education

Department of
Education
FREE STATE PROVINCE

PHYSICAL SCIENCES TRAINING MANUAL
CAPS

ACIDS AND BASES
GRADE 12

This document consists 33 of pages

Prescribed content

Acid-base theories

- Define *acids* and *bases* according to Arrhenius.
- Define acids and bases according to Lowry-Brønsted:

Reactions of acids and bases with water

- Distinguish between *strong acids/bases* and *weak acids/bases* with examples.
- Distinguish between *concentrated acids/bases* and *dilute acids/bases*.
- Write down the reaction equations of aqueous solutions of acids and bases.

  Examples: $HC\ell(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + C\ell^-(aq)$ (HCℓ is a monoprotic acid.)

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

  $H_2SO_4(aq) + 2H_2O(\ell) \rightarrow 2H_3O^+(aq) + SO_4^{2-}(aq)$ (H_2SO_4 is a diprotic acid.)

Conjugate acid-base pairs

- Identify conjugate acid-base pairs for given compounds.
- When the acid, HA, loses a proton, its conjugate base, A^-, is formed. When the base, A^-, accepts a proton, its conjugate acid, HA, is formed. These two are a conjugate acid-base pair.

Amphiprotic substances

Describe a substance that can act as either acid or base as amphiprotic or as an ampholyte. Water is a good example of an ampholyte. Write equations to show how an amphiprotic substance can act as acid or base.

Neutralisation reactions

Write down neutralisation reactions of common laboratory acids and bases.

Examples:

$HC\ell(aq) + NaOH(aq)/KOH(aq) \rightarrow NaC\ell(aq)/KC\ell(aq) + H_2O(\ell)$

$HC\ell(aq) + Na_2CO_3(aq) \rightarrow NaC\ell(aq) + H_2O(\ell) + CO_2(g)$

$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(\ell)$

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

$(COOH)_2(aq) + NaOH(aq) \rightarrow (COO)_2Na_2(aq) + H_2O(\ell)$

$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(\ell)$

NOTE: The above are examples of equations that candidates should be able to write from given information. However, any other neutralisation reaction can be given in the question paper to assess, e.g., stoichiometry calculations.

Hydrolysis

- Determine the approximate pH (equal to, smaller than or larger than 7) of salts in salt hydrolysis. Define *hydrolysis.*
- Hydrolysis of the salt of a weak acid and a strong base results in an alkaline solution, i.e. the pH > 7.
- Hydrolysis of the salt of a strong acid and a weak base results in an acidic solution, i.e. the pH < 7.
- The salt of a strong acid and a strong bases does not undergo hydrolysis and the solution of the salt will be neutral, i.e. pH = 7.

Acid-base titrations

- Motivate the choice of a specific indicator in a titration. Choose from methyl orange, phenolphthalein and bromothymol blue. Define the *equivalence point*.
- Define the *endpoint* of a titration.
- Perform stoichiometric calculations based on titrations of a strong acid with a strong base, a strong acid with a weak base and a weak acid with a strong base. Calculations may include percentage purity.
- For a titration, e.g. the titration of oxalic acid with sodium hydroxide:
  - List the apparatus needed or identify the apparatus from a diagram.
  - Describe the procedure to prepare a standard oxalic acid solution.
  - Describe the procedure to conduct the titration.
  - Describe safety precautions.
  - Describe measures that need to be in place to ensure reliable results.
  - Interpret given results to determine the unknown concentration.

Auto-ionisation of water
• Define K_w as the equilibrium constant for the ionisation of water or the ionic product of water or the ionisation constant of water, i.e. $K_w = [H_3O^+][OH^-] = 1 \times 10^{14}$ by 298 K. • Explain the *auto-ionisation of water*, i.e. the reaction of water with itself to form H_3O^+ ions and OH^- ions.
pH calculations
• Explain the pH scale as a scale of numbers from 0 to 14 used to express the acidity or alkalinity of a solution. • Calculate pH values of strong acids and strong bases using $pH = -\log[H_3O^+]$.
Comparison of strong and weak acids
• Interpret K_a values of acids to determine the relative strength of given acids. • Interpret K_b values of bases to determine the relative strength of given bases. • Compare strong and weak acids by looking at: o pH (monoprotic and diprotic acids) o Conductivity o Reaction rate

IMPORTANT TERMS AND DEFINITIONS	
Acid-base indicator	A dye used to distinguish between acidic and basic solutions by means of the colour changes it undergoes in these solutions.
Amphiprotic substance OR ampholyte	A substance that can act as either an acid or a base.
Arrhenius theory	An acid is a substance that produces hydrogen ions (H^+)/ hydronium ions (H_3O^+) when it dissolves in water. A base is a substance that produces hydroxide ions (OH^-) when it dissolves in water.
Auto-ionisation of water	A reaction in which water reacts with itself to form ions (hydronium ions and hydroxide ions).
Concentrated acids/bases	Contain a large amount (number of moles) of acid/base in proportion to the volume of water.
Conjugate acid-base pair	A pair of compounds or ions that differ by the presence of one H^+ ion. Example: CO_3^{2-} and HCO_3^- OR $HCℓ$ and $Cℓ^-$
Conjugate acid and base	A conjugate acid has one H^+ ion more than its conjugate base. Example: HCO_3^- is the conjugate acid of base CO_3^{2-}. CO_3^{2-} is the conjugate base of acid HCO_3^-.
Dilute acids/bases	Contain a small amount (number of moles) of acid/base in proportion to the volume of water.
Diprotic acid	An acid that can donate two protons. Example: H_2SO_4
Dissociation	The process in which ionic compounds split into ions.
Endpoint	The point in a titration where the indicator changes colour.
Equivalence point	The point in a reaction where equivalent amounts of acid and base have reacted completely.
Hydrolysis	The reaction of a salt with water. OR The reaction of an ion with water to produce the conjugate acid and a hydroxide ion or the conjugate base and a hydronium ion.
Ionisation	The process in which ions are formed during a chemical reaction.
Ion product of water	The product of the ions formed during auto-ionisation of water i.e. $[H_3O^+][OH^-]$ at 25 °C.
Ionisation constant of water (K_w)	The equilibrium value of the ion product $[H_3O^+][OH^-]$ at 25 °C.
K_a value	Ionisation constant for an acid

IMPORTANT TERMS AND DEFINITIONS	
K_b value	Dissociation or ionisation constant for a base
Lowry-Brønsted theory	An acid is a proton (H^+ ion) donor. A base is a proton (H^+ ion) acceptor.
Monoprotic acid	An acid that can donate one proton. Example: $HC\ell$
Neutralisation	The reaction of an acid with a base to form a salt (ionic compound) and water.
pH	The negative of the logarithm of the hydronium ion concentration in $mol \cdot dm^{-3}$. In symbols: $pH = -log[H_3O^+]$
pH scale	A scale from 0 – 14 used as a measure of the acidity and basicity of solutions where pH = 7 is neutral, pH > 7 is basic and pH < 7 is acidic.
Salt	The ionic compound that is the product of a neutralisation reaction.
Standard solution	A solution of precisely known concentration.
Strong bases	Dissociate completely in water to form a high concentration of OH^- ions. Examples: sodium hydroxide and potassium hydroxide.
Strong acids	Ionise completely in water to form a high concentration of H_3O^+ ions. Examples: hydrochloric acid, sulphuric acid and nitric acid
Titration	The procedure for determining the amount of acid (or base) in a solution by determining the volume of base (or acid) of known concentration that will completely react with it.
Weak acids	Ionise incompletely in water to form a low concentration of H_3O^+ ions. Examples: ethanoic acid and oxalic acid
Weak bases	Dissociate/ionise incompletely in water to form a low concentration of OH^- ions. Examples: ammonia, sodium carbonate, potassium carbonate, calcium carbonate and sodium hydrogen carbonate

Acids and bases

An acid-base reaction is any reaction in which **protons (H^+ ions) are transferred**. Acid-base reactions are also referred to as **protolytic reactions**.

1. Acid-base theories

1.1 Arrhenius theory of acids and bases

Arrhenius classified acids and bases in terms of their formulae and their behaviour in water.

- **Acids** are substances which **produce hydrogen ions (H^+ ions) in solution.**
- **Bases** are substances which **produce hydroxide ions (OH^- ions) in solution.**

Example:
An acid: $HC\ell(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + C\ell^-(aq)$
$HC\ell(g)$ is an acid because it has H in its formula and ionises in water to form hydronium ions, $H_3O^+(aq)$.

Note: An H^+ ion (the nucleus of the hydrogen atom) cannot exist separately in water. Therefore, when an acid is dissolved in water, the proton produced by the acid combines with water to produce the hydronium ion, H_3O^+.

A base: $NaOH(s) \xrightarrow{H_2O(\ell)} Na^+(aq) + OH^-(aq)$
$NaOH(s)$ is a base because it has OH in its formula and dissociates in water to form hydroxide ions, $OH^-(aq)$.

Limitations of the Arrhenius theory:
Some substances do not have discrete OH^- ions, but still behave like bases. Examples are NH_3 and Na_2CO_3. However, both will form OH^- ions in solution. The Arrhenius theory is limited to solutions and water has to be the solvent for acid-base reactions.

1.2 Lowry-Brønsted theory of acids and bases

This theory addressed the limitation of the Arrhenius theory which only classifies acids and bases when dissolved in water.

Lowry-Brønsted theory:
- **Acids** are **proton (H^+) donors**.
- **Bases** are **proton (H^+) acceptors**.

(The Lowry- Brønsted theory was named after Johannes Brønsted from Denmark and Thomas Lowry from England who independently suggested this theory in 1923.).

According to this theory, the only requirement for and acid-base reaction is that one species donates a proton and another species accepts a proton. Acid-base reactions can thus occur between gases, in non-aqueous solutions, in heterogeneous mixtures and in aqueous solution.

Examples:
- $HC\ell(g) + H_2O(\ell) \rightarrow H_3O^+(aq) + C\ell^-(aq)$
  acid base

In the forward reaction, $HC\ell$ loses a proton (H^+ ion) and $C\ell^-$ is formed. H_2O gains a proton (H^+ ion) and H_3O^+ is formed. $HC\ell$ is the proton donor and is thus the acid. H_2O is the proton acceptor and thus is the base.

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

In the forward reaction, H_2O loses a proton (H^+ ion) and OH^- is formed. NH_3 gains a proton (H^+ ion) and NH_4^+ is formed. H_2O is the proton donor and is thus the acid. NH_3 is the proton acceptor and thus is the base.

2. Reactions of acids and bases with water

2.1 Monoprotic and diprotic acids

HCℓ can **donate only one proton** (H^+ ion) in a reaction and is called a **monoprotic acid**.
HCℓ(aq) + H_2O(ℓ) → H_3O^+(aq) + $Cℓ^-$(aq)
(**1** mole of acid reacts with 1 mole of water to form **1** mole of H_3O^+ ions.)

Sulphuric acid **donate two protons** (two H^+ ions) in a reaction and is called a **diprotic acid**.

H_2SO_4(aq) + $2H_2O$(ℓ) → **2**H_3O^+(aq) + SO_4^{2-}(aq)
(**1** mole of acid reacts with 2 moles of water to form **2** moles of H_3O^+ ions.)

2.2 Strong and weak acids

When an acid dissolves in water, a proton (H^+ ion) is transferred to a water molecule to produce a hydronium ion and a negative ion depending on what acid you are starting from. In the general case: $HA + H_2O \rightleftharpoons H_3O^+ + A^-$

These reactions are reversible and equilibrium constants can be written for these reactions. In the case of acids, the equilibrium constant is called the **ionisation constant of acids** and the symbol $\mathbf{K_a}$ is used.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \qquad \text{(ionisation constant for acids)}$$

In some cases, the acid is so good at giving away hydrogen ions that we can think of the reaction as being one-way. The acid is virtually 100% ionised. In such cases we write the reaction with a single arrow showing that it proceeds basically in one direction and K_a will be very large because the concentration of products is very high and that of reactants is very low.

The strength of an acid depends on the type of acid and cannot be changed in the laboratory. For example, hydrochloric acid will always be a strong acid, regardless of how much water is present in a hydrochloric acid solution.

2.2.1 Strong acids

A strong acid **ionises completely** in water.

Strong acids are very good at giving away hydrogen ions that almost 100% of the acid ionises to produce hydronium ions and negative ions. We can write the equilibrium reactions of strong acids with a single arrow because the reactions proceed almost completely. **The ionisation constants (K_a) of strong acids are very large.**

The strong acids are hydrochloric acid (HCℓ), sulphuric acid (H_2SO_4) and nitric acid (HNO_3).

Example: HCℓ(aq) + H_2O(ℓ) → H_3O^+(aq) + $Cℓ^-$(aq)

HCℓ(g) is almost 100% ionised i.e. almost all the HCℓ(g) reacts with water to form ions. After ionisation no HCℓ(g) is present. The reaction is written with a single arrow representing an equilibrium position that lies to far right. K_a is very large.

All acids that ionise completely with K_a greater than one (K_a > 1), are referred to as strong acids. K_a values of H_2SO_4, HCℓ and HNO_3 strong acids are very large and no values are given for K_a. The K_a value for the hydronium ion (H_3O^+) is 1.

2.2.2 Weak acids

A weak acid **ionises incompletely** in water.

Weak acids are not good at giving away hydrogen ions an only a small percentage of the acid ionises to produce hydronium ions and negative ions. An equilibrium mixture is formed. **The ionisation constants (K_a) of weak acids are very small.**
The weak acids are ethanoic acid (CH_3COOH) and oxalic acid, $(COOH)_2$.

Example: $CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

CH_3COOH ionises to a very small extent to produce hydronium ions and ethanoate ions. The reverse reaction is more successful than the forward one. The ions react very easily to reform the acid and the water. At any one time, only about 1% of the ethanoic acid molecules have converted into ions. The rest remain as simple ethanoic acid molecules. An equilibrium mixture is formed. After ionisation, some of the CH_3COOH is still present.

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

All acids that ionise incompletely or only partially, with K_a smaller than one (Ka < 1), are referred to as weak acids.

Table 1: K_a values for some weak acids

Acid	Formula	K_a at 25 °C
Sulphurous acid	H_2SO_3	$1,2 \times 10^{-2}$
Hydrogen sulphate ion	HSO_4^-	$1,2 \times 10^{-2}$
Oxalic acid	$(COOH)_2$	$5,6 \times 10^{-2}$
Ethanoic acid (acetic acid)	CH_3COOH	$1,8 \times 10^{-5}$
Carbonic acid	H_2CO_3	$4,2 \times 10^{-7}$
Hydrogen sulphide	H_2S	1×10^{-7}
Ammonium ion	NH_4^+	$5,6 \times 10^{-10}$
Hydrogen carbonate ion	HCO_3^-	$4,8 \times 10^{-11}$

2.3 Strong and weak bases

When a base dissolves in water, it dissociates/ionises to form hydroxide ion (OH^- ions) in solution. In the general case: $B + H_2O \rightleftharpoons BH^+ + OH^-$

These reactions are reversible and equilibrium constants can be written for these reactions. In the case of bases, the equilibrium constant is called the **dissociation/ionisation constant of bases** and the symbol K_b is used.

$$K_b = \frac{[BH^+][OH^-]}{[B]} \qquad \text{(dissociation/ionisation constant for bases)}$$

Note: *Bases such as NaOH and KOH are ionic and dissociate in water i.e. solid ionic crystals are broken up into ions. Bases such as NH_3, ionise in water i.e. the molecule reacts with water to form ions in solution. Therefore K_b can be, depending on the base, referred to as either the dissociation constant or the ionisation constant for a base.*

Some bases dissociate in water to form 100% hydroxide ions in solution. Then the reaction is one-way and we write the equation with a single arrow showing that it proceeds in one direction. K_b will be very large because the concentration of products is very high and that of reactants is very low.

The strength of a base depends on the type of base and cannot be changed in the laboratory. For example, sodium hydroxide will always be a strong base, regardless of how much water is present in a sodium hydroxide solution.

2.3.1 Strong bases

A strong base **dissociates completely** in water.

Metal hydroxides, such as sodium hydroxide, are full ionic and split up 100% into metal ions and hydroxide ions in solution.

Example: $NaOH(s) \xrightarrow{H_2O(\ell)} Na^+(aq) + OH^-(aq)$

Each mole of sodium hydroxide dissolves to give a mole of hydroxide ions in solution. The equilibrium reactions of strong bases with a single arrow because the reactions proceed completely. **The dissociation constants (K_b) of strong bases are very large.**

The strong bases are sodium hydroxide (NaOH) and potassium hydroxide (KOH).

Example: $NaOH(s) \xrightarrow{H_2O(\ell)} Na^+(aq) + OH^-(aq)$

NaOH(s) dissociates 100% in water i.e. all the NaOH(s) forms OH^- ions in water. The reaction is written with a single arrow representing an equilibrium position that lies to far right. $K_b = [Na^+]]OH^-]$ is larger than 1.

All bases that dissociate/ionise completely with K_b greater than one ($K_b > 1$), are referred to as strong bases.

2.3.2 Weak bases

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

NH_3 ionises to a very small extent to produce ammonium ions and hydroxide ions. The reverse reaction is more successful than the forward reaction. The ions react very easily to reform the base and the water. After ionisation some of the NH_3 is still present.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

All bases that ionise incompletely or only partially with K_b smaller than one ($K_b < 1$), are referred to as weak bases.

The weak bases are ammonia (NH_3), sodium carbonate (Na_2CO_3), potassium carbonate (K_2CO_3), sodium hydrogen carbonate ($NaHCO_3$) and calcium carbonate ($CaCO_3$).

Table 2: K_b values for some bases

Acid	Formula	K_b at 25 °C
Hydroxide ion	OH^-	1,0
Carbonate ion	CO_3^{2-}	$2,1 \times 10^{-4}$
Ammonia	NH_3	$1,8 \times 10^{-5}$
Hydrogen carbonate ion	HCO_3^-	$2,4 \times 10^{-8}$
Ethanoate ion (acetate ion)	CH_3COO^-	$5,6 \times 10^{-10}$
Sulphate ion	SO_4^{2-}	$8,3 \times 10^{-13}$
Chloride ion	$C\ell^-$	Very small

2.4 Concentrated and dilute acids and bases

The concentration of acids and bases refer to the number of moles of acid or base present per volume of water. Concentration does not depend on the type of acid or base and can be changed in the laboratory.

Both strong /weak acids and bases can be concentrated when a small amount of the acid/base is added to a large volume of water.

Even when a strong acid/base is very dilute, it is still a strong acid/base and should be handled with care.

The concentration of acids and bases is given in $mol \cdot dm^{-3}$ and can calculated using the following formulae:

$$c = \frac{n}{V} \text{ OR } c = \frac{m}{MV}$$

3. Conjugate acid-base pairs

In all chemical equations written so far, a proton has been transferred to or from water. The Lowry-Brønsted theory focusses on reactants and products in an acid-base reaction and provides a new way of looking at acid-base reactions.

For example, consider the reaction of ethanoic acid with water:

$$CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

For the forward reaction CH_3COOH donates a proton and is the acid, whilst H_2O accepts a proton and is the base. When looking at the reverse reaction, H_3O^+ donates a proton and is the acid whilst CH_3COO^- accepts a proton and is the base.

Thus, when the acid CH_3COOH loses a proton, the base CH_3COO^- is formed. When the base H_2O gains a proton, the acid $H_3O^+$is formed. **Each of these acid-base pairs differ by the presence of one hydrogen ion and is a called a conjugate acid-base pair.**

$$CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

acid 1 base 2 acid 2 base 1

conjugate pair

conjugate pair

CH_3COO^- is the conjugate base of CH_3COOH and H_3O^+ is the conjugate acid of H_2O.

Every acid has a conjugate base and every base has a conjugate acid.
Note that:

- The conjugate base has one fewer H and one more minus charge than the acid.
- The conjugate acid has one more H and one fewer minus charge than the base.
- The net direction of an acid-base reaction depends on the relative strengths of the acids and bases involved. A reaction proceeds to the greater extent in the direction in which a stronger acid and a stronger base react to form a weaker acid and a weaker base. **The conjugate base of a weaker acid is a stronger base and the conjugate acid of a weaker base is a stronger acid.**

Example 1:

$$HNO_3(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + NO_3^-(aq)$$

stronger acid stronger base weaker acid weaker base

HNO_3 is a strong acid and thus completely ionised and the equilibrium position lies to far right because the forward reaction is favoured more than the reverse reaction. Therefore its conjugate base, NO_3^-, will be a weaker base than H_2O. NO_3^- has a poor tendency to accept a proton to form HNO_3.

Example 2:

$$CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

weaker acid weaker base stronger acid stronger base

CH_3COOH is a weak acid and is thus incompletely ionised. The equilibrium position lies to the left because the reverse reaction is favoured more than the forward reaction. The reverse reaction has a larger tendency to take place than the forward reaction. Therefore CH_3COO^- is a stronger base than H_2O because it readily accepts a proton from H_3O^+ to form CH_3COOH and H_2O.

Similarly, H_3O^+ is a stronger acid than CH_3COOH as it readily donates a proton to CH_3COO^- to form CH_3COOH H_2O.

4. Amphiprotic substances

Some substances can act as either an acid or a base depending on the other reactant. Such substances are called amphiprotic substances or ampholytes.

Example 1:

Reaction 1: Reaction of ammonia with water:

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

base 1 acid 2 acid 1 base 1

Reaction 2: Reaction of ethanoic acid with water

$CH_3COOH(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

acid 1 base 2 acid 2 base 1

In reaction 1, water acts as an acid. In reaction 2, water acts as base. When water reacts with an acid, it behaves like a base and accepts a proton, but when water reacts with a base, it acts like an acid and donates a proton. **Water can thus act as acid and base and is an amphiprotic substance.**

Example 2:

Reaction 1:

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

acid 1 base 2 acid 2 base 1

Reaction 2:

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

acid 1 base 2 acid 2 base 1

In reaction 1, HSO_4^- acts as a base. In reaction 2, HSO_4^- acts as an acid. HSO_4^- can **either donate or accept a proton and is an amphiprotic substance.**

Daily task 1: Homework/Classwork

Question 1: Multiple choice questions

1.1 Consider the equilibrium $H_2SO_4 + HSO_3^- \rightleftharpoons HSO_4^- + H_2SO_3$.
The two Lowry-Brønsted bases are:

A HSO_4^- and H_2SO_4
B HSO_3^- and H_2SO_3
C HSO_3^- and HSO_4^-
D HSO_4^- and H_2SO_3

1.2 Consider the following equilibrium: $HC_2O_4^- + HCO_3^- \rightleftharpoons H_2CO_3 + C_2O_4^{2-}$
Which ONE of the following correctly identifies the order of Lowry-Brønsted acids and bases in the above reaction?

A base, acid, acid, base
B acid, base, base, acid
C acid, base, acid, base
D base, acid, base, acid

1.3 Water can act as either an acid or a base. Which equation represents water reacting as an acid?

A $H_2O(\ell) + NH_3(g) \rightleftharpoons OH^-(aq) + NH_4^+(aq))$
B $H_2O(\ell) + HC\ell(aq) \rightleftharpoons H_3O^+(aq) + C\ell^-(aq)$
C $H_2O(\ell) \rightleftharpoons H_2(g) + ½O_2(g)$
D $H_2O(\ell) + C(s) \rightleftharpoons CO(g) + H_2(g)$

1.4 Which ONE of the following best describes the difference between a base and its conjugate acid?

A The base has an additional OH^- ion.
B The base has an additional H^+ ion.
C The conjugate acid has an additional OH^- ion.
D The conjugate acid has an additional H^+ ion.

1.5 The conjugate base of HBO_3^{2-} is:

A $H_2BO_3^-$ B BO_3^{2-} C BO_3^{3-} D HBO_3^-

1.6 What is the conjugate acid of $HC_2O_4^-$?

A $C_2O_4^{2-}$ B OH^- C $H_2C_2O_4^-$ D $H_2C_2O_4$

1.7 Consider the following equilibrium:

$$H_2C_2O_4(aq) + HPO_4^{2-}(aq) \rightleftharpoons HC_2O_4^-(aq) + H_2PO_4^-(aq)$$

In the above equilibrium, a conjugate pair is:

A $H_2C_2O_4$ and $H_2PO_4^-$
B HPO_4^{2-} and $HC_2O_4^-$
C HPO_4^{2-} and $H_2PO_4^-$
D $H_2C_2O_4$ and HPO_4^{2-}

1.8 Consider the equilibrium: $H_3BO_3 + CO_3^{2-} \rightleftharpoons HCO_3^- + H_2BO_3^-$

Which conjugate pair contains the strongest acid and weakest base?

	Strongest acid	**Weakest base**
A	H_3BO_3	$H_3BO_3^-$
B	CO_3^{2-}	HCO_3^-
C	HCO_3^-	$H_2BO_3^-$
D	H_3BO_3	HCO_3^-

1.9 Which ONE of the following species is not amphiprotic?

A $H_2C_6H_5O_7^-$ B H_2O C H_3BO_3 D $H_2PO_4^-$

1.10 Which of the following amphiprotic ions will act predominantly as a base in solution?

A $H_2PO_4^-$ B HSO_3^- C HSO_4^- D HPO_4^{2-}

1.11 Consider the following reaction:

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

A conjugate acid-base pair in this reaction is:

A H_3O^+ and H_2O B H_2SO_4 and H_3O^+

C HSO_4^- and H_3O^+ D HSO_4^- and H_2O

1.12 Which ONE of the following can act either as an acid or a base?

A H_3O^+(aq) B CO_3^{2-}(aq) C $C\ell^-$(aq) D HSO_4^-

Contextual questions

Question 2

2.1 For each of the following, write down the formula of the conjugate acid:

2.1.1 NH_3 2.1.2 NH_2^- 2.1.3 $C_{10}H_{14}N_2$ 2.1.4 H_2O

2.2 For each of the following, write down the formula of the conjugate base:

2.2.1 $HC\ell$ 2.2.2 H_2CO_3 2.2.3 H_2O 2.2.4 HPO_4^{2-}

Question 3

3.1 In each equation, label the acids, bases and conjugate pairs.

3.1.1 $HC\ell + H_2O \rightarrow H_3O^+ + C\ell^-$

3.1.2 $O^{2-} + H_2O \rightarrow OH^- + OH^-$

3.1.3 $NH_4^+ + BrO_3^- \rightarrow NH_3 + HBrO_3$

3.1.4 $NH_3 + H_3PO_4 \rightarrow NH_4^+ + H_2PO_4^-$

3.1.5 $HPO_4^{2-} + HSO_4^- \rightarrow H_2PO_4^- + SO_4^{2-}$

3.2 Identify amphiprotic substances in the above reactions.

Question 4

The two reactants in an acid−base reaction are HNO_2(aq) and HCO_3^-(aq).

4.1 Write the balanced equation for the above reaction.

4.2 Define the term conjugate acid−base pair.

4.3 Write down the formulae for any conjugate acid−base pair for the above reaction.

Question 5

NH_4^+ ions are mixed with HCO_3^- ions.

5.1 Write a balanced equation for the reaction that takes place.

5.2 Identify the two bases in the above reaction.

5.3 Predict whether the reaction will favour the reactants or products. Explain the answer.

Question 6

6.1 Arrange the following bases in order of increasing strength:

$C\ell^-$ OH^- CH_3COO^-

6.2 Arrange the following acids in order of increasing strength:

NH_4^+ H_3O^+ HSO_4^- (Refer to Table 1)

5. Neutralisation reactions

A neutralisation reaction takes place when an **acid reacts with a base** in aqueous solution **to form a salt and water**.

A salt is any ionic compound whose cation (positive ion) comes from a base and whose anion (negative ion) comes from an acid.

Example:

$HC\ell(aq) + NaOH(aq) \rightarrow NaC\ell(aq) + H_2O(\ell)$

acid base salt water

5.1 Reaction of a strong acid and a strong base

When a strong acid reacts with a strong base, the salt that forms will be neutral.

$HC\ell$, a strong acid, reacts with NaoH, a strong base, as follows:

$$HC\ell(aq) + NaOH(aq) \rightarrow \underset{\text{salt}}{NaC\ell(aq)} + H_2O(\ell)$$

Analysis of the salt, $NaC\ell$, shows the following:

The Na^+ ion that forms in solution is cation of a strong base, NaOH, and will have no measurable effect on the pH of the solution. The $C\ell^-$ ion is the conjugate base of $HC\ell$, a strong acid, and thus is a very weak base and will also have no effect on the pH of the solution. Therefore the salt solution will be neutral.

When **a strong acid reacts with a** (stoichiometrically equivalent amount of a) **strong base, the resulting salt solution is neutral.**

5.2 Reaction of a weak acid with a strong base

When a weak acid reacts with a strong base, the salt solution will be basic (an alkaline solution).

Ethanoic acid, a weak acid, reacts with sodium hydroxide, a strong base, as follows:

$$CH_3COOH(aq) + NaOH(aq) \rightarrow \underset{\text{salt}}{CH_3COONa(aq)} + H_2O(\ell)$$

Analysis of the salt, CH_3COONa, shows the following:

The Na^+ ion that forms in solution is cation of a strong base, NaOH, and will have no measurable effect on the pH of the solution. The CH_3COO^- ion is the conjugate base of CH_3COOH, a weak acid, and thus is a base. Therefore the salt solution will be basic.

When **a weak acid reacts with a** (stoichiometrically equivalent amount of a) **strong base, the resulting salt solution is basic (or alkaline).**

5.3 Reaction of a strong acid with a weak base

When a strong acid reacts with a weak base, the salt solution will be acidic.

Hydrochloric acid, a strong acid, reacts with ammonia, a weak base, as follows:

$HC\ell(aq) + NH_3(aq) \rightarrow NH_4C\ell(aq)$ OR $HC\ell(aq) + NH_4OH(aq) \rightarrow NH_4C\ell(aq) + H_2O(\ell)$

Note: *An ammonia solution consists of NH_4^+ ions and OH^- in an aqueous solution.*

Analysis of the salt, $NH_4C\ell$, shows the following:

The NH_4^+ ion that forms in solution is the cation of a weak base, NH_3, and thus is an acid. The $C\ell^-$ ion is the conjugate base of $HC\ell$, a strong acid, and thus is a very weak base and will also have no effect on the pH of the solution. Therefore the salt solution will be acidic.

When **a strong acid reacts with a** (stoichiometrically equivalent amount of a) **weak base, the resulting salt solution is acidic**.

Table 3: Summary of aqueous solutions of salts

Cation of salt	Anion of salt	pH of solution
From strong base (e.g. Na^+, K^+, Ca^{2+}, Mg^{2+})	From strong acid (e.g. $C\ell^-$, NO_3^-)	pH = 7 (neutral)
From strong base (e.g. Na^+, K^+, Ca^{2+}, Mg^{2+})	From weak acid (e.g. CH_3COO^-; $(COO)_2^{2-}$; CO_3^{2-})	pH > 7 (basic)
From weak base (e.g. NH_4^+)	From strong acid (e.g. $C\ell^-$, NO_3^-)	pH < 7 (acidic)

Table 4: Acid and base properties of some ions in aqueous solution

	Neutral	Basic	Acidic
Anions	$C\ell^-$, NO_3^-	CH_3COO^-; $(COO)_2^{2-}$; CO_3^{2-} ; HCO_3^-	HSO_4^-
Cations	Na^+, K^+, Ca^{2+}, Mg^{2+}	-	NH_4^+

6. Hydrolysis

Hydrolysis is **the reaction of a salt with water.**

Hydrolysis can be used to explain the pH of salt solutions. The following steps are followed:

Step 1: Split the salt into its ions and decide from which acid and which base the ions are coming. Also determine whether each of the acid and base is strong or weak. If an ion comes from either strong base or a strong acid, it will not undergo hydrolysis.

Step 2: Decide which ion, positive or negative, will undergo hydrolysis:

- If the **acid/base that is formed as product is strong** (completely ionised), the reverse reaction will be favoured and therefore there will be no forward reaction and thus **no hydrolysis** will take place.
- If the **acid/base that is formed as product is weak** (incompletely ionise), the forward reaction will be favoured and **hydrolysis will take place**. This reaction will then determine the pH of the salt solution.

Step 3: Write an equation for the reaction of ion that will undergo hydrolysis with H_2O.
The reactants will be the ion and water.
The products will be the acid from which the ion comes and OH^- OR the base from which the ion comes and H_3O^+.
If the ion accepts a proton from H_2O to form the acid it comes from, OH^- will be the other product.

Example: $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$

H^+

If the ion donates a proton to H_2O to form the base it comes from, H_3O^+ will be the other product.

Example: $NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

H^+

Step 4: The pH of the salt solution is finally determined by the hydrolysis reaction that will take place. If OH^- is formed during hydrolysis, the salt is basic. If H_3O^+ is formed during hydrolysis, the salt is acidic.

Example 1:
Will a solution of $NH_4C\ell$ be basic, acidic or neutral? Use hydrolysis to fully explain the answer.

Step 1: The two ions present in this salt are NH_4^+ and $C\ell^-$.
NH_4^+ comes from a weak base, NH_3. $C\ell^-$ comes from a strong acid, $HC\ell$.

Step 2: $C\ell^-$ will not undergo hydrolysis because it is the conjugate base of a strong acid. NH_4^+ is the conjugate acid of a weak base and will undergo hydrolysis.

Step 3: NH_4^+ will undergo hydrolysis according to the following equation:
$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$
The forward reaction will be favoured above the reverse reaction, because NH_3 is a weak base and will ionise only partially.

Step 4: The salt is acidic because H_3O^+ ions are formed during hydrolysis. The pH of the salt will be less than 7.

Example 2:
Will a solution of $NaNO_3$ be basic, acidic or neutral? Use hydrolysis to fully explain the answer.

Step 1: The two ions present in this salt are Na^+ and NO_3^-.
Na^+ comes from a strong base, NaOH. NO_3^- comes from a strong acid, HNO_3.

Step 2: NO_3^- will not undergo hydrolysis because it is the conjugate base of a strong acid. Na^+ will not undergo hydrolysis because it comes from a strong base.

Step 3: None of the two ions will undergo hydrolysis.

Step 4: No H_3O^+ or OH^- ions are formed. Therefore the salt is neutral and the pH = 7.

Example 3:
Will a solution of $NaHCO_3$ be basic, acidic or neutral? Use hydrolysis to fully explain the answer.

Step 1: The two ions present in this salt are Na^+ and HCO_3^-.
Na^+ comes from a strong base, NaOH. HCO_3^- comes from a weak acid, H_2CO_3.

Step 2: Na^+ comes from a strong base and will not undergo hydrolysis.
HCO_3^- is the conjugate base of a weak acid and will undergo hydrolysis.

Step 3: HCO_3^- will undergo hydrolysis according to the following equation:
$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$
The forward reaction will be favoured above the reverse reaction, because H_2CO_3 is a weak acid and will ionise incompletely.

Step 4: The salt is basic because OH^- ions are formed during hydrolysis. The pH of the salt will be greater than 7.

SUMMARY

- Hydrolysis of the **salt of a weak acid and a strong base** forms an alkaline solution, i.e. the **pH > 7**. Examples of such salts are sodium ethanoate, sodium oxalate and sodium carbonate.
- Hydrolysis of the **salt of a strong acid and a weak base** forms acidic solution, i.e. the **pH < 7**. An example of such a salt is ammonium chloride.
- The **salt of a strong acid and a strong bases** does not undergo hydrolysis and the solution of the salt will be neutral, i.e. **pH = 7**.

Daily task 2: Homework/Classwork

Question 1: Multiple choice questions

1.1 Which ONE of the following formulae represents a salt?

A KOH B KCℓ C CH_3OH D CH_3COOH

1.2 Consider the equation $H^+(aq) + OH^- \rightleftharpoons H_2O(\ell)$. Which type of reaction does this equation represent?

A Esterification
B Decomposition
C Hydrolysis
D Neutralisation

1.3 When HCℓ(aq) is exactly neutralised by NaOH(aq), the hydrogen ion concentration in the resulting mixture is …

A always less than the concentration of the hydroxide ions.
B always greater than the concentration of the hydroxide ions.
C always equal to the concentration of the hydroxide ions.
D sometimes greater and sometimes less than the concentration of the hydroxide ions.

1.4 Which of the following statements about weak acids is false?

A Weak acids ionise only slightly in dilute aqueous solution.
B The K_a values for weak acids are numbers that are greater than 1.
C The ionization constant for a weak acid does not include a term for the concentration of water.
D Many weak acids are familiar to us in everyday use.

1.5 Consider the ionisation constants for four weak acids, I, II, III and IV.

Weak acid	K_a
I	1×10^{-3}
II	3×10^{-5}
III	$2,6 \times 10^{-7}$
IV	4×10^{-9}

The anion of which ONE of the above acids is the weakest base?

A I B II C III D IV

1.6 Which ONE of the following salts produces neutral solutions when dissolved in water?

A $NaNO_3$ B CH_3COONa C Na_2CO_3 D $NH_4C\ell$

1.7 Consider the following salts:
I KNO_3 II $Na_2(COO)_2$ III Na_2CO_3 IV KCℓ V $(NH_4)_2SO_4$

Which response includes all the above salts that hydrolyse in dilute aqueous solution?

A I, III, and IV
B II, III, and V
C II, IV, and V
D I, II, III, and V

1.8 When solid K_2CO_3 is added to water, the pH …
A becomes less than 7 because of hydrolysis of K^+.
B becomes greater than 7 because of hydrolysis of K^+.
C becomes greater than 7 because of hydrolysis of CO_3^{2-}.
D becomes less than 7 because of hydrolysis of CO_3^{2-}.

1.9 Which ONE of the following represents the complete neutralisation of H_3PO_4 by NaOH?

A $H_3PO_4 + NaOH \rightarrow NaH + HPO_4 + H_2O$

B $H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O$

C $H_3PO_4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$

D $H_3PO_4 + 2NaOH \rightarrow Na_2HPO_4 + 2H_2O$

1.10 Which ONE of the following is the net ionic equation for the reaction of nitric acid with NaOH(aq)?

A $H^+(aq) + NO_3^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + H_2O(\ell)$

B $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

C. $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(\ell)$

D $HNO_3(aq) + NaOH(aq) + H_2O(\ell) \rightarrow NaNO_3(aq) + H_3O^+(aq) + OH^-(aq)$

1.11 Which ONE of the following is the net ionic equation for the reaction of NH_3(aq) with HNO_3(aq)?

A $NH_3(aq) + H^+(aq) + NO_3^-(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq)$

B $H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$

C $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$

D $NH_3(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq)$

1.12 Which ONE of the following is the dissociation equation for $Ca(HCO_3)_2$(s) in water?

A $HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

B $Ca(HCO_3)_2(s) \rightarrow Ca^{2+}(aq) + 2HCO_3^-(aq)$

C $Ca(HCO_3)_2(s) \rightarrow Ca^{2+}(aq) + (HCO_3)^{2-}(aq)$

D $HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$

1.13 Consider the reaction of CH_3COOH(aq) with NaOH(aq). Which ONE of the following net equations accounts for the pH of the salt?

A $CH_3COOH(aq) + NaOH(aq) \rightleftharpoons NaCH_3COO(aq) + H_2O(\ell)$

B $H^+(aq) + OH^-(aq) \rightleftharpoons H_2O(\ell)$

C $CH_3COO^-(aq) + H_2O(\ell) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$

D $CH_3COOH(aq) + OH^-(aq) \rightleftharpoons CH_3COO^-(aq) + H_2O(\ell)$

Contextual questions

Question 2

Decide whether each of the following salts will give rise to an acidic, basic or neutral solution in water. Explain how you arrived at each answer.

2.1 $NaNO_3$ 2.2 $NaHCO_3$ 2.3 K_2CO_3 2.4 $NH_4C\ell$

Question 3

The salt sodium carbonate is dissolved in water.

3.1 Define the term *hydrolysis*.

3.2 Write down the formula of the acid and the base that reacts to form this salt.

3.3 Classify each of the acid and base in QUESTION 3.2 as either weak or strong.

3.4 From your answer to QUESTION 3.3, predict whether the pH of the salt solution will be EQUAL TO 7, GREATER THAN 7 or LESS THAN 7.

3.5 Write down a balanced equation to explain the answer in QUESTION 3.4.

Question 4

The salt ammonium nitrate is dissolved in water.

4.1 Write down the formula of the acid and the base that reacts to form this salt.

4.2 Classify each of the acid and base in QUESTION 4.1 as either weak or strong.

4.3 From your answer to QUESTION 4.2, predict whether the pH of the salt solution will be EQUAL TO 7, GREATER THAN 7 or LESS THAN 7.

4.4 Write down a balanced equation to explain the answer in QUESTION 4.3.

7. Acid-base titrations

An **acid-base titration** is a procedure for determining the amount of acid (or base) in a solution by determining the volume of the base (or acid) of known concentration that will completely react with it.

The solution of known concentration is called a **standard solution**. The process of determining the unknown concentration by measuring volumes is also referred to a **volumetric analysis**.

Apparatus needed for an acid-base titration

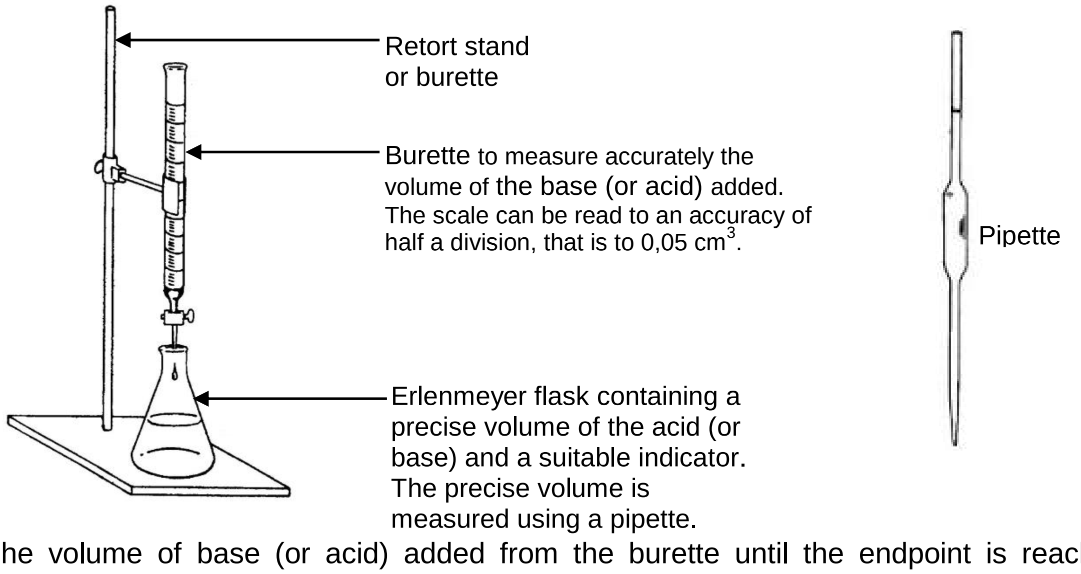
The apparatus needed to do an acid-base titration in the laboratory is shown in the diagram below. In a titration, the pipette is used to transfer a precise volume (usually to ±0,05 cm^3) of an acid (or base) into an Erlenmeyer flask/conical flask.

A base (or acid) that reacts with the pipetted acid (or base) in the Erlenmeyer flask/conical flask is carefully added from a burette until it has all exactly reacted i.e. the moles of the base (or acid) equal the moles of the acid (or base). This is called the **equivalence point** of the reaction and the **end point** of the titration. There needs to be a way of knowing when the end point is reached. An **indicator** that changes colour within the range of the equivalence point of the reaction or endpoint of the titration is added to the acid (or base) in the Erlenmeyer flask/conical flask.

Table 5: Acid-base indicators

Indicator	pH range	Colour in acid	Colour in base
Bromothymol blue	6,0 – 7,6	yellow	blue
Phenolphthalein	8,3 – 10,0	colourless	dark pink
Methyl orange	2,5 – 4,4	red	yellow

The endpoint and equivalence point are not necessarily equal, but they do represent the same idea. An endpoint is indicated by an indicator at the end of a titration. An equivalence point is when the moles of a standard solution equal the moles of the solution of unknown concentration.



The volume of base (or acid) added from the burette until the endpoint is reached, is measured and then used to calculate the unknown concentration.

Steps when calculating the unknown concentration:

Step 1: Calculate the number of moles of the solution of known concentration:
$n = cV$

Step 2: Use the mole ratio in the balanced equation to determine the number of moles of the solution of unknown concentration that have reacted.

Step 3: Calculate the unknown concentration: $c = \frac{n}{V}$

Precautionary measures when doing a titration

- Thoroughly clean all glassware (burette, pipette, Erlenmeyer flask) before the titration.
- Rinse the pipette first with distilled water and then twice with the acid (or base) to be measured. Rotate the pipette to ensure that the solution contacts the entire inner surface of the pipette.
- When adding the solution from the pipette into the Erlenmeyer flask, hold the tip of the pipette against the inner surface of the flask to avoid splatter. When the flow of liquid from the pipette stops, continue to hold the pipette in a vertical position for 15 seconds to allow draining of the pipette. Do not blow through the pipette to expel the last drop.
- Before filling the burette, it should be rinsed at least twice with the base (or acid) with which it will be filled. Drain the rinse solution through the tip of the burette.
- Use a funnel to fill the burette to above the top calibration mark on the burette. After eliminating air bubbles in the tip of the burette, open the tap and lower the meniscus of the solution until it is at a point on the calibrated portion of the burette.
- When measuring the volume of solution in the burette, make sure that your eye is in level with the meniscus to prevent a parallax mistake.
- When adding the solution from the burette into the solution in the flask, swirl the flask with the right hand whilst the tap of the burette is controlled with the left hand. This will ensure thorough mixing of the two solutions and prevent over-titration.
- The approach of the endpoint will be signalled by a temporary colour change of the indicator. Smaller amounts of solution from the burette should then be added and later drop by drop to prevent over-titration.
- Wash any titrant (solution in burette) spilled against the sides of the Erlenmeyer flask down with distilled water before the endpoint is reached. This water will not change the number of moles of acid (or base) present in the flask and thus will not influence the results.
- The most accurate endpoint is obtained when the intensity of the new colour of the indicator is the faintest that can be seen and remains for at least 30 seconds.
- For reliable results, the titration should be repeated at least three times.

Safety precautions when doing acid-base titrations

- Always add acid to water and not water to acid.
- Wear protective equipment: gloves, goggles, lab gown, etc.
- Clean all spills.
- When chemicals are spilled on skin, wash for at least 15 minutes with running water.
- Always have an emergency eye wash nearby.

7.1 Titration of a strong acid and a strong base

- When a strong acid reacts with a stoichiometrically equivalent amount of a strong base, the resulting salt solution is neutral (see section 5 and 6 on p 13 & 14).
- The endpoint of the titration of a strong acid (e.g. $HC\ell$) with a strong base (e.g. $NaOH$) is at pH = 7.
- The best choice of indicator will be **bromothymol blue** because the pH at the endpoint of the titration falls within the range in which the indicator will change colour i.e. pH 6,0 - 7,6.

Example 1:

A standard solution is prepared in a 200 cm^3 volumetric flask by dissolving 4,9 g of pure sulphuric acid in water and filling the flask to the mark. During a titration, 20,7 cm^3 of this standard solution completely neutralise 10,0 cm^3 of a sodium hydroxide solution of unknown concentration.

1.1 Write the balanced equation for the titration reaction.

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

1.2 Calculate the concentration of the sulphuric acid solution.

First calculate the molar mass of H_2SO_4 and convert the volume of acid i.e. 200 cm^3 to dm^3.

$$c = \frac{m}{MV} = \frac{4,9}{(98)(0,2)} = 0,25\ mol \cdot dm^{-3}$$

1.3 **Calculate the number of moles of sulphuric acid (n_a) neutralised.**

Convert the volume of acid neutralised i.e. 20,7 cm^3 to dm^3.

$n_a = c_aV_a = (0,25)(20,7 \times 10^{-3}) = 5,175 \times 10^{-3}$ mol

1.4 Calculate the number of moles of sodium hydroxide (n_b) neutralised.

The balanced equation shows that 1 mole H_2SO_4 requires 2 moles NaOH to completely react. This is the required stoichiometric factor to obtain the amount of NaOH that has reacted.

At the equivalence point: $\frac{n_a}{n_b} = \frac{1}{2}$ $\therefore$ $n_b = 2n_a = 2(5,175 \times 10^{-2}) = 0,01035$ mol

1.5 Calculate the concentration of the sodium hydroxide.

Convert the volume of base neutralised i.e. 10,0 cm^3 to dm^3.

$$c_b = \frac{n_b}{V_b} = \frac{0,01}{10 \times 10^{-3}} = 1,035\ mol \cdot dm^{-3}$$

7.2 Titration of a weak acid and a strong base

- When a weak acid reacts with a stoichiometrically equivalent amount of a strong base, the resulting salt solution is basic (see section 5 and 6 on p 13 & 14).
- The endpoint of the titration of a weak acid (e.g. CH_3COOH) with a strong acid (e.g. NaOH) is at pH > 7.
- The best choice of indicator will be **phenolphthalein** because the pH at the endpoint of the titration falls within the range in which the indicator will change colour i.e. pH 8,3 – 10,0.

Example 2:

A 1,034 g sample of impure oxalic acid is dissolved in water and an acid-base indicator is added. The sample requires 34,47 cm^3 of 0,485 $mol \cdot dm^{-3}$ NaOH to reach the equivalence point. Calculate the mass of oxalic acid and hence the percentage purity of the sample.

Step 1: Write the balanced equation for the reaction.

$(COOH)_2(aq) + 2NaOH(aq) \rightarrow (COO)_2Na_2(aq) + 2H_2O(\ell)$

Step 2: Calculate the number of moles of sodium hydroxide (n_b) neutralised.

Convert the volume of base neutralised i.e. 34,47 cm^3 to dm^3.

$n_b = c_bV_b = (0,485)(34,47 \times 10^{-3}) = 1,67 \times 10^{-2}$ mol

Step 3: Calculate the number of moles of oxalic acid (n_a) neutralised.

The balanced equation shows that 1 mole $(COOH)_2$ *requires 2 moles NaOH to completely react. This is the required stoichiometric factor to obtain the amount of* $(COOH)_2$ *that has reacted.*

At the equivalence point: $\frac{n_a}{n_b} = \frac{1}{2}$ $\therefore$ $n_a = ½n_b = ½(1,67 \times 10^{-2}) = 8,36 \times 10^{-3}$ mol

Step 4: Calculate the mass of oxalic acid that has reacted.

$n = \frac{m}{M} \therefore 8{,}36 \times 10^{-3} = \frac{m}{90} \therefore m = 0{,}75\ g$

Step 5: Calculate the percentage of oxalic acid in the sample.

$$\text{Percentage purity} = \frac{\text{mass reacted}}{\text{mass of sample}} \times 100$$
$$= \frac{0{,}75}{1{,}034} \times 100$$
$$= 72{,}76\%$$

7.3 Titration of a strong acid with a weak base

- When a strong acid reacts with a stoichiometrically equivalent amount of a weak base, the resulting salt solution is acidic (see section 5 and 6 on p 13 & 14).
- The endpoint of the titration of a strong acid (e.g. $HC\ell$) with a weak base (e.g. Na_2CO_3) is at pH < 7.
- The best choice of indicator will be **methyl orange** because the pH at the endpoint of the titration falls within the range in which the indicator will change colour i.e. pH 2,5 – 4,4.

Example 3:

A bulk solution of hydrochloric acid is standardised using pure anhydrous sodium carbonate (Na_2CO_3, a primary standard). 13,25 g of sodium carbonate is dissolved in about 150,0 cm^3 of distilled water in a beaker. The solution is then transferred, with appropriate washings, into a volumetric flask, and the volume of water made up to 250 cm^3, and thoroughly shaken (with stopper on!) to ensure complete mixing.

25,0 cm^3 of the sodium carbonate solution is pipetted into a conical flask and methyl orange indicator is added. During a titration it is found that 24,65 cm^3 of a hydrochloric acid solution, of unknown concentration, is needed to completely neutralise it.

3.1 Calculate the concentration of the prepared sodium carbonate solution.
First calculate the molar mass of Na_2CO_3 and convert the volume of acid i.e. 250 cm^3 to dm^3.

$c = \frac{m}{MV} = \frac{13{,}25}{(106)(0{,}25)} = 0{,}5\ mol \cdot dm^{-3}$

3.2 Write the balanced equation for the titration reaction.
$Na_2CO_3(aq) + 2HC\ell(aq) \rightarrow 2NaC\ell(aq) + H_2O(\ell) + CO_2(g)$

3.3 Calculate the number of moles of sodium carbonate titrated.
Convert the volume of base neutralised i.e. 20,7 cm^3 to dm^3.

$n_b = c_bV_b = (0{,}5)(25{,}0 \times 10^{-3}) = 0{,}0125\ mol$

3.4 Calculate the number of moles of hydrochloric acid used in the titration.
The balanced equation shows that 2 mole $HC\ell$ requires 1 moles Na_2CO_3 *to completely react. This is the required stoichiometric factor to obtain the amount of $HC\ell$ that has reacted.*

At the equivalence point: $\frac{n_a}{n_b} = \frac{2}{1} \therefore n_a = 2n_b = 2(0{,}0125) = 0{,}025\ mol$

3.5 Calculate the concentration of the hydrochloric acid.
Convert the volume of acid neutralised i.e. 24,56 cm^3 to dm^3.

$c_b = \frac{n_b}{V_b} = \frac{0{,}025}{24{,}56 \times 10^{-3}} = 1{,}02\ mol \cdot dm^{-3}$

Daily task 3: Homework/Classwork

Question 1: Multiple choice questions

1.1 Which ONE of the following is a characteristic of basic solutions?

A They turn litmus a blue colour.
B They accept OH^- ions.
C They have a pOH greater than 7.
D They react with Mg to produce H_2 gas.

1.2 One of the species in the chemical indicator HIn^- exhibits a yellow colour. If acid is added, the indicator turns red. Which ONE of the following is correct?

	Red	Yellow
A	H_2In	HIn^-
B	In^{2-}	H_2In
C	In^{2-}	HIn^-
D	HIn^-	H_2In

1.3 The equilibrium between the acid form (HIn) and the base form (In^-) of an acid-base indicator is represented by the equation below.

$$HIn + H_2O \rightleftharpoons H_3O^+ + In^-$$

Consider the following statements about the transition point of all indicators described by the above equilibrium:

I The moles of H_3O^+ equals the moles of OH^-.
II $[HIn] = [In^-]$
III $[H_3O^+] = 1{,}0 \times 10^{-7}\ mol{\cdot}dm^{-3}$

Which of the above the above statements is/are true?

A I and II only
B II only
C II and III only
D I, II and II

1.4 Consider the following indicator equilibrium:

$$\underset{\text{colourless}}{HIn} + H_2O \rightleftharpoons H_3O^+ + \underset{\text{blue}}{In^-}$$

What is the effect of adding HCℓ to a blue sample of this indicator?

	Shift of equilibrium position	Colour change
A	To the right	More blue
B	To the left	Less blue
C	To the left	More blue
D	To the right	Less blue

1.5 Which of the following titrations would have a pH > 7 at the equivalence point?

A HCOOH with NaOH
B HNO_3 with KOH
C HCℓ with NH_3
D H_2SO_4 with $NaHCO_3$

1.6 Which ONE of the following 0,1 mol·dm^{-3} solutions will have the lowest hydrogen ion concentration?

A $HC\ell(aq)$
B $KOH(aq)$
C $H_2SO_4(aq)$
D $CH_3COOH(aq)$

Contextual questions

Question 2
A 25,0 mℓ sample of the weak acid H_2S is titrated with 31,8 mℓ of 0,30 mol·dm^{-3} NaOH. Calculate the concentration of the acid.

Question 3
Calculate the volume of 0,300 mol·dm^{-3} HNO_3 needed to completely neutralise 25,0 mℓ of 0,250 mol·dm^{-3} $Sr(OH)_2$.

Question 4
A learner uses a standard solution of sodium hydrogen carbonate to determine the concentration of a sulphuric acid solution.

4.1 Write down the definition of a standard solution.
4.2 Why is H_2SO_4 regarded as a strong acid?
4.3 Write the balanced equation for the reaction of sulphuric acid with water.

In a titration, the learner finds that 20 cm^3 of a 0,2 mol·dm^{-3} solution of sodium hydrogen carbonate neutralises 12 cm^3 of the sulphuric acid solution. The balanced equation for this reaction is:

$$2NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2CO_2 + 2H_2O$$

4.4 How many moles of $NaHCO_3$ are present in 20 cm^3 of the 0,2 mol.dm^{-3} $NaHCO_3$ solution?
4.5 Determine the number of moles of H_2SO_4 that are neutralised by 20 cm^3 of the 0,2 mol·dm^{-3} $NaHCO_3$ solution.
4.6 Calculate the concentration of the H_2SO_4 solution.
4.7 From the table below, select the most suitable indicator for use in this titration.

Indicator	pH range
Methyl orange	2,5 – 4,4
Bromothymol blue	6,0 – 7,6
Phenolphthalein	8,3 – 10,0

Question 5
5.1 Define a strong acid.
5.2 Write an equation to show the hydrolysis of sodium carbonate
5.3 A monoprotic acid has a concentration of 0,01 mol·dm^{-3} and a pH of 3.
5.3.1 Calculate the concentration of the hydrogen ions in this solution.
5.3.2 Is this acid comparable (similar) in strength to HCℓ? (Answer YES or NO)
5.3.3 Give a reason for the answer to QUESTION 5.3.2
5.4 A learner adds a sample of calcium carbonate to 50,0 cm^3 of hydrochloric acid of concentration 1,0 mol·dm^{-3}. The hydrochloric acid is in excess. The balanced equation for the reaction that takes place is:

$$CaCO_3 + 2HC\ell \rightarrow CaC\ell_2 + CO_2 + H_2O$$

The excess HCℓ is now neutralised by 28,0 cm^3 of a 0,5 mol·dm^{-3} sodium hydroxide solution. The balanced equation for this reaction is: $HC\ell + NaOH \rightarrow NaC\ell + H_2O$
Calculate the mass of calcium carbonate in this sample.

8. Auto-ionisation of water

Although pure water is often considered a nonelectrolyte i.e. a non-conductor of electricity, precise measurements do show a very small conduction. This conduction results from auto-ionisation or self-ionisation i.e. a reaction in which two like molecules react to give ions.

In the case of water, a proton from one H_2O molecule is transferred to another H_2O molecule, leaving behind an OH^- ion and forming a hydronium ion, $H_3O^+(aq)$.

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

The equilibrium constant, referred to as the ion product of water (K_w), for this reaction can be written as: $K_w = [H_3O^+][OH^-] = 1{,}0 \times 10^{-14}$ at 25 °C

The small value of K_w shows the small extent to which water reacts with itself.

Using K_w, you can determine the concentrations of H_3O^+ and OH^- ions in pure water. These ions are produced in equal numbers in pure water, so their concentrations are equal.

Let x = $[H_3O^+] = [OH^-]$

Substitution into the equation for the ion product constant:

$K_w = [H_3O^+][OH^-] = x^2 = 1{,}0 \times 10^{-14}$

$\therefore x = 1{,}0 \times 10^{-7}$ mol·dm^{-3}

Acidic solutions will have hydronium ion concentrations greater than $1{,}0 \times 10^{-7}$.
Basic solution will have hydronium ion concentrations smaller than $1{,}0 \times 10^{-7}$.

9. pH calculations

9.1 Definition of pH

Whether an aqueous solution is acidic, basic or neutral depend on the hydronium ion concentration. These concentration values are very small and therefore it is more convenient to give acidity in terms of pH.

- pH is defined as the negative of the logarithm of the hydronium ion concentration in mol·dm^{-3}:

  In symbols: $pH = -\log[H_3O^+]$

  Note: *You can also write* $pH = -\log[H^+]$

- Similarly we can define the pOH of a solution as the negative of the logarithm of the hydroxide ion concentration in mol·dm^{-3}.

  In symbols: $pOH = -\log[OH^-]$

- If we take the negative logarithms of both sides of the expression for the ion product of water, we can obtain another useful equation: pH + pOH = 14

  Derivation: $K_w = 1{,}0 \times 10^{-14} = [H_3O^+][OH^-]$

  $-\log K_w = -\log(1{,}0 \times 10^{-14}) = -\log([H_3O^+][OH^-])$

  $pK_w = 14 = -\log[H_3O^+] + -\log[OH^-]$

  $pK_w = 14 = pH + pOH \therefore pH + pOH = 14$

9.2. pH of neutral, acidic and basic solutions

pH of a neutral solution:
$pH = -\log[H_3O^+] = -\log(1{,}0 \times 10^{-7}) = 7$.

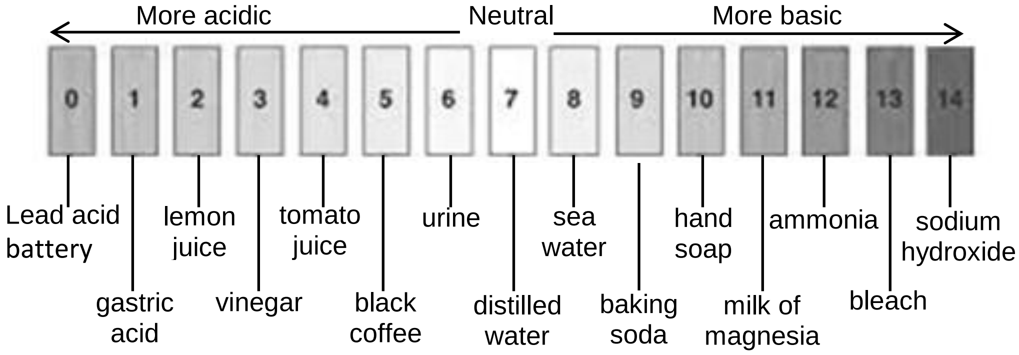
pH of an acidic solution i.e. $[H_3O^+] > 1{,}0 \times 10^{-7}$:
$pH = -\log[H_3O^+] = -\log(1{,}0 \times 10^{-3}) = 3$.

pH of a neutral solution i.e. $[H_3O^+] < 1{,}0 \times 10^{-7}$:
$pH = -\log[H_3O^+] = -\log(1{,}0 \times 10^{-10}) = 10$.

From the above it can be seen that the pH values of acidic solutions are smaller than 7 and that of basic solutions are greater than 7.

9.3 pH scale

The pH scale is a scale from 0 to 14 used as a measure of the acidity and basicity of solutions.



9.4 Calculating pH

9.4.1 pH of strong acids

Strong acids ionise completely and therefore the hydronium ion concentration will be equal to the concentration of the acid.

Example 1:

A hydrochloric acid solution has a concentration of 4×10^{-3} mol·dm^{-3} at 25 °C.
Calculate the:

1.1 H_3O^+ concentration in the solution

HCℓ is a strong acid and ionise completely: $HC\ell(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + C\ell^-(aq)$

$[H_3O^+] = [HC\ell] = 4 \times 10^{-3}$ mol·dm^{-3}

1.2 OH^- concentration in the solution

$K_w = [H_3O^+][OH^-] = 1,0 \times 10^{-14}$

$\therefore (4 \times 10^{-3})[OH^-] = 1,0 \times 10^{-14}$

$\therefore [OH^-] = 2,5 \times 10^{-12}$ mol·dm^{-3}

1.3 pH of the solution

$pH = -\log[H_3O^+] = -\log(4 \times 10^{-3}) = 2,39$

Example 2:

A solution has a pH of 3,75, Calculate the hydronium ion concentration in the solution.

$pH = -\log[H_3O^+] = 3,75$

$\therefore [H_3O^+] = 10^{-pH} = 10^{-3,75} = 1,78 \times 10^{-4}$ mol·dm^{-3}

9.4.2 pH of strong bases

Strong acids dissociate/ionise completely and therefore the hydroxide ion concentration will be equal to the concentration of the base.

Example 3:

Calculate the pH of 0,10 mol·dm^{-3} NaOH(aq) at 25 °C.

NaOH is a strong base and dissociates completely: $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

Method 1:

$[OH^-] = [NaOH] = 0,1$ mol·dm^{-3}

$K_w = [H_3O^+][OH^-] = 1,0 \times 10^{-14}$

$\therefore [H_3O^+][0,1] = 1,0 \times 10^{-14}$

$\therefore [H_3O^+] = 1 \times 10^{-13}$ mol·dm^{-3}

$pH = -\log[H_3O^+] = -\log(1 \times 10^{-13}) = 13$

Method 2:

$[OH^-] = [NaOH] = 0,1$ mol·dm^{-3}

$pOH = -\log[OH^-] = -\log(0,1) = 1$

$pH + pOH = 14$

$\therefore pH + 1 = 14$

$\therefore pH = 13$

Daily task 4: Homework/Classwork

Question 1: Multiple choice questions

1.1 The ionisation of water can be represented by:

A $2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ B $H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^{-(}aq)$

C $H_2O(\ell) \rightleftharpoons 2H^+(aq) + O^{2-}(aq)$ D $2H_2O(\ell) \rightleftharpoons 2H_2(g) + O_2(g)$

1.2 Consider the following equilibrium at 25 °C: $2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
What happens to $[OH^-]$ and pH as 0,1 mol·dm^{-3} HCℓ is added?

A $[OH^-]$ increases and pH decreases. B $[OH^-]$ decreases and pH increases.

C $[OH^-]$ decreases and pH decreases. D $[OH^-]$ increases and pH increases.

1.3 The ionization of water is endothermic. Which ONE of the following could be correct if the temperature of water is decreased?

	K_w	pH	Classification
A	Remains the same	7,0	Neutral
B	Increases	7,1	Basic
C	Decreases	6,8	Acidic
D	Increases	7,1	Neutral

1.4 A drop in pH level of 2 in an aquarium would mean that the acidity, as measured by $[H^+]$, had changed by a factor of:

A 2 B 10 C 100 D 100

1.5 Which of the following solutions has the greatest hydroxide ion concentration?

A A buffer solution with pH = 5 C 0,1 mol·dm^{-3} HCℓ

B 0,1 mol·dm^{-3} CH_3COOH D Pure water

Contextual questions

Question 2

A sample of orange juice has a hydronium ion concentration of 2,9 x 10^{-4} mol·dm^{-3}.

2.1 Calculate the pH of the juice.

2.2 Is the solution acidic or basic?

Question 3

Calculate the hydronium ion concentration in:

3.1 Arterial human blood with a pH of 7,4

3.2 A brand of carbonated beverage with a pH of 3,16

Question 4

Calculate the:

4.1 pH of a 2,5 mol·dm^{-3} KOH solution

4.2 $[H_3O^+]$ of a KOH solution that has a pH of 13,48

4.3 pH of a 1,5 mol·dm^{-3} H_2SO_4 solution

4.4 Hydroxide ion concentration in 0,01 mol·dm^{-3} solution of ammonia with a pH of 10,6

Question 5

A learner adds 0,06 mole of NaOH to 1 dm^3 of a 0,5 mol·dm^{-3} HCℓ solution.

5.1 Write down a balanced equation for the reaction that takes place.

5.2 Calculate the initial number of moles of HCℓ present in the solution.

5.3 Write down the number of moles of NaOH needed to react with the acid.

5.4 Which one of the two substances is in excess?

5.5 Calculate the pH of the final solution.

10. Comparison of strong and weak acids

10.1 Ionisation and K_a values

Strong acids are completely ionised. This implies that the hydronium ion concentration in a strong acid is equal to the original concentration of the acid.
For example, if the concentration of HCℓ(aq) is 0,1 mol·dm^{-3}, the concentration of H_3O^+(aq) will also be 0,1 mol·dm^{-3}

$$\underset{0{,}1\ \text{mol·dm}^{-3}}{HC\ell(aq)} + H_2O(\ell) \rightarrow \underset{0{,}1\ \text{mol·dm}^{-3}}{H_3O^+(aq)} + C\ell^-(aq)$$

The equilibrium position in the above reaction lies to the far right and the ionisation constant K_a will be larger than 1. A large value for K_a indicates ionisation products are strongly favoured.

Weak acids are only partially ionised. This implies that the hydronium ion concentration of a weak acid is much lower than the original acid concentration. The hydronium ion concentration is smaller than if the acid was a strong acid of the same concentration.
For example, if the concentration of ethanoic acid is 0,1 mol·dm^{-3}, the concentration of H_3O^+(aq) will be LESS THAN 0,1 mol·dm^{-3}

$$\underset{0{,}1\ \text{mol·dm}^{-3}}{CH_3COOH(aq)} + H_2O(\ell) \rightarrow \underset{<<\ 0{,}1\ \text{mol·dm}^{-3}}{H_3O^+(aq)} + C\ell^-(aq)$$

The equilibrium position in the above reaction lies to the left and the ionisation constant K_a will be less than 1. A small value for K_a indicates reactants are favoured.

10.2 pH

pH depends on the hydronium ion concentration. For the same acid concentration, the $[H_3O^+]$ of a **strong acid** will be higher than that of a weak acid. pH = -log$[H_3O^+]$, therefore, the pH of the strong acid will be lower than that of the weak acid.

10.3 Conductivity

Strong acids ionise completely to form a high concentration of ions. Therefore they are strong electrolytes and conduct a current well.
Weak acids ionise only partially the ion concentration will be low as most molecules remain intact. Therefore they are weak electrolytes and conduct only a small current.

10.4 Reaction rate

The rate of a reaction depends on the concentration of reactants. For example, when zinc reacts with hydrochloric acid and ethanoic acid respectively, hydrogen gas will be released in each case. The ion concentration of the strong acid, HCℓ, is higher than that of the weak acid, CH_3COOH. Therefore the release of hydrogen gas will be faster when HCℓ is used.

Table 4: Comparison of strong and weak acids

	Strong acid	**Weak acids**
Ionisation	Completely ionised	Partially ionised
K_a value	$K_a > 1$	$K_a < 1$
$[H_3O^+]$ for acids of same concentration	Equal to the original acid concentration	Much less than the original acid concentration
pH of acids of same concentration	Lower pH	Higher pH
Conductivity	Strong electrolyte	Weaker electrolyte
Reaction rate	Faster rate	Slower rate

Daily task 5: Homework/Classwork

Question 1: Multiple choice questions

1.1 Which solution would have the greatest electrical conductivity?

A $0{,}1\ mol\cdot dm^{-3}\ NH_3$
B $0{,}1\ mol\cdot dm^{-3}\ HC\ell$
C $0{,}1\ mol\cdot dm^{-3}\ CH_3OH$
D $0{,}1\ mol\cdot dm^{-3}\ H_2SO_4$

1.2 Which of the following would be the same when comparing equal volumes of $1{,}0\ mol\cdot dm^{-3}$ $HC\ell(aq)$ and $1{,}0\ mol\cdot dm^{-3}$ $CH_3COOH(aq)$?

A The moles of base required for neutralization.
B The titration curve for reaction with a base.
C The electrical conductivity.
D The pH.

1.3 When comparing equal volumes of $0{,}10\ mol\cdot dm^{-3}$ $HNO_3(aq)$ with $0{,}10\ mol\cdot dm^{-3}$ $HC\ell(aq)$, what would be observed?

A The volumes of $0{,}10\ mol\cdot dm^{-3}$ NaOH needed for neutralisation would be different.
B The pH values would be the same.
C The electrical conductivities would be different.
D The effects on blue litmus paper would be different.

1.4 Consider the following statements about acids.

I Acids increase the concentration of hydrogen ions in solution.
II Acids increase the concentration of hydroxide ions in solution.
III Acids increase the pH of a solution.
IV Acids react with magnesium to produce hydrogen gas.

Which of the above statements are true?

A I and IV
B II and IV
C I, II, and IV
D II, III, and IV

1.5 Which of the following processes can cause the pH of a basic solution to decrease?

I. Addition of water
II. Addition of acid
III. Addition of the same basic solution

A I only
B II only
C I and II
D I, II and III

Daily task 6: Homework/Class work

Question 1: Multiple choice questions

1.1 If equal volumes of 0,10 mol·dm^{-3} $HC\ell(aq)$ solution and 0,10 mol·dm^{-3} $CH_3COOH(aq)$ solution are compared, which would be true of the $CH_3COOH(aq)$?

A It would have a higher hydronium ion concentration.
B It would have a higher pH.
C It would produce a larger volume of hydrogen gas when reacted with zinc.
D It would require a greater volume of 0,10 mol·dm^{-3} $NaOH(aq)$ solution for neutralisation.

1.2 Water can act as either an acid or a base. Which ONE of the following equations represents water reacting as an acid?

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

B $H_2O(\ell) + HC\ell(aq) \rightleftharpoons H_3O^+(aq) + C\ell^-(aq)$

C $H_2O(\ell) \rightleftharpoons H_2(g) + ½O_2(g)$

D $H_2O(\ell) + C(s) \rightleftharpoons CO(g) + H_2(g)$

1.3 Which ONE of the following is a property of acids?

A Acidic solutions feel slippery.
B Acids taste bitter.
C Acids react with certain metals to generate hydrogen.
D Acids turn red litmus paper blue.

1.4 Which ONE of the following salts will dissolve in water to give a basic solution?

A $(NH_4)_2CO_3$
B $NH_4C\ell$
C Na_2SO_4
D Na_2CO_3

1.5 According to Arrhenius, what does the reaction $Ba(OH)_2(s) \rightarrow ¼Ba^{2+}(aq) + 2OH^-(aq)$ represent?

A Dissociation of an acid
B Dissociation of a base
C Formation of an acidic solution
D Formation of a neutral solution

1.6 According to the Brønsted-Lowry theory, a base is a/an …

A hydrogen ion (proton) acceptor.
B electrolyte.
C nonelectrolyte.
D substance that increases the hydrogen (hydronium) ion concentration

1.7 Which equation shows an acid-base neutralisation reaction?

A $Zn(s) + 2HC\ell(aq) \rightarrow ¼H_2(g) + ZnC\ell_2(aq)$
B $H_2CO_3(aq) \rightarrow ¼CO_2(aq) + H_2O(\ell)$
C $2NaOH(aq) + CaC\ell_2(aq) \rightarrow ¼2NaC\ell(aq) + Ca(OH)_2(s)$
D $NaOH(aq) + HC\ell\ (aq) \rightarrow ¼\ NaC\ell\ (aq) + H_2O(\ell)$

1.8 Consider the following reaction: $H_2O(\ell) + HPO_4^{2-}(aq) \rightleftharpoons H_2PO_4^-(aq) + OH^-(aq)$

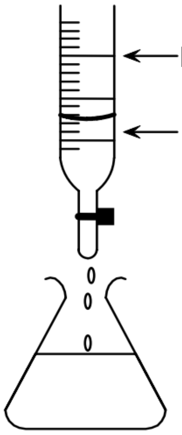
Which ONE of the following pairs is the Brønsted-Lowry acids in this reaction?

A $HPO_4^{2-}(aq)$ and $OH^-(aq)$
B $H_2O(\ell)$ and $HPO_4^{2-}(aq)$
C $H_2O(\ell)$ and $H_2PO_4^-(aq)$
D $H_2O(\ell)$ and $OH^-(aq)$

1.9 A sodium hydroxide solution of concentration 0,1 mol·dm^{-3} is added dropwise to an ethanoic acid solution of concentration 0,1 mol·dm^{-3}. Which ONE of the following substances will increase in concentration as sodium hydroxide is added dropwise?

A H_3O^+ B OH^- C CH_3COO^- D H_2O

1.10 In a titration involving HCℓ and NaOH, as shown in the sketch below, a learner accidentally exceeded the endpoint.



Which ONE of the following is correct for the solution that is now in the flask?

A $[H^+] < [OH^-]$ and pH < 7 B $[H^+] > [OH^-]$ and pH < 7
C $[H^+] < [OH^-]$ and pH > 7 D $[H^+] > [OH^-]$ and pH > 7

1.11 0,1 mol·dm^{-3} solutions of four substances below are prepared. Which ONE will have the lowest pH?

A NH_4NO_3
B NH_4OH
C KNO_3
D KOH

1.12 Which one of the following solutions will have a pH less than 7?
A KCℓ(aq)
B $NaNO_3$(aq)
C $NH_4C\ell$(aq)
D CH_3COONa(aq)

Contextual questions

Question 2
A sample of sodium hydrogen carbonate was tested for purity using the following method:
0,4 g of the solid was dissolved in 100 cm^3 of water and titrated with 0,2 mol·dm^{-3} hydrochloric acid using methyl orange indicator. 23,75 cm^3 of acid was required for complete neutralisation.

2.1 Write the balanced equation for the titration reaction.
2.2 Calculate the number of moles of acid used in the titration and the number of moles of sodium hydrogen carbonate titrated.
2.3 Calculate the mass of sodium hydrogen carbonate titrated and hence the purity of the sample.

Question 3
Calculate the volume of a 0,53 mol·dm^{-3} NaOH solution needed to neutralise 37 mℓ of a 0,33 mol·dm^{-3} H_2SO_4 solution.

Question 4

A learner determined the pH of a number of solutions at 25 °C. She obtained the following results:

Solution	Battery acid	Orange juice	Bicarbonate of soda
pH	1	4,2	12

4.1 Which solution contains the highest concentration of hydrogen ions?
4.2 Calculate the concentration of hydroxide ions in orange juice.
4.3 How will the pH of battery acid change when:
(Only write INCREASES, DECREASES or STAYS THE SAME.)
4.3.1 Distilled water is added to it
4.3.2 Some of the bicarbonate of soda solution is added to it

Question 5

A learner adds a sample of calcium carbonate to 50,0 cm^3 of hydrochloric acid of concentration 1,0 $mol \cdot dm^{-3}$. The hydrochloric acid is in excess.

The balanced equation for the reaction that takes place is:

$$CaCO_3 + 2HC\ell \rightarrow CaC\ell_2 + CO_2 + H_2O$$

The excess $HC\ell$ is now neutralised by 28,0 cm^3 of a 0,5 $mol \cdot dm^{-3}$ sodium hydroxide solution. The balanced equation for this reaction is: $HC\ell + NaOH \rightarrow NaC\ell + H_2O$

5.1 Hydrochloric acid is a strong acid. Explain the term strong acid.
5.2 Calculate the mass of calcium carbonate in this sample.

Question 6

6.1 Write down:
6.1.1 The meaning of the term diprotic acid
6.1.2 The formula of a diprotic acid

6.2 Magnesium hydroxide ($Mg(OH)_2$) is often used as medicine to relieve an upset stomach. The pH of the $HC\ell(aq)$ in a person's stomach is 1.
6.2.1 Calculate the concentration of the hydrochloric acid in the person's stomach.
6.2.2 Will the pH in the stomach INCREASE, DECREASE or STAY THE SAME after taking in a dose of $Mg(OH)_2$?

6.3 Sodium carbonate crystals ($Na_2CO_3.10H_2O$) are used to neutralise a hydrochloric acid solution with a pH of 1.
6.3.1 Write down the balanced equation for the neutralisation reaction.
6.3.2 Calculate the mass of sodium carbonate crystals that will be required to neutralise 200 cm^3 of the hydrochloric acid solution.
6.3.3 Choose from the following table the most suitable indicator for the reaction.

Indicator	**pH range in which the colour changes**
Methyl red	4,8 – 6,0
Neutral red	6,8 – 8,0
Chlorophenol red	7,0 – 8,8

6.3.4 Give a reason for the answer to QUESTION 6.3.3.

Question 7

The acid-base equilibrium for phenolphthalein ($C_{20}H_{14}O_4(aq)$) in solution is represented by the equation below.

$$\underset{\text{Colourless}}{C_{20}H_{14}O_4(aq)} + 2H_2O(\ell) \rightleftharpoons 2H_3O^+(aq) + \underset{\text{pink}}{C_{20}H_{12}O_4^{2-}(aq)}$$

Explain, using Le Chatelier's principle, why phenolphthalein is colourless when a solution of hydrochloric acid is added to it.

Question 8

A certain species of fish cannot survive in water having a pH less than 5,5. In a river the hydrogen ion concentration was measured to be $3,2 \times 10^{-5}$ $mol{\cdot}dm^{-3}$.

8.1 Will this species of fish survive in this river? Show clearly how you arrived at the answer.

8.2 A certain compound has ammonium chloride as the main ingredient. A group of learners investigated the amount of ammonium chloride present in a sample of the compound. The learners added 100 cm^3 of a 1,0 $mol{\cdot}dm^{-3}$ solution of sodium hydroxide to the sample in a flask.

8.2.1 Calculate the amount (in moles) of sodium hydroxide.

This mixture was warmed to remove the ammonia formed. The excess sodium hydroxide was then neutralised through a titration by 45 cm^3 of a 0,3 $mol{\cdot}dm^{-3}$ sulphuric acid solution. The relevant equations are:

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

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

Calculate the following:

8.2.2 The number of moles of sodium hydroxide that was neutralised by the sulphuric acid solution

8.2.3 The mass (in gram) of ammonium chloride in the sample

The following indicators are available for the titration.

Indicator	**pH range in which the colour changes**
Methyl orange	3,2 – 4,4
Bromothymol blue	6,0 – 7,6
Phenolphthalein	8,2 – 10,0

8.2.4 Select from the list and write down the best indicator to use for this investigation.

8.2.5 Give a reason for your choice in QUESTION 8.2.4.

Question 9

9.1 Eight grams (8,0 g) of sodium hydroxide are dissolved in 350 cm^3 of distilled water. 15 cm^3 of this solution neutralises 20 cm^3 of a sulphuric acid solution.
The balanced equation for this reaction is:

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

Calculate the concentration of the sulphuric acid solution.

9.2 An environmental disaster threatens the small town of Bafanaville. There has been a large spillage of concentrated hydrochloric acid ($HC\ell$) into the town's only water storage dam.
The pH of the water has decreased to 4.

9.2.1 Explain, with the aid of a chemical equation, why the pH of the dam water decreased.

The Municipality decides to add quantities of soda ash (Na_2CO_3) to the water of the dam, hoping that the pH will be restored to a value close to 7.

9.2.2 Calculate the mass of soda ash (Na_2CO_3) required to neutralise each 1 dm^3 of the acidified dam water.

9.2.3 Besides neutralisation, what other effect will the addition of the Na_2CO_3 have on the water in the dam?