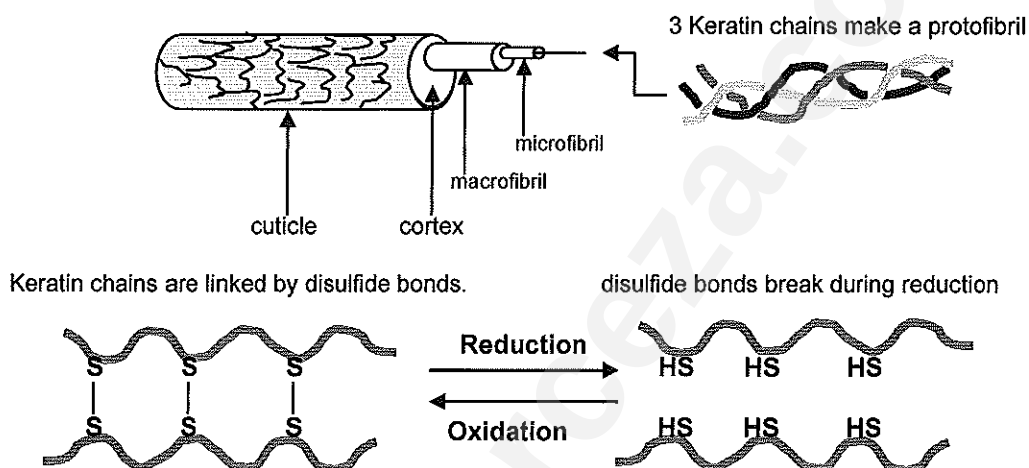
- In the reaction $X + H_2O \rightleftharpoons H_3O^+ + HSO_4^-$ X represents the following:-
A. the acid SO_4^{2-} B. the base SO_4^{2-} C. the acid H_2SO_4 D. the base H_2SO_4
- The pH of a certain solution is 8. What will be the pH if a few drops of sodium hydroxide are added?
A. more than 8 B. less than 8 C. 8 D. 0
- The pH of a $0,1 \text{ mol.dm}^{-3}$ sodium chloride solution will be
A. 1 B. 2 C. 7 D. 13
- Pure water is neutral at 25°C because
A. water is an ampholyte C. $[H_3O^+] = [OH^-]$
B. the water molecule is polar D. $[H_3O^+][OH^-] = 1 \times 10^{-14}$ at 25°C
- In a certain aqueous solution at 25°C the hydronium ion concentration is given by $[H_3O^+] = 5,0 \times 10^{-6} \text{ mol.dm}^{-3}$. What is the hydroxyl ion concentration of the solution in the units of mol.dm^{-3} ?
A. $2,0 \times 10^{-8}$ B. $2,0 \times 10^{-6}$ C. $2,0 \times 10^{-9}$ D. $5,0 \times 10^{-9}$
- Which of the following is a salt formed from the neutralization of a strong acid by a weak base?
A. CH_3COONa B. NH_4Cl C. $(CH_3COO)_2Mg$ D. Na_2SO_4
- Which one of the following statements regarding a strong acid is correct?
A. it has a high pH value C. it is always concentrated
B. it ionizes completely in water D. it becomes weaker when diluted.
- Consider the following ionisation equilibrium $2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
The ionisation constant of water (K_w) in the above reaction increases from 1×10^{-14} at 25°C to $9,6 \times 10^{-14}$ at 60°C . Which one of the following statements is correct?
A. $[H_3O^+] > [OH^-]$ at 60°C C. the pH increases with an increase in temperature
B. the reaction is exothermic D. the pH decreases with an increase in temperature
- When 10 cm^3 of HNO_3 is added to 990 cm^3 of pure water, which of the following statements is true?
A. $[OH^-]$ remains constant C. $[H_3O^+]$ remains constant
B. $[H_3O^+][OH^-]$ remains constant D. K_w increases.
- Which of the following indicators would be best to use in the titration of ethanoic acid against a solution of sodium hydroxide?
A. congo red; pH change 2,8 – 4,8 C. chlorophenol red; pH change 5,2 – 6,8
B. cresol red; pH change 7,0 – 9,1 D. methyl orange; pH change 3,1 – 4,4
- Which one of the following will exactly neutralize $0,1 \text{ dm}^3$ of sulfuric acid H_2SO_4 solution of concentration 1 mol.dm^{-3} ?
A. $0,1 \text{ mol LiOH}$ B. $0,1 \text{ mol NH}_3$ C. $0,1 \text{ mol Na}_2CO_3$ D. $0,1 \text{ mol KOH}$
- Which of the following statements concerning neutralization reactions are true?
I proton transfer occurs II the products are water and a salt
III the reaction is exothermic
A. I and II B. II and III C. I and III D. I, II and III
- Which of the following is NOT a neutralization reaction?
A. $H_3O^+ + OH^- \rightarrow 2H_2O$ C. $Zn + HCl \rightarrow H_2 + ZnCl_2$
B. $CuO + H_2SO_4 \rightarrow CuSO_4 + H_2O$ D. $LiOH + HCl \rightarrow LiCl + H_2O$

Curls or No Curls

At the heart of curly hair lies a covalent bond formed between two sulfur atoms. Hair consists of a fibrous protein called keratin which contains large numbers of the sulfur containing amino acid cysteine. The disulfide bonds between cysteine molecules cross link the chains of the protein. Three keratin chains twist together to make a protofibril; 11 protofibrils coil together to make a microfibril and hundreds of microfibrils pack together to make macrofibrils which form the inner cortex of the hair.

The cuticle or outer layer of the hair is made up of cells which overlap each other like the scales on a fish. These cells are cemented together by lipids (oils) and form an impenetrable layer which protects the cortex.

An electron microscope photo of a human hair



The disulfide bonds allow for some elasticity in hair. The keratin chains can slide past each other but return to their original positions when the stress is removed. In order to straighten or curl hair permanently, it is necessary to penetrate through the cuticle to the cortex and to alter its chemical structure by temporarily breaking the disulfide bonds.

To penetrate the cuticle the hair is first exposed to a high pH (alkaline) solution. This removes oils and causes the cuticle to swell and open.

Oxidation/reduction treatments are used in both perms and relaxers. They contain a reducing agent (ammonium thioglycolate) which breaks the disulfide linkages. After manipulating the hair to curl or straighten, an oxidizing agent (hydrogen peroxide) is applied. This allows the disulfide bonds to reform. The pH range for this treatment is about 9 to 10. So called acid perms use glycerol monothioglycolate (instead of ammonium thioglycolate) which allows for a lower pH (6.5 to 8.2) but requires heat.

Alkaline relaxers work at a very high pH (12 to 14). They completely de-fat the cuticle and randomly hydrolyze many of the proteins in the cortex. When the relaxer is washed out with a neutralizing shampoo some of the disulfide bonds re-form and set the straightened hair. At the same time the shampoo closes the cuticle. Lye containing relaxers use sodium hydroxide while no-lye relaxers use potassium hydroxide; lithium hydroxide or guanidine hydroxide. Although relaxers using sodium hydroxide act quickly and are able to straighten very curly hair, the caustic nature of this compound can damage hair leaving it brittle and coarse to the touch. In some cases this strong alkali also alters hair colour.

Where concentrations of hydroxides are high it is important that the product contains a "base". The term "base relaxer" means that the scalp and hair line must be coated with a protective coating such as petrolatum, mineral oil and lanolin, before applying the hair relaxer.

Always use reputable products in hair treatments, obey all instructions and allow time for the hair to recover between treatments.