Western Cape Government
BETTER TOGETHER.

CTSC
CapeTownScienceCentre

ACIDS AND BASES

Developed by the
Cape Town Science Centre

In collaboration with the
Western Cape Education Department

Understanding Acids and Bases

Acids and bases are found everywhere!

To better understand the chemistry of acid-base reactions, it is important to know the properties of acids and bases and the scientific models which define what is an acid and a base.

PROPERTIES

Acids

- Tastes sour
- It turns BLUE litmus paper RED
- Increases the concentration of hydrogen ions (H^+) in a solution
- Decreases the concentration of hydroxide ions (OH^-) in a solution
- It has a pH values of **LESS THAN 7**

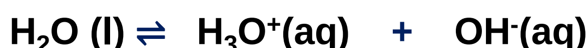
Bases

- Tastes bitter and has a soapy feel
- Turns RED litmus paper BLUE
- It has a pH value of **MORE THAN 7**
- Decreases the concentration of hydrogen ions (H^+) in a solution
- Increases the concentration of hydroxide ions (OH^-) in a solution

Scientific Models

Arrhenius Theory – Only explains acids & bases when dissolved **IN WATER**

Arrhenius noticed that water dissociates (splits up) into hydronium and hydroxide ions according to the following reaction:



$$H_2O\,(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

- **Acid** – a substance that produces H^+/H_3O^+ ions in an aqueous solution
- **Base -** a substance that produces OH^- ions in an aqueous solution

Bronsted-Lowry Theory – Explains acid & bases in both **SOLID** and **LIQUID PHASE**

Bronsted and Lowry broadened the acid/base definition of Arrhenius to not need water

- **Acid** is a proton (H^+) **DONOR**
- **Base** is a proton (H^+) **ACCEPTOR**

The proton exchange, called *protolysis*, is simultaneous

Proton transfer reaction – General equation during acid-base reaction:

$$HA + B \longrightarrow BH^+ + A^- \quad \text{OR} \quad HA + B^- \longrightarrow BH + A^-$$

Western Cape Government
BETTER TOGETHER.

CTSC
CapeTownScienceCentre

Conjugate Acid-Base Pairs

The Lowry-Bronsted Theory involves an acid-base protolytic reaction in which a proton transfer takes place. This proton transfer is simultaneous!

Therefore a pair of substances will differ from one another by a proton within an acid-base reaction. This pair is called a **CONJUGATE ACID-BASE PAIR**.

Conjugate comes from the Latin word *coniugātiō* which means to "yoke together".

When an ACID donates a proton, a CONJUGATE BASE is produced.

When a BASE accepts a proton, a CONJUGATE ACID is produced.

When a **BASE** has accepted a proton, the formed product is called a **CONJUGATE ACID** because it can donate a proton in the reverse reaction again

The conjugate acid of an base

The conjugate base of an acid

When an **ACID** has donated a proton, the remaining ion is called a **CONJUGATE BASE** because it can accept a proton in the reverse reaction again

EXAMPLES

Remove a Proton from the acid	
ACID	CONJUGATE BASE
H_2O	OH^-
HCl	Cl^-
HSO_4^-	SO_4^{2-}
HPO_4^{2-}	PO_4^{3-}

Add a proton to the base	
BASE	CONJUGATE ACID
H_2O	H_3O^+
NH_3	NH_4^+
HSO_4^-	H_2SO_4
SO_4^{2-}	HSO_4^-

AMPHIPROTIC substances (ampholyte) are able to react as either an acid or a base.

In presence of a STRONG acid, an amphiprotic substance reacts as a base.

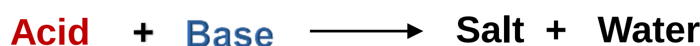
In presence of a STRONG base, an amphiprotic substance reacts as a acid.

Western Cape Government
BETTER TOGETHER.

CTSC
CapeTownScienceCentre

Reactions with metals

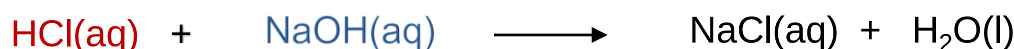
The general reaction mechanism for acid base reaction, results in the formation of a **salt and water**, regardless of what acid or base was used.



A salt is a compound made up of a metal and non-metal portion. It is a product of an acid-base reaction where hydrogen in the acid molecule is replaced by a metal cation of the base.

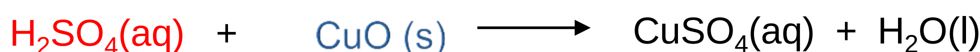
Acid and metal Hydroxide

Example

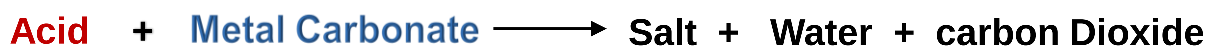


Acid and Metal Oxide

Example

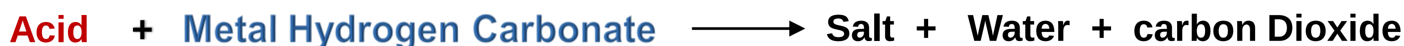


Acid and Metal Carbonate



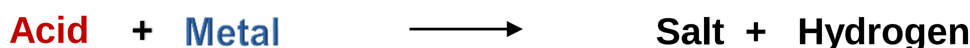
Example

Acid and Metal Hydrogen Carbonate

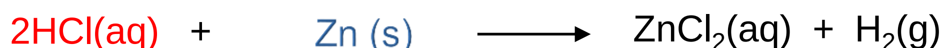


Example

Acid and Metal



Example



Western Cape Government
BETTER TOGETHER.

CTSC
CapeTownScienceCentre

Understanding Acid-Base Strength

The strength is important in understanding acid-base chemistry.

The strength of an acid or base refers to extent of *ionisation* or *dissociation* that takes place in a solution.

Acids are molecular structures (covalent), which will undergo **ionisation**.

Bases are ionic structures, which will undergo **dissociation**.

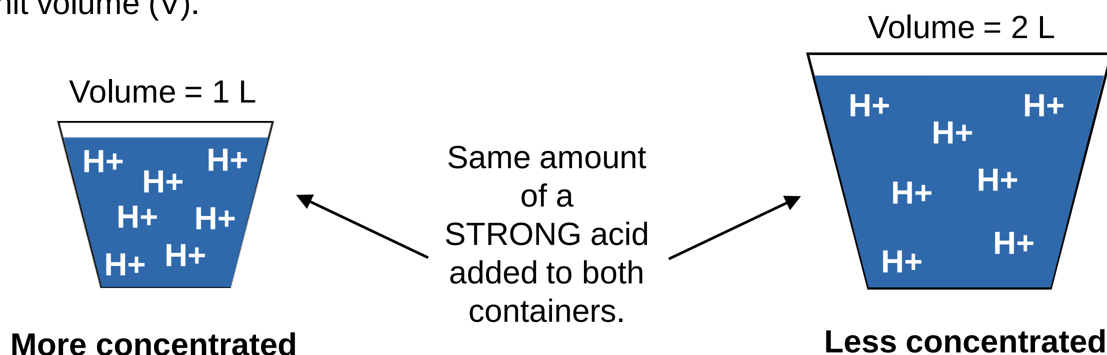
Ionisation – Chemical process where covalent molecules produce ions in solution.

Dissociation – Chemical process where ionic compounds produce ions in solution.

Strong acids ionise **completely** in solution to form a **high** concentration of **H_3O^+ ions**	**Weak acids** ionise **incompletely** in solution to form a **low** concentration of **H_3O^+ ions**
Examples Hydrochloric acid (HCl) Sulfuric acid (H_2SO_4) Nitric Acid (HNO_3)	**Examples** Ethanoic acid (CH_3COOH) Hydrofluoric acid (HF) Phosphoric acid (H_3PO_4)

Strong bases dissociate **completely** in solution to form a **high** concentration of **OH^- ions**	**Weak bases** dissociate **incompletely** in solution to form a **low** concentration of **OH^- ions**
Examples: Sodium hydroxide (NaOH) Potassium hydroxide (KOH) Lithium hydroxide (LiOH)	**Examples:** Ammonium hydroxide (NH_4OH) Calcium hydroxide ($Ca(OH)_2$) Magnesium hydroxide ($Mg(OH)_2$)

Acid/Base strength must NOT be confused with **concentration** (c) which refer to the amount of acid/base with certain volume of solution, defined as the number of moles (n) per unit volume (V).



How concentrated or dilute an **acid** or **base** may be is a measure of the amount of water present in the system.

Western Cape Government
BETTER TOGETHER.

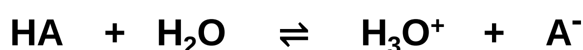
CTSC
CapeTownScienceCentre

Identifying Strong & Weak Acids/Bases

The strength of acids and bases can be identified by using the Equilibrium Constant

Strong and Weak Acid

When acids are dissolved in water, they ionise according to their general equation:



The equilibrium constant is:

$$K_c = \frac{[H_3O^+][A^-]}{[HA]} = K_a$$

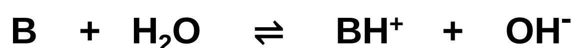
As this equilibrium is focused only on acids, the K_c becomes **K_a**, which is the **ionisation constant of an acid**.

For a strong acid, where acid ionises completely, the K_a value is high (>1).
This is because the denominator concentration [HA] is low and the numerator concentration $[H_3O^+][A^-]$ is high.

For a weak acid, where acid ionises partially, the K_a value is low (<1).
This is because the denominator concentration [HA] is high and the numerator concentration $[H_3O^+][A^-]$ is low.

Strong and Weak Base

When acids are dissolved in water, they ionise according to their general equation:



The equilibrium constant is:

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

As this equilibrium is focused only on bases, the K_c becomes **K_b**, which is the **ionisation constant of a base**.

For a strong base, where the base dissociates completely, the K_b value is high (>1).
This is because the denominator concentration [B] is low and the numerator concentration $[BH^+][OH^-]$ is high.

For a weak base, where the base dissociates partially, the K_b value is low (<1).
This is because the denominator concentration [B] is high and the numerator concentration $[BH^+][OH^-]$ is low.

Equilibrium Constant for Water (K_w)

Water is an amphiprotic substance, which is able to act as both an acid and a base.

Two water molecules can undergo auto-protolysis or auto-ionisation where two molecules react with one another and were one acts an acid (H^+) and the other a base (proton acceptor).

$$H_2O\ (l) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$$

The equilibrium constant is:

$$K_c = [H_3O^+][OH^-] = K_w$$

As this equilibrium is focused only on auto-ionisation of water, the K_c becomes K_w, which is the **ionisation constant of water**.

In pure water, $[H_3O^+] = 1 \times 10^{-7}$ mol.dm^{-7} and $[OH^-] = 1 \times 10^{-7}$ mol.dm^{-7}

Therefore $K_w = [H_3O^+].[OH^-] = 1 \times 10^{-14}$ at room temperature

The auto-ionisation process of water is weak as evidenced by the extremely low value of 1×10^{-14}

The pH Scale

Due to the low concentrations of the hydroxide and hydronium ions, it is simpler to refer to their negative logarithm, which allows us to work with whole numbers.

This is the pH scale, ranging from 0 to 14, and indicates the degree of acidity of a solution.

$$pH = -\log [H_3O^+] \qquad pOH = -\log [OH^-]$$

$$pH + pOH = 14$$

Acidic Solution	Neutral Solution	Basic Solution
$[H_3O^+] > [OH^-]$	$[H_3O^+] = [OH^-]$	$[H_3O^+] < [OH^-]$
$[H_3O^+] > 1x10^{-7}$	$[H_3O^+] = 1x10^{-7}$	$[H_3O^+] < 1x10^{-7}$

The pH of a substance can only be determined when it is in an **aqueous solution.**

pH	Examples of solutions
0	Battery acid, strong hydrofluoric acid
1	Hydrochloric acid secreted by stomach lining
2	Lemon juice, gastric acid, vinegar
3	Grapefruit juice, orange juice, soda
4	Tomato juice, acid rain
5	Soft drinking water, black coffee
6	Urine, saliva
7	"Pure" water
8	Sea water
9	Baking soda
10	Great Salt Lake, milk of magnesia
11	Ammonia solution
12	Soapy water
13	Bleach, oven cleaner
14	Liquid drain cleaner

http://cnx.org/content/col11496/1.6/

Indicators

An indicator is substance that changes colour in the presence of an acid or base.

Indicator	Colour in acid	Colour in base	Range
Methyl orange	Orange	Yellow	3.1 – 4.4
Methyl red	Red	Yellow	4.4 – 6.2
Bromothymol blue	Yellow	Blue	6 – 7.6
Phenolphthalein	Colourless	Pink	8.3 – 10

ACID RED   BASE BLUE

Litmus turns red/pink in an acidic solution and blue in a basic solution

pH Calculations

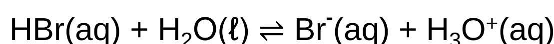
Titrations is an experimental technique used to determine the concentration of an acid or a base using a standard solution.

Using volumetric analysis, the unknown concertation of a solution (acid or base) may be determined.

What to Consider When Calculating the pH
Use the equations for pH
Use the equation $$pH = -\log[H_3O^+]$$ Other useful equations include $$[H_3O^+][OH^-] = 1 \times 10^{-14}$$ $$pH = 14 - p[OH^-]$$ $$pH = 14 - (-\log[OH^-])$$
Use the equations for concentration
Use the equation $$c = \frac{n}{V} = \frac{mol}{dm^3}$$ Remember moles (n) can be calculated using mass of a substance (m) and its molar mass (M): $$n = \frac{m}{M}$$
Use Mole Ratios
○ Write down the full balanced reaction ○ Identify the acid/base

WORKED Exam Question — Paper 2, Oct/Nov 2019, Q.7

A hydrogen bromide solution, HBr(aq), reacts with water according to the following balanced chemical equation:



$$HBr(aq) + H_2O(\ell) \rightleftharpoons Br^-(aq) + H_3O^+(aq)$$

The K_a value of HBr(aq) at 25 °C is 1 x 10^9.

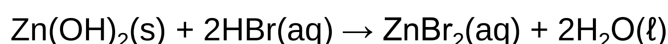
7.1 Is hydrogen bromide a STRONG ACID or a WEAK ACID? Give a reason for the answer. (2)

Strong acid. Large K_a value K_a > 1 (HBr) ionises completely

7.2 Write down the FORMULAE of the TWO bases in the above reaction. (2)

H_2O and Br^-

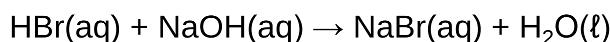
7.3 HBr(aq) reacts with $Zn(OH)_2$(s) according to the following balanced equation:



$$Zn(OH)_2(s) + 2HBr(aq) \rightarrow ZnBr_2(aq) + 2H_2O(\ell)$$

An unknown quantity of $Zn(OH)_2$(s) is reacted with 90 cm^3 of HBr(aq) in a flask.
(Assume that the volume of the solution does not change during the reaction.)

The EXCESS HBr(aq) is then neutralised by 16,5 cm^3 of NaOH(aq) of concentration 0,5 $mol\cdot dm^{-3}$. The balanced equation for the reaction is:



$$HBr(aq) + NaOH(aq) \rightarrow NaBr(aq) + H_2O(\ell)$$

7.3.1 Calculate the pH of the HBr solution remaining in the flask AFTER the reaction with $Zn(OH)_2$(s). (7)

$n(NaOH)_{reacted}$ = cV = 0.5 (0.0165) = 0.00825 mol

$n(HBr)_{excess}$ = n(NaOH) = 0.00825 mol

$c(H_3O^+) = \frac{n}{V} = \frac{0.00825}{0.09}$ = 0.092 $mol.dm^{-3}$

pH = $-\log[H_3O^+]$

= -log(0.092)

= 1.04

Continued… Paper 2, Oct/Nov 2019, Q.7

7.3.2 Calculate the mass of $Zn(OH)_2$(s) INITIALLY present in the flask if the initial concentration of HBr(aq) was 0,45 mol·dm^{-3}. (6)

$n(HBr)_{initial}$ = cV

= 0.45 (0.09)

= 0.0405 mol

n(HBr reacted with $Zn(OH)_2$) = 0.0405 – 0.00825 = 0.03224 mol

$n(Zn(OH)_2) = \frac{1}{2}$ n(HBr)

$= \frac{1}{2}$ (0.03224)

= 0.016125 mol

$m(Zn(OH)_2)$ = nM

= 0.016125 mol (99)

= 1.596 g

Past Exam Question Paper 2, May/June 2019, Q.7

7.1 Define a *base* in terms of the Arrhenius theory. (2)

7.2 Explain how a *weak base* differs from a *strong base*. (2)

7.3 Write down the balanced equation for the hydrolysis of $NaHCO_3$. (3)

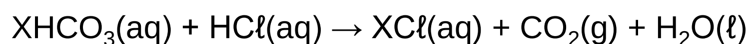
7.4 A learner wishes to identify element **X** in the hydrogen carbonate, $XHCO_3$. To do this she dissolves 0,4 g of $XHCO_3$ in 100 cm^3 of water. She then titrates all of this solution with a 0,2 mol dm^{-3} hydrochloric acid (HCℓ) solution. Methyl orange is used as the indicator during the titration.

7.4.1 Calculate the pH of the hydrochloric acid solution. (3)

7.4.2 Give a reason why methyl orange is a suitable indicator in this titration. (1)

At the endpoint she finds that 20 cm^3 of the acid neutralised ALL the hydrogen carbonate solution. The balanced equation for the reaction is:

$$XHCO_3(aq) + HC\ell(aq) \rightarrow XC\ell(aq) + CO_2(g) + H_2O(\ell)$$



7.4.3 Identify element X by means of a calculation. (6)

Past Exam Question

Paper 2, Oct/Nov 2017, Q.7

7.1 Ammonia ionises in water to form a basic solution according to the following balanced equation:

$$NH_3(g) + H_2O(\ell) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

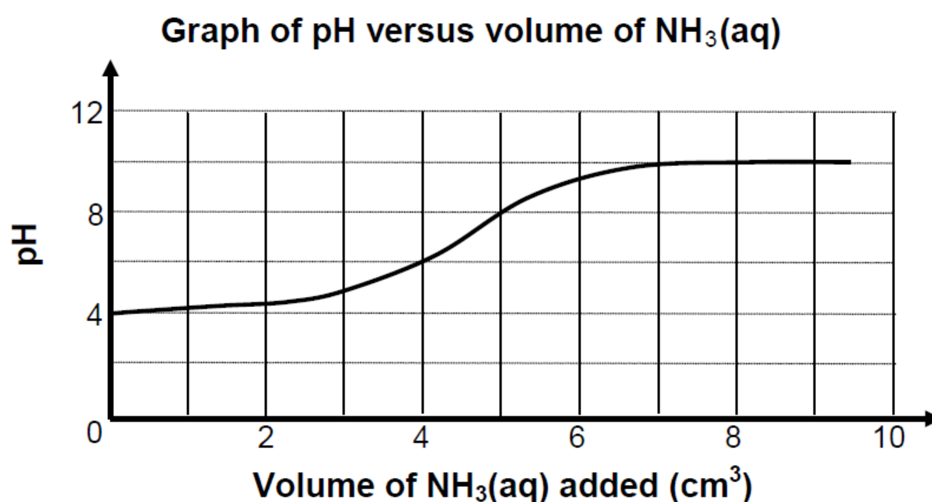
7.1.1 Is ammonia a WEAK or a STRONG base? Give a reason for the answer. (2)

7.1.2 Write down the conjugate acid of NH_3 (g). (1)

7.1.3 Identify ONE substance in this reaction that can behave as an ampholyte in some reactions. (1)

7.2 A learner adds distilled water to a soil sample and then filters the mixture. The pH of the filtered liquid is then measured.

He then gradually adds an ammonia solution, NH_3(aq), to this liquid and measures the pH of the solution at regular intervals. The graph below shows the results obtained.



7.2.1 Is the soil sample ACIDIC or BASIC? Refer to the graph above and give a reason for the answer. (2)

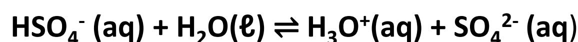
7.2.2 Calculate the concentration of the hydroxide ions (OH^-) in the reaction mixture after the addition of 4 cm^3 of NH_3 (aq). (4)

Past Exam Question

Paper 2, Oct/Nov 2018, Q.7

7.1 Sulphuric acid is a strong acid present in acid rain. It ionises in two steps as follows:

$$H_2SO_4(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HSO_4^-(aq)$$



$$HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

7.1.1 Define an *acid* in terms of the Lowry-Brønsted theory. (2)

7.1.2 Write down the FORMULA of the conjugate base of $H_3O^+(aq)$ (1)

7.1.3 Write down the FORMULA of the substance that acts as an ampholyte in the ionisation of sulphuric acid. (2)

7.2 Acid rain does not cause damage to lakes that have rocks containing limestone ($CaCO_3$). Hydrolysis of $CaCO_3$ results in the formation of ions, which neutralise the acid.

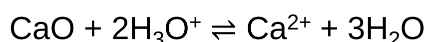
7.2.1 Define *hydrolysis* of a salt. (2)

7.2.2 Explain, with the aid of the relevant HYDROLYSIS reaction, how limestone can neutralise the acid. (3)

7.3 The water in a certain lake has a pH of 5.

7.3.1 Calculate the concentration of the hydronium ions in the water. (3)

The volume of water in the lake is 4×10^9 dm^3. Lime, CaO, is added to the water to neutralise the acid according to the following reaction:



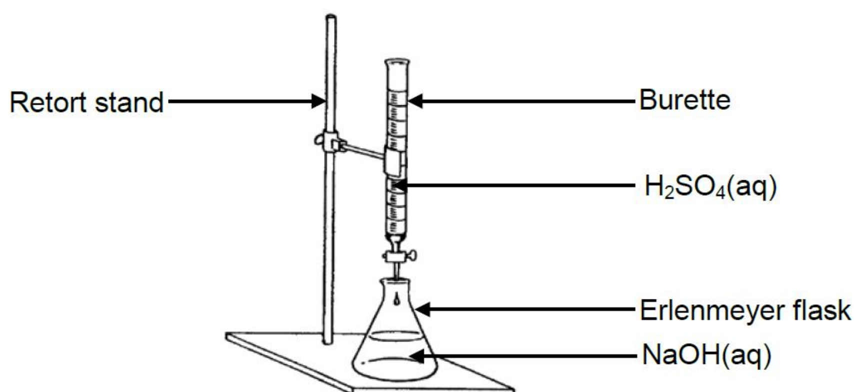
$$CaO + 2H_3O^+ \rightleftharpoons Ca^{2+} + 3H_2O$$

7.3.2 If the final amount of hydronium ions is $1{,}26 \times 10^3$ moles, calculate the mass of lime that was added to the lake. (7)

Past Exam Question

Paper 2, MayJune 2018, Q.7

The reaction between a sulphuric acid (H_2SO_4) solution and a sodium hydroxide (NaOH) solution is investigated using the apparatus illustrated below.



7.1 Write down the name of experimental procedure illustrated above. (1)

7.2 What is the function of the burette? (1)

7.3 Define an acid in terms of the Arrhenius theory. (2)

7.4 Give a reason why sulphuric acid is regarded as a strong acid. (1)

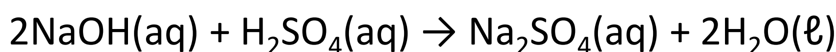
7.5 Bromothymol Blue is used as an indicator. Write down the colour change that will take place in the Erlenmeyer flask on reaching the endpoint of the titration.

Choose from the following:

BLUE TO YELLOW YELLOW TO BLUE GREEN TO YELLOW

During the titration a learner adds 25 cm^3 of NaOH(aq) of concentration 0,1 $mol{\cdot}dm^{-3}$ to an Erlenmeyer flask and titrates this solution with H_2SO_4(aq) of concentration 0,1 $mol{\cdot}dm^{-3}$.

The balanced equation for the reaction that takes place is:



$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(\ell)$$

7.6 Determine the volume of H_2SO_4 (aq) which must be added to neutralise the NaOH(aq) in the Erlenmeyer flask completely. (4)

7.7 If the learner passes the endpoint by adding 5 cm^3 of the same H_2SO_4(aq) in excess, calculate the pH of the solution in the flask. (7)