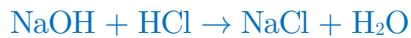


A sample of impure CaCO_3 weighing 0.245g is dissolved in 100mL of standardised 0.0905mol/L HCl solution. After the reaction is complete, it's found that the excess acid required 17mL of NaOH solution for complete neutralisation using phenolphthalein as an indicator. Earlier it was found 25mL of NaOH solution required an average 69mL of standardised HCl solution for a bromothymol blue indicator to change colour. Find the % of CaCO_3 in the impure sample.

[10 marks]



$$n(\text{HCl}) = 0.069 \times 0.0905 = 0.00624 \text{ mol}$$

$$n(\text{NaOH}) \text{ excess in 25mL} = n(\text{HCl}) = 0.00624 \text{ mol}$$

$$n(\text{NaOH}) \text{ reacted with excess acid} = 0.00624 \times \frac{17}{25} = 0.00425 \text{ mol}$$

$$n(\text{HCl}) \text{ excess} = n(\text{NaOH}) \text{ reacted with excess acid} = 0.00425 \text{ mol}$$

$$n(\text{HCl}) \text{ initial} = 0.1 \times 0.0905 = 0.0905 \text{ mol}$$

$$n(\text{HCl}) \text{ reacted} = 0.00480 \text{ mol}$$



$$n(\text{H}_3\text{O}^+) = n(\text{HCl}) = 0.00480 \text{ mol}$$

$$n(\text{CaCO}_3) = 0.00480 \times 0.5 = 0.00240 \text{ mol}$$

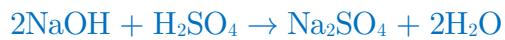
$$m(\text{CaCO}_3) = 0.00240 \times M(\text{CaCO}_3) = 0.240 \text{ g}$$

$$\% \text{ purity} = \frac{0.240}{0.245} \times 100 = 98.12\% = 98\%$$

Answer: ~98%

10mL of 0.1mol/L NaOH is added to 3mL of 0.05mol/L H_2SO_4 . Calculate the pH of the resultant solution (assume 100% ionisation of H_2SO_4).

[7 marks]



$$n(\text{NaOH}) = 0.1 \times 0.01 = 0.001 \text{ mol}$$

$$\text{mol ratio } (\text{NaOH}) = \frac{n(\text{NaOH})}{2} = \frac{0.001}{2} = 0.0005 \text{ mol}$$

$$n(\text{H}_2\text{SO}_4) = 0.05 \times 0.003 = 0.00015$$

$$\text{mol ratio } (\text{H}_2\text{SO}_4) = \frac{n(\text{H}_2\text{SO}_4)}{1} = \frac{0.00015}{1} = 0.00015 \text{ mol}$$

Hence NaOH is the LR as less is present on a mol-to-mol basis

$$n(\text{NaOH}) \text{ reacted} = n(\text{H}_2\text{SO}_4) \times 2 = 0.0003 \text{ mol}$$

$$n(\text{NaOH}) \text{ left} = 0.0007 \text{ mol}$$

$$n(\text{OH}^-) \text{ left} = n(\text{NaOH}) = 0.0007 \text{ mol}$$

$$[\text{OH}^-] = \frac{0.0007}{0.01+0.003} = 0.0538 \text{ mol/L}$$

$$\text{pOH} = 1.27$$

$$\text{pH} = 12.7$$

Answer: 12.7

Electrolytes

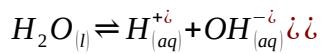
Electrolyte: A substance that dissolves in water and forms ions.

Strong electrolytes are good conductors of electricity in aqueous solution because water causes them to dissociate (ionic substances) or completely ionise (covalent molecular substances). Strong acids and ionic substances are strong electrolytes.

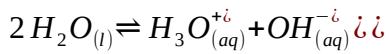
Weak electrolytes are poor conductors of electricity in aqueous solution. Only a small percentage of molecules ionise/dissociate when dissolved in water and so only a small concentration of ions is produced to conduct charge. Weak acids and weak covalent molecular bases tend to be weak electrolytes.

Non-electrolytes are non-conductors of electricity because no ions are produced when they dissolve in water.

Self-Ionisation of Water



or



In water at 25°C, $[H^+] = [OH^-]$ so the water is neutral (both are $1.00 \times 10^{-7} \text{ mol L}^{-1}$).

Equilibrium constant for the self-ionisation of water: $K_w = [H^+][OH^-] = [H_3O^+][OH^-]$

$K_w = 1.00 \times 10^{-14}$ at 25°C.

The self-ionisation of water is an endothermic process. An increase in temperature will favour the endothermic (forward) reaction which would result in an increased equilibrium constant. (e.g., at 40 °C, $K_w = 2.916 \times 10^{-14}$).

Similarly, at lower temperatures, the exothermic (reverse) reaction is favoured and so the equilibrium constant is reduced (e.g., at 10 °C, $K_w = 2.93 \times 10^{-15}$).

If it's pure water, no matter what the value of K_w is, $[H^+] = [OH^-]$.

Since $[H_3O^+][OH^-] = 1.00 \times 10^{-14}$, for all solutions at 25 °C, the concentration of one of these ions can be calculated given the concentration of the other.

Acidic solutions	$[H^+] > [OH^-]$
or	$[H^+] > 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25°C
and	$[OH^-] < 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25°C
Basic solutions	$[H^+] < [OH^-]$
or	$[H^+] < 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25°C
and	$[OH^-] > 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25°C
Neutral solutions	$[H^+] = [OH^-]$
or	$[H^+] = 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25°C
and	$[OH^-] > 1.00 \times 10^{-7} \text{ mol L}^{-1}$ at 25°C

Acidity constant (K_a) and acid strength:

An acid-base transfer reaction is an example of an equilibrium process. The extent to which an acid ionises in aqueous solutions can be determined from the equilibrium constant K_a also known as the acidity constant. It is a measure to which the proton transfer goes to completion.

The larger the K_a the greater the tendency of the acid to donate a proton to water i.e., the stronger the acid is the greater the degree of its ionisation.

A K_a value at 25°C of 1.8×10^{-5} indicates the reaction only proceeds to a very limited extent and only about 1% of ethanoic acid molecules are ionised.

The larger the value of K_a the more the equilibrium position favours products, hence the greater the tendency of the acid to form hydronium ions and so the stronger the acid is.

The K_a value also indicates the strength of the acid's conjugate base to accept a proton back from water so, if K_a is large (strong acid) its conjugate base is weak. As K_a is progressively smaller, the acid is progressively weaker, and the conjugate base becomes progressively stronger. (The stronger the acid is the weaker its conjugate-base).

Even though pure water at temperatures above 40 °C has a pH less than 7, it isn't acidic (this is because a neutral pH of 7 only applies at 25 °C). Pure water at, whatever temperature, is always neutral because the $[H_3O^+] = [OH^-]$.

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

An increase in pH by 1 means that the solution is only $\frac{1}{10}$ as acidic.

Arrhenius theory:

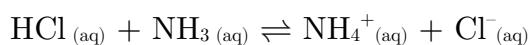
- Acids have H in their formula and produce hydrogen ions in aqueous solution.
- Bases have OH in their formula and produce hydroxide ions in aqueous solution.
- Neutralisation occurs when H⁺ ions and OH⁻ ions combine to form H₂O (l).
- Limited to aqueous solutions.

Problems with the theory:

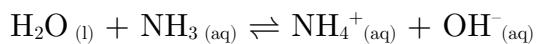
- Some bases produce OH⁻ ions in solution yet do not have OH in their formula (eg NH₃ and CO₃²⁻).
- Restricted to aqueous solutions.
- Not all salts are neutral.
- Does not allow for the existence of hydronium ions (H₃O⁺).

Bronsted-Lowry theory:

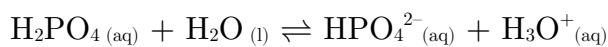
- Acids act as proton (H⁺) donors.
- Bases act as proton (H⁺) acceptors.



HCl is acting as an acid by donating a proton to NH₃. NH₃ is acting as a base as it's accepting a proton from H₂O.

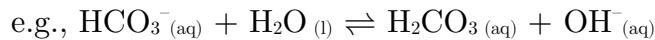


H₂O is acting as an acid by donating a proton to NH₃. NH₃ is acting as a base as it's accepting a proton from H₂O.



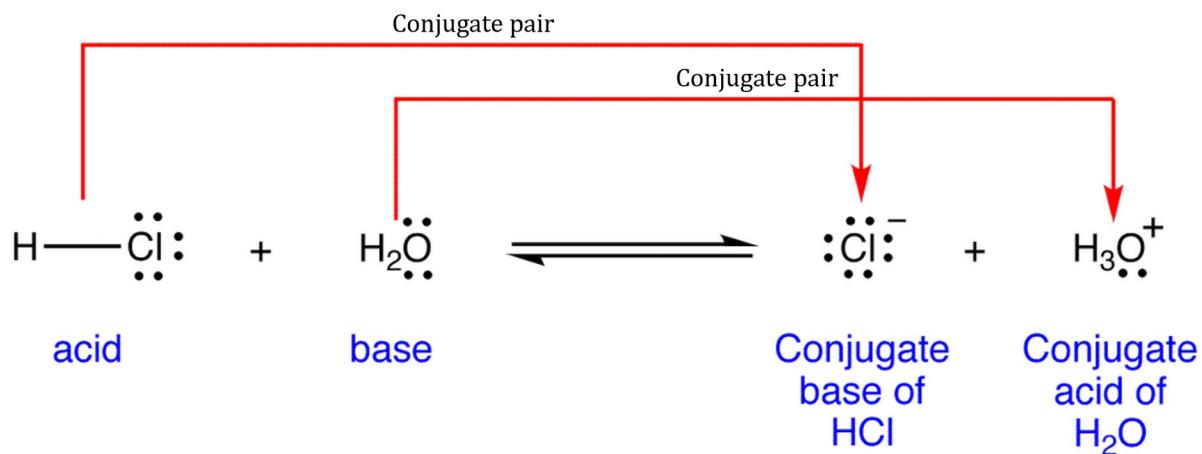
H_2PO_4 is acting as an acid by donating a proton to H_2O . H_2O is acting as a base as it's accepting a proton from H_2PO_4 .

Neutralisation: Reaction between a proton donor and a proton acceptor.



When an acid donates a proton, it has the potential to act as a base, hence the substance formed is its conjugate base. When a base accepts a proton, it has the potential to act as an acid, hence the substance formed is its conjugate acid.

The acid and the base it forms (or the base and the acid it forms) are called a conjugate acid-base pair.



In any conjugate acid- base pair the formula of the acid is always greater by the equivalent of H^+ .

The stronger the acid the weaker its conjugate base.

Acids & Bases

Strong acids completely ionise in water (e.g., $\text{HCl} \rightarrow \text{H}^{+\text{(aq)}} + \text{Cl}^{-\text{(aq)}}$).

Weak acids partially ionise in water (e.g., $\text{CH}_3\text{COOH} \rightleftharpoons \text{H}^{+\text{(aq)}} + \text{CH}_3\text{COO}^{-\text{(aq)}}$).

Strong acids:	Weak acids:
Hydrochloric acid (HCl)	Oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$)
Sulfuric acid (H_2SO_4)	Phosphoric acid (H_3PO_4)
Nitric acid (HNO_3)	Hydrofluoric acid (HF)
Hydrobromic acid (HBr)	Ethanoic acid (CH_3COOH)
Hydroiodic acid (HI)	Carbonic acid (H_2CO_3)
Perchloric acid (HClO_4)	Hypochlorous acid (HClO)

Strong bases completely dissociate in water to produce hydroxide ions (e.g., $\text{NaOH}_{(\text{aq})} \rightarrow \text{Na}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$).

Weak bases partially dissociate in water to produce hydroxide ions (e.g., $\text{NH}_3_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{NH}_4^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$).

Strong bases:	Weak bases:
Soluble:	Ammonia (NH_3)
Sodium hydroxide (NaOH)	Carbonate ion (CO_3^{2-})
Potassium hydroxide (KOH)	Ethanoate ion (CH_3COO^-)
Slightly soluble:	Fluoride ion (F^-)
Calcium hydroxide (Ca(OH)_2)	Phosphate ion (PO_4^{3-})
Magnesium hydroxide (Mg(OH)_2)	

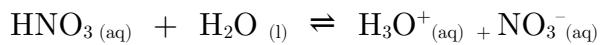
Hydrolysis of Salts

Hydrolysis: The process in which an ion reacts with water to produce H_3O^+ or OH^- ions. Ions that produce H_3O^+ ions are said to be acidic while ions that produce OH^- are said to be basic.

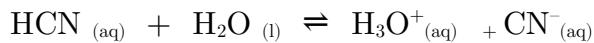
	Acidic:	Basic:	Neutral:
Anion	HSO_4^- , H_2PO_4^-	Other anions	Cl^- , NO_3^- , I^- , Br^-
Cation	NH_4^+ , Fe^{3+} , Sn^{4+} , Cr^{3+} , Al^{3+}	Not common	Na^+ , Li^+ , Mg^{2+} , K^+ , Ca^{2+} , Ba^{2+}

Acidic anions	HSO_4^- and H_2PO_4^- (i.e., anion from a polyprotic acid)
Basic anions	CH_3COO^- , HPO_4^{2-} , PO_4^{3-} , SO_4^{2-} , HCO_3^- , CO_3^{2-} , ClO^- , HS^- , CN^- , S^{2-} , F^- (anions from a weak acid)
Neutral anions	Cl^- , NO_3^- , Br^- , I^- (anions from strong acids)
Acidic cations	NH_4^+ , Al^{3+} , Fe^{3+} , Cr^{3+} (cation from a weak base or from aquated metal ions)
Basic cations	None covered in the syllabus
Neutral cations	Na^+ , K^+ , Ca^{2+} , Mg^{2+} , Li^+ , Ba^{2+} Group I and II metal cations

Acid-base reactions tend to occur in the direction in which the stronger acid and stronger base react to form a weaker acid and weaker base e.g.,



Equilibrium lies to the right.



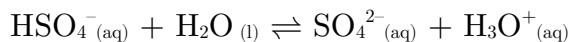
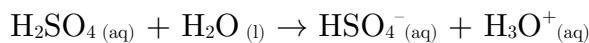
Equilibrium lies to the left.

Monoprotic acids can donate only one proton (hydrogen ion) per molecule (CH_3COOH).

While weak monoprotic acids produce fewer hydrogen ions in solution than strong acids, they react to the same extent with bases provided concentration is the same.

Diprotic acids contain 2 acidic protons per molecule.

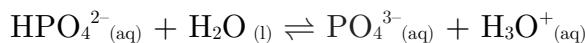
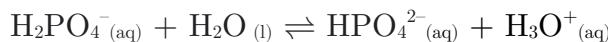
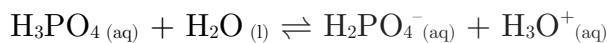
Q: Explain (using equations) why one mole of H_2SO_4 yields between one and two moles of H_3O^+ ions in solution.



Successive ionisation: The process of becoming ions over multiple steps.

Polyprotic acids contain 2 or more acidic protons per molecule.

Q: Use equations to show the successive ionisation of phosphoric acid.



Acid-Base Properties of Salts



Although this is called a neutralisation reaction, the salt formed isn't always neutral. An ion present in the salt can donate (becoming acidic) or accept (becoming basic) a proton via reacting with water so can be acidic or neutral depending on the reaction.

The reaction of the ions with water is called hydrolysis.

Buffer Solutions & Buffer Capacity

Buffer solution: A mixture of a weak acid and its conjugate base or of a weak base and its conjugate acid.

Buffer solutions can maintain a relatively constant pH when quantities of acidic or basic materials are added to them.

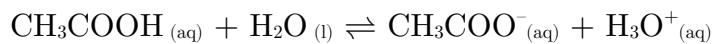
They are often prepared by adding the salt of the conjugate acid or base to the weak base or acid (respectively).

For a buffer to work, the amount of weak acid and its conjugate base must be large compared to the amount of H^+ and OH^- added or removed.

Buffer capacity: The ability of the buffer to resist change changes in pH. It refers to the amount of H^+ or OH^- that can be added without causing a 1 unit change to the pH of the buffer solution.

Buffering capacity increases as the concentration of the acid/conjugate base pair increases. The greater the concentration of the acid/conjugate base pair, the greater the amount of H^+ or OH^- that can be added without overwhelming the buffer's capacity to maintain a constant pH.

Q: A buffer solution was made by dissolving enough CH_3COOH and NaCH_3COO in water to produce a solution that contained 0.4mol of CH_3COOH and 0.2mol of CH_3COO^- .



[a] Would the buffer operate within its capacity if 0.01mol of NaOH was added?

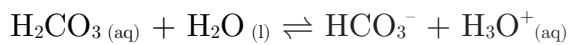
0.01mol of NaOH would neutralise 0.01mol of H_3O^+ . The forward reaction would be favoured to partially counteract this change. As there's 0.4mol of CH_3COOH available for the production of H_3O^+ (as per the forward reaction) the buffer would be operating within its capacity and be able to maintain a constant pH.

[b] Would the buffer operate within its capacity if 0.5mol of NaOH was added?

0.5mol of NaOH would neutralise 0.5mol of H_3O^+ . The forward reaction would be favoured to partially counteract this change. As there's 0.4mol of CH_3COOH available for the production of H_3O^+ (as per the forward reaction) the buffer would

be overwhelmed. It wouldn't have sufficient moles of CH₃COOH to replace the H₃O⁺ consumed by the NaOH. The pH of the buffer solution would change markedly.

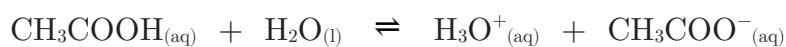
Exercise tends to increase the acidity of blood. The presence of H₂CO₃ and its conjugate base HCO₃⁻ help to maintain the blood's pH at 7.4.



During exercise, more H₃O⁺ is released into the blood from the muscles. Le Châtelier's principle predicts that the reverse reaction would be favoured to counteract the increased [H₃O⁺] and effectively maintain a relatively constant blood pH.

2mol of ethanoic acid (CH₃COOH) is mixed with 2mol sodium ethanoate (NaCH₃COO) and making the solution volume up to 1.00L. This solution contains a high concentration of weak acid, CH₃COOH, and its conjugate base, CH₃COO⁻ (from NaCH₃COO).

The equilibrium set up is given as:



Q: Explain buffering using collision theory.

If acid is added to the system this increases the H⁺/ H₃O⁺ concentration.

This increases the frequency of collisions between H₃O⁺ and CH₃COO⁻ and so increases the rate of the reverse reaction relative to the forward reaction.

This will consume H⁺/ H₃O⁺ and reduce their concentration.

(While this occurs the concentration of the reactants is increasing and so the rate of the forward reaction increases until the rates of the forward and reverse reactions are

the same and) equilibrium is re-established where H_3O^+ concentration is very close to what it was originally maintaining pH.

If base is added to the system this decreases the $\text{H}^+/\text{H}_3\text{O}^+$ concentration.

This reduces the frequency of collisions between H_3O^+ and CH_3COO^- and so reduces the rate of the reverse reaction relative to the forward reaction.

This will increase the $\text{H}^+/\text{H}_3\text{O}^+$ concentration.

(The rate of the reverse reaction then increases until the rates of the forward and reverse reactions are the same and) equilibrium is re-established where H_3O^+ concentration is very close to what it was originally maintaining pH.

The buffer capacity of a solution depends on 2 factors:

1. The relative concentrations of the weak acid (base) and its conjugate base (acid) – equal concentrations of these two offers the best buffering capacity.
2. The concentration of the weak acid (base) and its conjugate base (acid) – the greater the concentrations (in general) the greater the buffering capacity.

Acid-Base Titrations

Titration: The procedure of adding one solution to the other for a complete reaction.

2 types of titrations:

1. Acid-base titrations.
2. Redox titrations.

The point at which the acid-base reaction is complete is called the equivalence point when reactants are in stoichiometric proportions. This is often visible as acids and bases are colourless, so an indicator or pH meter is required.

Since the products of an acid-base reaction may be acidic, basic or neutral, it's important to choose an indicator so that the end point matches the equivalence

point. An appropriately chosen indicator will show a colour change (end point) when the reaction is complete (equivalence point). A pH meter can also be used.

Indicators are weak acids or bases whose colour changes as they change to their conjugate form. They're chosen so that this colour change occurs at equivalence point.

Indicator:	pH range:	Colour change acid/base:	Suitable for:
Methyl orange	3.1–4.4	Red → yellow	Strong acid + weak base
Bromophenol blue	4–4.6	Yellow → purple	
Methyl red	4.2–6.3	Red → yellow	Strong acid + strong base
Phenolphthalein	8.3–10	Colourless → pink	Weak acid + strong base

The pH of solutions of salts depends on the relative strengths of the anion as a base and the cation as an acid. Salts produced by reactions between a weak acid and weak base are very close to neutral. The equivalence point is not distinct (no great change in pH). As such, a pH meter is definitely the more preferred choice to monitor this type of acid-base titration.

Indicator:	Colour in acid:	Colour in base:	pH range:
Universal indicator	Red	Purple	0–14
Methyl orange	Red	Yellow	3.1–4.4
Bromophenol blue	Yellow	Purple	3–4.6
Methyl red	Red	Yellow	4.4–6.2
Litmus	Red	Blue	4.5–8.3
Bromothymol blue	Yellow	Blue	6–7.6
Phenolphthalein	Colourless	Pink	8.3–10
Alizarine yellow R	Yellow	Red	10.2–12

For an indicator (which is a weak acid) in solution, the following reaction takes place:



The colour changes are due to the difference in the colours of HIn and In^- .

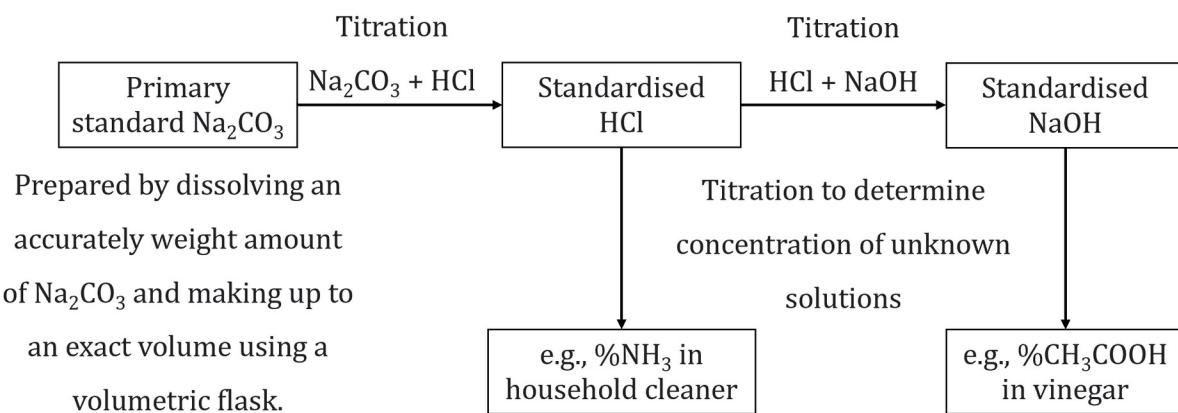
A primary standard is required to establish the concentration of unknown solutions. Its concentration must be accurately known. To be a useful indicator, the end point should match the equivalence point.

Characteristics of a primary standard:

- High degree of purity and known formula.
- Stable – Doesn't decompose in air or react with gases in air e.g., CO_2 , N_2 , O_2 , etc.
- Relatively high molar mass – to minimise weighing errors (minimising percentage uncertainty).
- Soluble in water.
- The reactions it's involved in are known.
- Not hygroscopic (absorbs water from the atmosphere) or deliquescent (absorbs so much water from the atmosphere that it eventually dissolves in the water it absorbs).

2 commonly used primary standards are anhydrous sodium carbonate (Na_2CO_3), potassium hydrogen phthalate ($\text{KHC}_8\text{H}_4\text{O}_4$) and hydrated oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$).

Standardising solutions:

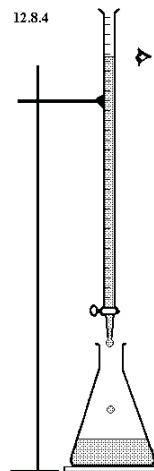


5 main pieces of equipment are used in titrations:

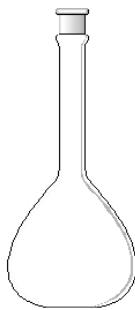
1. Pipette – Used to accurately deliver a known volume (aliquot) of liquid. It's rinsed with the solution to be delivered from it prior to use.



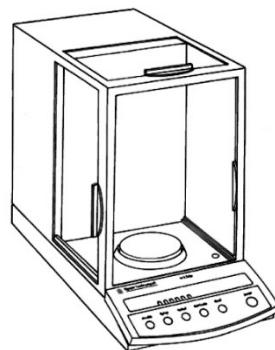
2. Burette – Used to accurately deliver a variable volume (titre) of liquid. It's rinsed with the solution to be delivered from it prior to use.



3. Volumetric flask – Used to hold an accurately known volume of solution **under the burette**(is it right?). It's rinsed with de-ionised or distilled water prior to use.



4. Analytical balance – Used to weigh out an accurately known mass.



5. Conical flask – Used to hold the solutions during the titration. It's rinsed with de-ionised or distilled water prior to use.



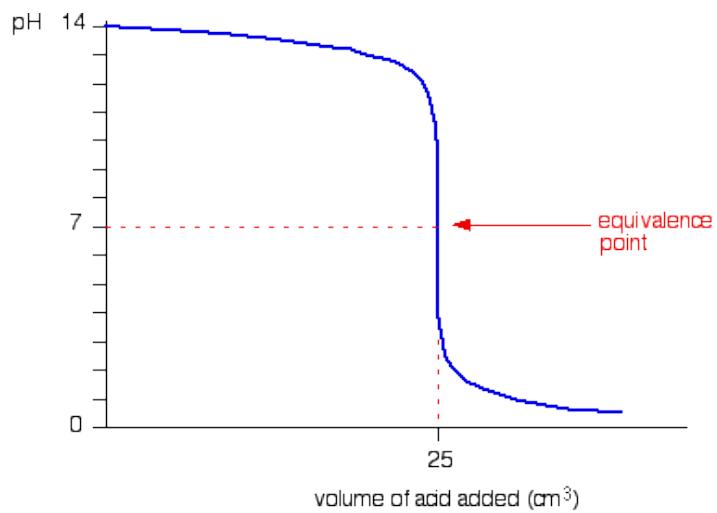
Titration process:

1. An aliquot of the standard solution is added to a conical flask.
2. A titre of the unknown solution is then carefully added from a burette into the conical flask until the reaction between the acid and base is complete.
3. This theoretical point in the titration when neither acid nor base remain is called the equivalence point.
4. The equivalence point occurs when stoichiometrically equivalent amounts of acid and base (reactants) have reacted.

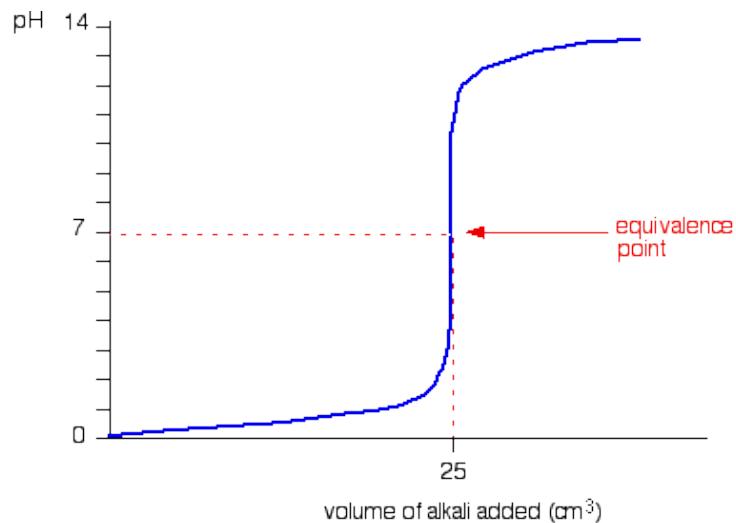
In a back titration, a known volume of an excess reactant is added to the reactant being analysed. A titration is then carried out to determine the amount of excess reactant remaining after the reaction.

Changes in pH During Acid-Base Titrations

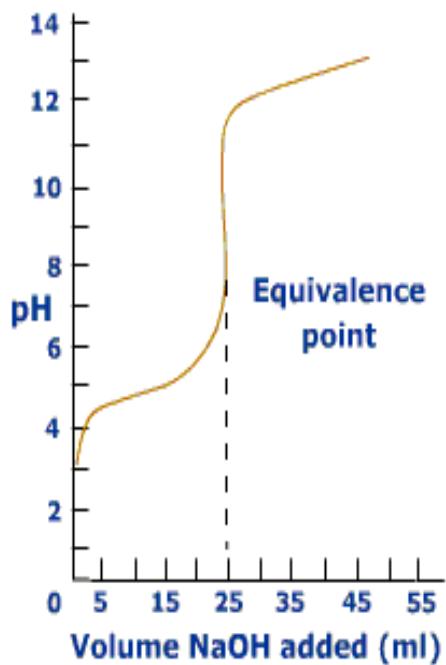
Running strong acid into strong base:



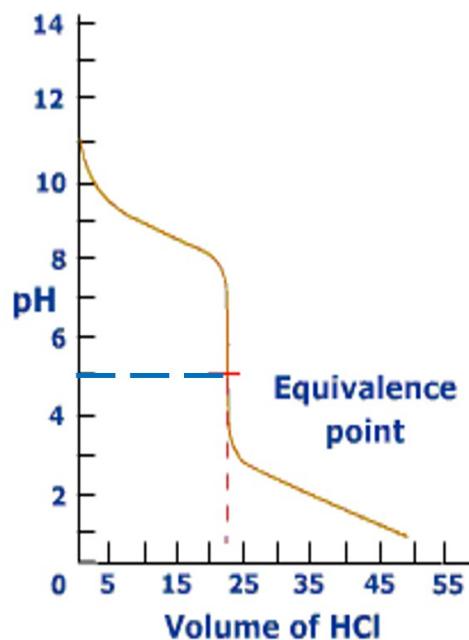
Running strong base into strong acid:



Running strong base into weak acid:

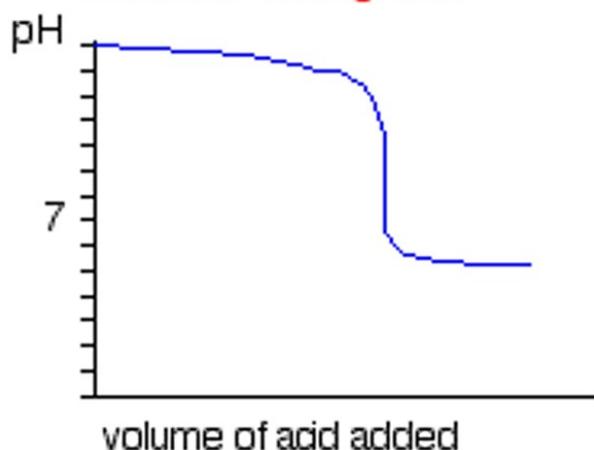


Running strong acid into weak base:



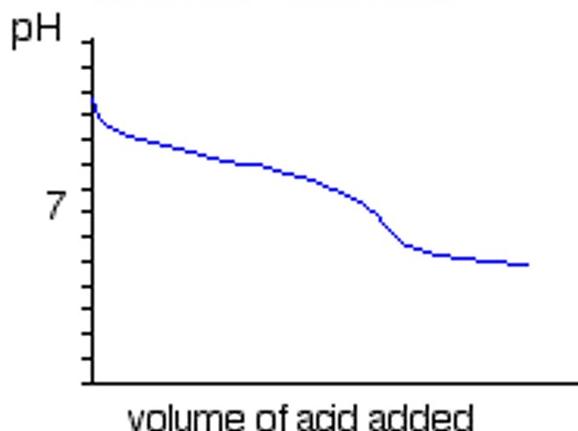
Running weak acid into strong base:

weak acid - strong base

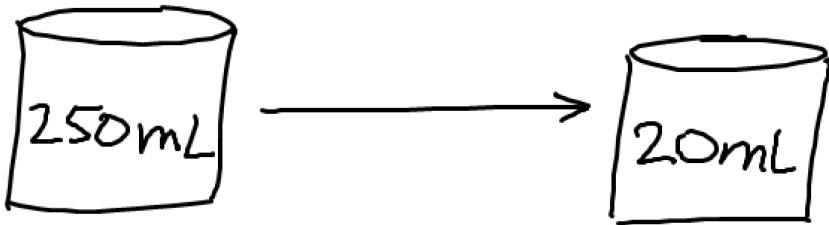


Running weak acid into weak base:

weak acid - weak base

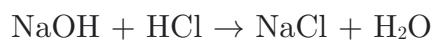


Q: A 1.252g sample of an oven cleaner containing mostly sodium hydroxide was dissolved and made up to 250mL using a volumetric flask. A 20mL aliquot of this solution was titrated against 0.095mol L⁻¹ HCl and required 23.85mL of this acid to reach the end point. Calculate the percentage (by mass) of NaOH in the oven cleaner.



1.252g
glass cleaner

23.85mL HCl
at 0.095mol L⁻¹



$$n(\text{HCl}) = cV = (0.095)(0.02385) = 0.00227\text{mol}$$

$$n(\text{NaOH}) = n(\text{HCl}) = 0.00227\text{mol}$$

$$n(\text{NaOH}) \text{ in } 250\text{mL} = 0.00277 \times \frac{250}{20} = 0.0283\text{mol}$$

$$m(\text{NaOH}) \text{ in } 250\text{mL} = nM = (0.0283)(22.99+16+1.008) = 1.133\text{g}$$

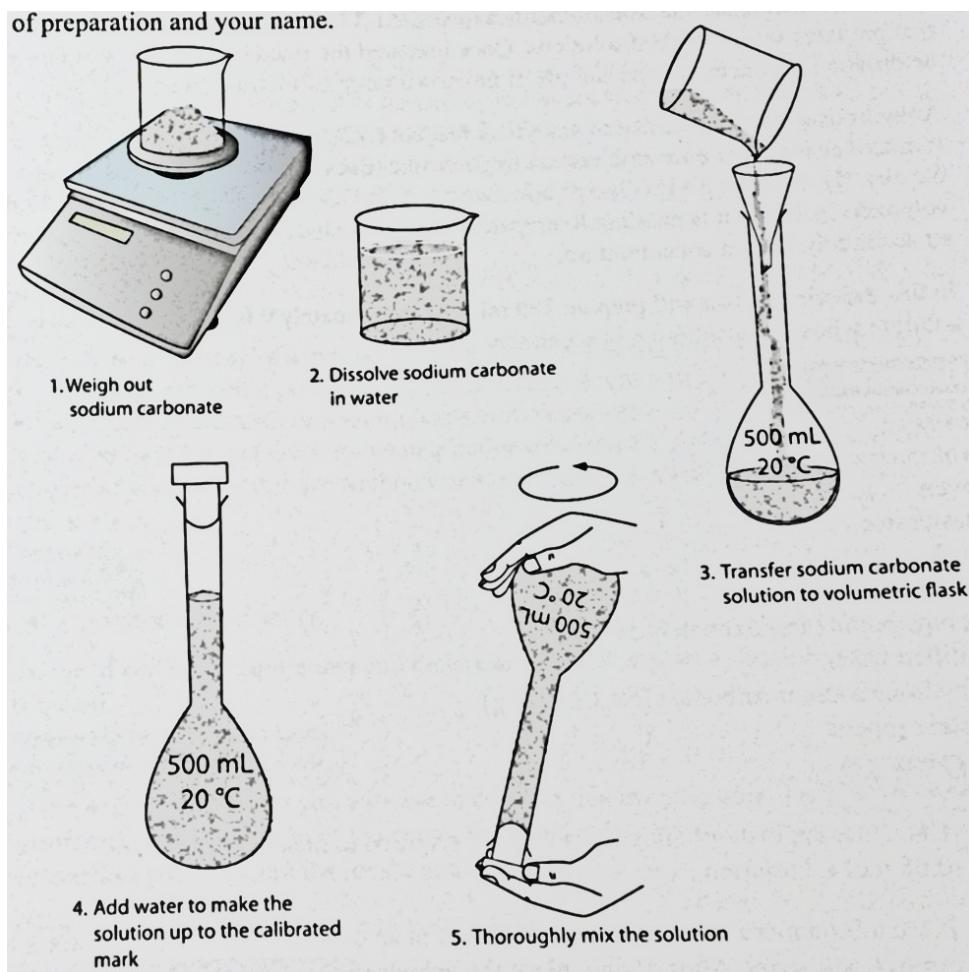
$$\% \text{NaOH} = \frac{1.133}{1.252} \times 100 = 90.5\%$$

STAWA Experiments 12

Method:

1. Calculate the mass of Na₂CO₃ ($m = nM = cVM$).
2. Place a little more than the required amount in an oven at 270°C for 30 minutes.
3. Place the Na₂CO₃ in a desiccator to cool.
4. Put the required mass of Na₂CO₃ (step 1) into a 250mL beaker.
5. Dissolve the solid in 100mL (some) distilled water.
6. Transfer this solution to a 500mL volumetric flask – rinse the beaker with about 20mL portions of distilled water.
7. Add distilled water to make the solution up to the calibrated mark (last few mL with a pipette).

8. Place the stopper in the volumetric flask and mix the solution thoroughly.
9. Transfer the solution to a clean storage bottle which should be rinsed with a bit of the Na_2CO_3 solution.



$$c = \frac{\frac{m(\text{Na}_2\text{CO}_3)}{M(\text{Na}_2\text{CO}_3)} * 1}{V}$$

Possible errors: Calibration error, evaporation of the Na_2CO_3 solution.

$$\% \text{ error} = \frac{|\text{true value} - \text{experimental value}|}{\text{true value}} \times 100$$

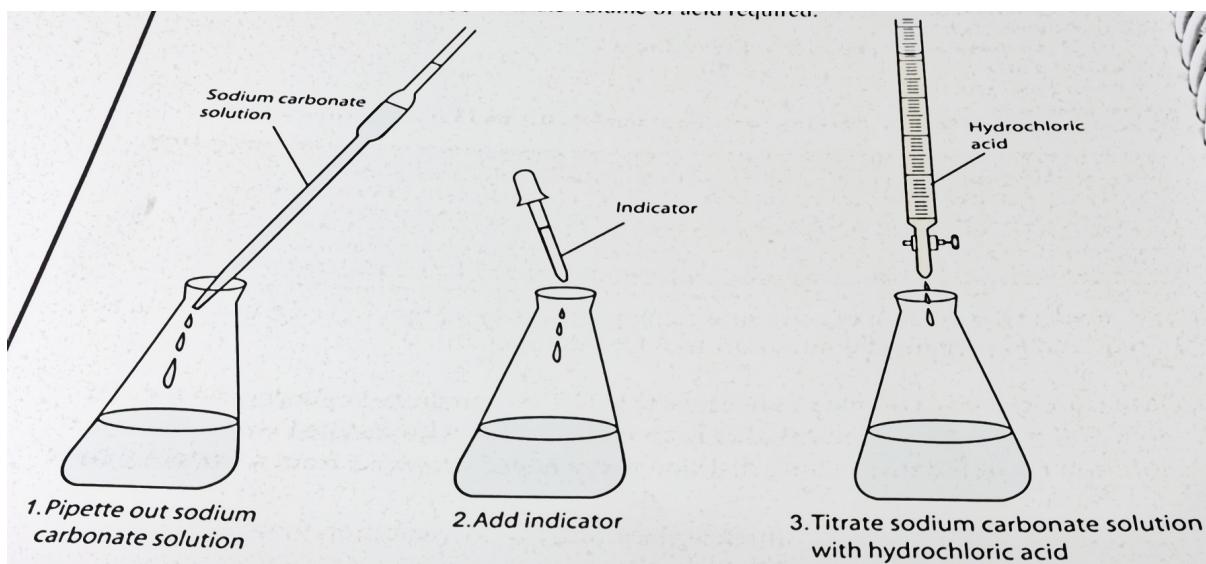
STAWA Experiment 13

Method – Part A: Making 0.1mol L⁻¹ HCl

1. Calculate volume of concentrated HCl to prepare 500mL of 0.1mol L⁻¹ solution ($c_1V_1 = c_2V_2$).
2. Transfer this volume of concentrated HCl to a 500mL volumetric flask that's about half filled with distilled water (last few mL with pipette).
3. Place the stopper in the volumetric flask and mix the solution thoroughly.
4. Transfer the approximately 0.1mol l⁻¹ HCl to a clean storage bottle that's been rinsed with a little bit of the HCl solution.

Method – Part B: Standardisation of the HCl solution

1. Place 100mL of standard Na₂CO₃ solution into a clean beaker. If wet, rinse with a little bit of the Na₂CO₃ solution first.
2. Rinse a clean 20mL pipette with some of the Na₂CO₃ solution.
3. Pipette 20mL of Na₂CO₃ solution into a 250mL conical flask.
4. Rinse a clean burette with some of the HCl solution and then fill the burette with the solution.
5. Add 2-3 drops of your chosen indicator to the flask.
6. Rinse a clean burette with some of the HCl solution and then fill the burette with the solution.
7. Record the level of acid in the burette. Obtain a rough estimate of the titration volume by running acid quickly from the burette while constantly swirling the liquid in the conical flask. Stop delivery of acid as soon as permanent colour change is obtained.



8. Table of results:

	Rough estimate	Accurate titrations		
		1	2	3
Final reading (mL)				
Initial reading (mL)				
Titre (mL)				

9. Prepare another conical flask containing a 20mL aliquot of the Na_2CO_3 solution and 2-3 drops of indicator. This time add the acid quickly from the burette with constant swirling of the flask until the volume added is within 2-3mL of the approximate volume required. Rinse the inside of the conical flask with a jet of distilled water to return any splashed solution to the bulk solution. Continue adding acid drop by drop and with constant swirling until the addition of one drop is sufficient to produce permanent colour change. Note and record the level of the acid in the burette at the end point.

10. Repeat the accurate titration with further 20mL aliquots of Na_2CO_3 solution until consistent titration volumes are obtained. These should be within 0.2mL of each other.

STAWA Experiment #13 – acidic end point → methyl orange.

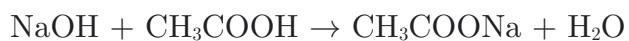
STAWA Experiment #14 – neutral end point → bromothymol blue.

STAWA Experiment #15 – basic end point → phenolphthalein.

Rinse burette and pipette with the solution.

Rinse conical flask with water.

STAWA Experiment #15:



Part 1:

1. Put vinegar in beaker.
2. Put blue thing on the pipette (shorter side). Brace the pipette.
3. Rinse pipette with vinegar (don't let it go out the top).
4. Put contents of pipette into a separate beaker.
5. Fill up the pipette until meniscus is on the graduated line (25mL).
6. Wipe the excess vinegar off the surface of pipette.
7. Take blue thing off and immediately seal the top hole with thumb.
8. Roll thumb to one side until meniscus is on graduated line if not already.
9. Put vinegar in volumetric flask.
10. Make the pipette at an angle to the vertical volumetric flask.

11. Count to 3 at the end (some left is normal).
12. Rinse the volumetric flask down with some distilled water (long tube).
13. Fill volumetric flask with distilled water until meniscus is under graduated line (500mL) (without long tube).
14. Last few mL with pipette until meniscus is on the graduated line.
15. Put stopper on volumetric flask and mix it around (inverting) (keep stopper on it).

Part 2:

1. Rinse burette with a bit of vinegar solution.
2. Put contents of pipette into the same separate beaker as initially.
3. Fill the burette with vinegar (with glass funnel) until above the zero mark.
4. Record initial reading on burette (doesn't have to be exactly at zero) (0.50).

Part 3:

1. Fill pipette with NaOH until above the graduated line (20mL).
2. Put NaOH into separate beaker until meniscus is on graduated line.
3. Put NaOH solution in conical flasks that are already rinsed with NaOH solution from beaker.
4. Count to 3 at the end (some left is normal).
5. Wipe excess solution off the surface of the pipette.
6. Add 2 drops of phenolphthalein into conical flask of NaOH and of (rough titration?).
7. Put burette filled with vinegar into conical flask filled with NaOH until one more drop would result in permanent colour change.
8. Record final reading on burette (28.70).

Part 4:

1. Fill burette with vinegar solution until around the 26mL mark.
2. Record initial reading of burette filled with the vinegar solution (1.30).
3. Fill the other flask (with phenolphthalein) with the vinegar from burette until one more drop would result in permanent colour change (around 27).
4. Record final reading on burette (28.80).

	Rough titration	Accurate titrations		
		1	2	3
Initial reading (mL)				
Final reading (mL)				
Titre (mL)				

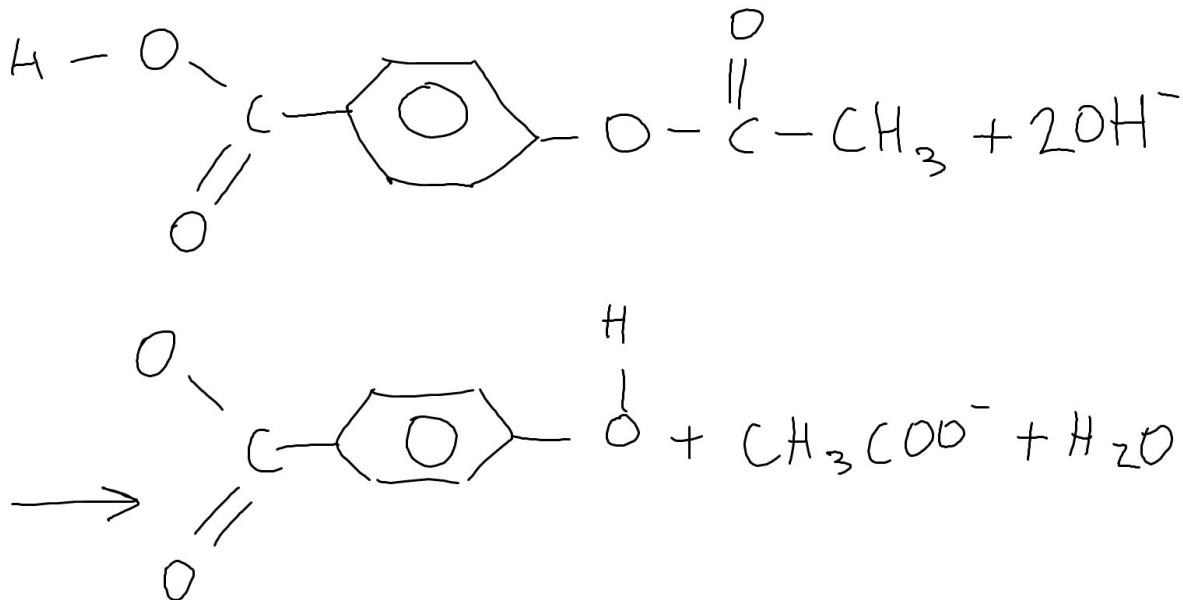
	Rough titration:	Titration 1:	Titration 2:	Titration 3:
Initial reading (mL)				
Final reading (mL)				
Titre (mL)				

Q: What's the effect on the calculated concentration when using the wrong indicator?

If you use the wrong indicator, then the equivalence point will be reached earlier or later depending on where the equivalence point is reached and the type of indicator used.

The primary standard, Na_2CO_3 , is basic. As the standard solution, HCl, is added, the pH progressively decreases with an equivalence point in the acidic range, since a strong acid is being added to a weak base. If bromothymol blue is used as an indicator, which has a neutral point, then the end point of the indicator will be reached too early, underestimating the volume of HCl required, hence increasing the calculated [HCl] as concentration is inversely proportional to volume.

You can't use NaOH as a primary standard because NaOH absorbs a lot of moisture from the atmosphere. That's why we have to use Na_2CO_3 to determine the



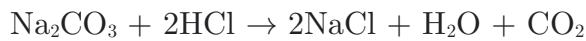
An analysis of Aspirin involves a back titration where NaOH is used to react with the $-\text{COOH}$ group and the $-\text{C}=\text{O}$ group in the molecule shown above.

A 2.00g Aspirin tablet was crushed and then boiled with 150.0mL of $4.52 \times 10^{-1}\text{mol L}^{-1}$ NaOH solution until dissolved. A 20.0mL aliquot of the solution was then titrated

with a HCl solution and required 17.65mL to reach the endpoint. Earlier, 25.0mL of the HCl solution was titrated against a standard 1.62×10^{-1} mol L⁻¹ Na₂CO₃ solution and required 28.9mL to reach the endpoint. Calculate the percentage purity of the Aspirin tablet.

[11 marks]

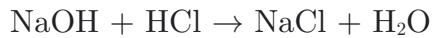
1. Find n(HCl) in the 25mL of HCl via $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$.
2. Find n(HCl) in the 17.65mL of HCl from step 1.
3. State that n(HCl) = n(NaOH) via $\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$.
4. Find n(NaOH)_{reacted} in the 150mL solution from step 3 with (initial – excess).
5. State that n(OH⁻) = n(NaOH).
6. Find n(aspirin) = n(OH⁻)/2 from the given equation.
7. Find m(aspirin) = n(aspirin) x M(aspirin).
8. Find % mass.



$$n(\text{Na}_2\text{CO}_3) = 0.162 \times 0.0289 = 0.00468 \text{ mol}$$

$$n(\text{HCl}) \text{ in } 25\text{mL} = 2 \times n(\text{Na}_2\text{CO}_3) = 0.00936 \text{ mol}$$

$$n(\text{HCl}) \text{ in } 20\text{mL} = 0.00936 \times \frac{17.65}{25} = 0.00661 \text{ mol}$$



$$n(\text{NaOH}) \text{ in } 20\text{mL} = n(\text{HCl}) = 0.00661 \text{ mol}$$

$$n(\text{NaOH}) \text{ in } 150\text{mL} = 0.00661 \times \frac{150}{20} = 0.0496 \text{ mol}$$

$$n(\text{NaOH}) \text{ reacted} = n(\text{NaOH}) \text{ initial} - n(\text{NaOH}) \text{ excess} = 0.0182 \text{ mol}$$

$$n(\text{acid}) = n \cancel{\text{HCl}} = 0.00911 \text{ mol}$$

$$m(\text{acid}) = 1.64 \text{ g}$$

$$\% \text{ mass} = \frac{1.64}{2} \times 100 = 82.1\%$$

Q: Name a suitable indicator that could've been used in the titration between HCl and Na₂CO₃. Justify your choice using appropriate chemical equations.

[6 marks]

Methyl orange.

The salt produced is NaCl from the titration.

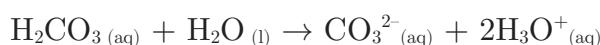


As HCl is a strong acid and Na₂CO₃ is a weak base, the resulting salt is neutral.

However, the CO₂ gas reacts with water, resulting in the formation of H₂CO₃.



H₂CO₃ then undergoes hydrolysis, forming H₃O⁺ ions.

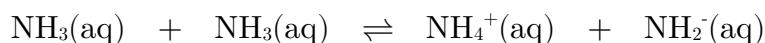


The increase in [H₃O⁺] means the solution is slightly acidic as [H₃O⁺]>[OH⁻]. As such, methyl orange is a suitable indicator as it changes colour in the range 3.0–4.5
(you can approximate 3.0–4.5 as the pH range for both methyl orange and bromophenol blue).

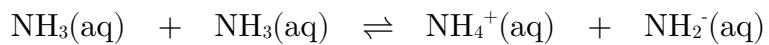
Theory Past Exam Questions

WACE 3AB 2010 Q28:

Like water, ammonia is able to react with itself, in the process known as ‘self-ionisation’. The equation for the self-ionisation of ammonia is below.



- (a) Identify the conjugate acid and base pairs in the reaction. Join each pair with a line, and label the conjugate acid and base of each pair appropriately.
mark)



(b) At standard temperature and pressure, the equilibrium constant, K, for this reaction is about 1×10^{-30} . The self-ionisation of ammonia is an endothermic process. Will the value of K be less than or greater than 1×10^{-30} at temperatures greater than 0°C ? Explain.

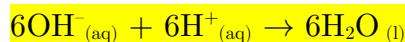
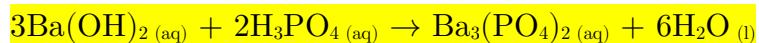
The reaction is endothermic.

Hence at $T > 0^\circ\text{C}$ (the temperature for which the equilibrium constant is given) the forward reaction will be favoured.

This means the concentration of products will increase relative to reactants, meaning the K value will increase, meaning the K value will be greater than 1×10^{-30} .

WACE 2AB 2014 Q38:

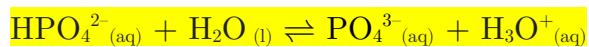
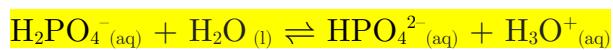
(b) Write an ionic equation for the reaction between phosphoric acid and barium hydroxide solution. Include state symbols.



(c) Phosphoric acid is a polyprotic acid. With the aid of equations, use phosphoric acid as an example to explain the term ‘polyprotic’

A polyprotic acid is an acid with more than one proton/ H^+ ion available for ionisation.

H_3PO_4 , a polyprotic acid, has 3 donatable hydrogens and hence can undergo 3 consecutive ionisation reactions, each donating one H^+ ion.



- (d) Phosphoric acid is a weak acid and a weak electrolyte. Barium hydroxide is a strong base and a strong electrolyte. Using equations containing phosphoric acid and barium hydroxide, explain the difference between the terms ‘strong’ and ‘weak’ when referring to electrolytes.

A strong electrolyte (e.g., $\text{Ba}(\text{OH})_2$) fully ionises in solution, resulting in a completely ionic solution.



A weak electrolyte (e.g., H_3PO_4) partially ionises in solution, resulting in a solution containing both ionic products and the original reactant.



WACE 2AB 2011 Q30:

The poisonous compound oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$) is found in significant quantities in the leaves of the rhubarb plant, while its stalks contain only trace amounts, making them safe to eat. Oxalic acid is a polyprotic acid.

- (a) Explain what is meant by the term ‘polyprotic’.

- **1 mark:** Acid with more than one ionisable/‘donatable’ proton/ H^+

- (b) Complete the table below by giving appropriate formulae.

Substance	Example
A polyprotic acid (other than oxalic	Any polyprotic acid. E.g. H_2SO_4 ,

acid)	H_3PO_4
A monoprotic acid	Any monoprotic acid. E.g. HCl , HNO_3

(c) Write the equations for the successive ionisation of oxalic acid.



(d) Explain why a 0.1 mol L^{-1} solution of oxalic acid would have a higher pH than a 0.1 mol L^{-1} solution of sulfuric acid.

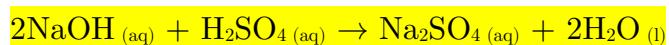
A strong acid e.g., H_2SO_4 will fully ionise in solution whereas a weak acid e.g., $\text{H}_2\text{C}_2\text{O}_4$ will partially ionise in solution.

Hence the full ionisation of H_2SO_4 will result in a greater $[\text{H}^+]$ than the partial ionisation of $\text{H}_2\text{C}_2\text{O}_4$ despite both acids being diprotic.

HSC 1996 Q25:

Understanding of acids and bases has changed since Arrhenius first developed his theory. Although an acid-base reaction is known as neutralisation, the resulting salt solution is not always neutral. For example, a solution of the salt sodium sulfate is neutral, but a solution of sodium ethanoate (acetate) is basic.

(a) Write an equation to describe the formation of sodium sulfate from an acid-base reaction. Name the reactants.

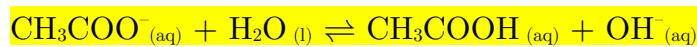


Reactants: Sodium hydroxide, sulfuric acid.

(b) Explain why a solution of sodium ethanoate (CH_3COONa) is basic, which a sodium sulfate solution of the same concentration has a pH of 7.0. Write ionic equations to describe any reactions.

CH_3COONa :

Na^+ ion is neutral whereas CH_3COO^- ion is basic.



The increase in $[\text{OH}^-]$ causes pH to increase above 7, creating a basic solution.

Na_2SO_4 :

Na^+ ion and SO_4^{2-} ion are both neutral.

Neither undergoes hydrolysis and so there's no change in pH.

TEE 2001 Q5:

(a) A 0.1 mol L⁻¹ solution of Na_2HPO_4 has a pH of about 10. Explain this, using an equation or equations.

Na^+ ion is neutral whereas HPO_4^{2-} ion is basic.

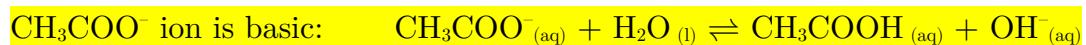


The increase in $[\text{OH}^-]$ causes pH to increase above 7, creating a basic solution.

(b) A 0.1 mol L⁻¹ solution of $\text{NH}_4\text{CH}_3\text{COO}$ (ammonium acetate) has a pH of approximately 7. Explain this, using at least two equations.

NH_4^+ ion is acidic:





The acidity of the NH_4^+ ion is countered by the basicity of the CH_3COO^- ion as both equations have similar K values and hence pH is around 7.

WACE 3AB 2015 Q40:

Hydrogen fluoride, HF, is a highly dangerous and corrosive liquid that boils at near room temperature. It readily forms hydrofluoric acid in the presence of water and is an ingredient used to produce many important compounds, including medicines and polymers.

- (b) The equilibrium constant (K) for the dissociation of hydrofluoric acid is 6.8×10^{-4} , and for hydrochloric acid K is very large. To make a solution of hydrofluoric acid with the same pH as hydrochloric acid, a greater concentration of hydrofluoric acid is required. Explain why this is so.

To have the same pH, both acids need to have the same $[\text{H}^+]$.

HF doesn't ionise to the same extent as HCl.

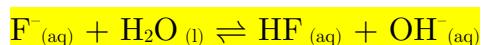
Hence a greater [HF] is needed to give the required $[\text{H}^+]$.

- (c) The salts, sodium chloride and sodium fluoride, readily dissolve in water. At 25.0°C the pH of the sodium chloride solution is equal to 7 whereas the pH of the sodium fluoride solution is greater than 7. Explain this difference in pH. Include any relevant equation(s) to support your answer.

Na^+ ion is neutral.

Cl^- ion is neutral and hence NaCl is neutral.

F^- ion is basic and hence NaF is basic.



The increase in $[\text{OH}^-]$ causes pH to increase above 7, creating a basic solution, creating a basic solution.

WACE 3AB 2012 Q36:

Water is able to react with itself in the process known as ‘self-ionisation’ or ‘auto-ionisation’.

- (a) Write the equation for the self-ionisation of water.



- (b) At 25 °C, the value of K_w is approximately 1.0×10^{-14} . At 10 °C, the value of K_w is approximately 2.9×10^{-15} .

What are the relative concentrations of H^+ and OH^- ions in a neutral water solution at 25 °C?

Circle the correct answer.

[H⁺] > [OH⁻] [H⁺] < [OH⁻] [H⁺] = [OH⁻]

What are the relative concentrations of H^+ and OH^- ions in a neutral water solution at 10 °C?

Circle the correct answer.

[H⁺] > [OH⁻] [H⁺] < [OH⁻] [H⁺] = [OH⁻]

(c) Consider the values of K_w at 10 °C and 25 °C, and state whether the self-ionisation of water is an endothermic or exothermic process. Give a reason to support your answer.

The K_w value is greater at 25°C than at 10°C.

K_w is the ratio of concentration of products to reactants.

Hence the forward reaction is favoured when temperature increases.

According to LCP, the endothermic reaction is favoured when temperature increases.

∴ The self-ionisation of water is an endothermic process.

TEE 2008 Q9:

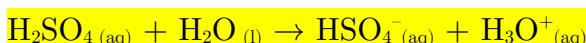
A student was given a 0.100 mol L⁻¹ sulfuric acid solution and a 0.200 mol L⁻¹ hydrochloric acid solution. She tested the pH of the solutions using a pH meter and found that the pH of the sulfuric acid solution higher than that of the hydrochloric acid solution. Explain this observation. Include equations in your answer.

As HCl is a strong acid, its ionisation in solution goes to completion.

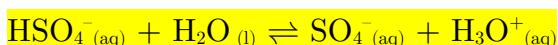


$$[\text{H}^+] = [\text{HCl}] = 0.2\text{mol L}^{-1}$$

H_2SO_4 is diprotic and undergoes 2 consecutive ionisations. The first goes to completion:



The second does not:



If both ionisation reactions went to completion then $[\text{H}^+] = 2 \times [\text{H}_2\text{SO}_4] = 0.2\text{mol L}^{-1}$

However, the second one doesn't go to completion and hence $[\text{H}^+] < 0.2\text{mol L}^{-1}$

$[H^+]$ of H_2SO_4 is less than $[H^+]$ of HCl

\therefore pH of H_2SO_4 is greater than pH of HCl

WACE 3AB 2015 Q40:

Propanoic acid, CH_3CH_2COOH , is a weak monoprotic acid. When 0.500 mol of sodium propanoate ($NaCH_3CH_2COO$) is dissolved in 1.00 L of 0.500 mol L^{-1} propanoic acid at 25.0 °C a buffer solution is formed.

- (d) (i) Addition of 10.0 mL of 1.00 mol L^{-1} $HCl(aq)$ to this buffer does not significantly change its pH. Explain this observation, including any relevant equation(s).



As hydrogen ions are added to the buffer equilibrium, the equilibrium will shift left to use up the added H^+ ions

Due to the shift in equilibrium, there is very little change to the overall concentration of hydrogen ions, and so the pH change is insignificant

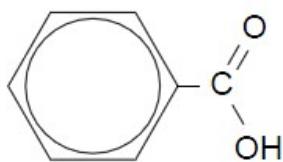
- (ii) State **two** conditions required to ensure that this system has a high buffering capacity.

One: Equal concentrations of acid and conjugate base

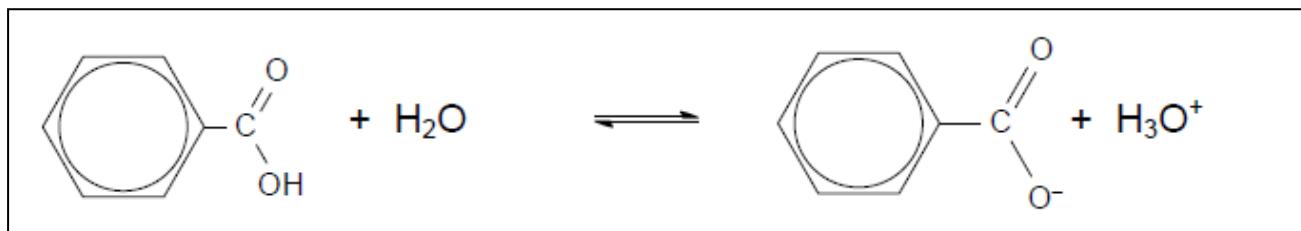
Two: High concentrations of acid and conjugate base

WACE 3AB 2010 Q29:

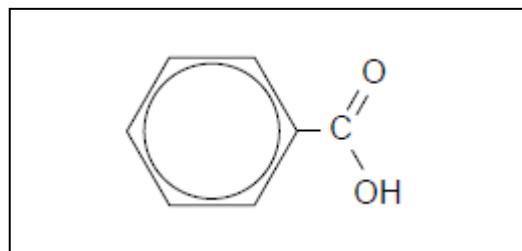
Benzoic acid is found in many berries and some other fruits, and is used as a food preservative. The structure of benzoic acid is shown below. In an aqueous environment, benzoic acid ionises and exists in equilibrium with the benzoate ion.



(a) Write the equation for the reaction between benzoic acid and water.



(b) Draw the structure (either benzoic acid or the benzoate ion) that would predominate in the acidic environment of the stomach.



Examiner's comments: The majority of candidates did not recognise that benzoic acid (rather than benzoate ion) would predominate in the acidic environment of the stomach; perhaps candidates decided that the acidic environment of the stomach would push the equilibrium to the right.

(c) Show, using equations and the principles of equilibrium, how a solution of benzoic acid and the benzoate ion may behave as a buffer.

When H^+ is added it will react with benzoate. This consumes the added H^+ , meaning no significant changes to $[\text{H}^+]$ or pH



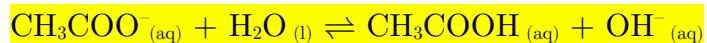
When OH^- is added it will react with benzoic acid. This consumes the added OH^- , meaning no significant changes to $[\text{OH}]$ or pH



HSC 2015 Q24:

- (a) Explain why the salt, sodium acetate, forms a basic solution when dissolved in water. Include an equation in your answer.

Na^+ ion is neutral whereas CH_3COO^- ion is basic.



The increase in $[\text{OH}^-]$ causes pH to increase, creating a basic solution, creating a basic solution.

- (b) A solution is prepared by using equal volumes and concentrations of acetic acid and sodium acetate.

Explain how the pH of this solution would be affected by the addition of a small amount of sodium hydroxide solution. Include an equation in your answer.



The added OH^- causes $[\text{H}_3\text{O}^+]$ to decrease.

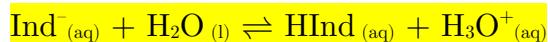
According to LCP, this causes the forward reaction to be favoured which produces H_3O^+ , and the equilibrium shifts right until a new equilibrium is established, minimising the change in pH.

HSC 2013 Q25:

An indicator is placed in water. The resulting solution contains the green ion, Ind^- , and the red molecule, $HInd$.

Explain why this solution can be used as an indicator. In your response, include a suitable chemical equation that uses Ind^- and $Hind$.

An indicator needs to change colour in different pH conditions.



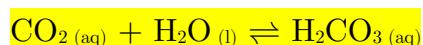
When a base is present, $[H_3O^+]$ is reduced. This causes the equilibrium to shift right, increasing $[HInd]$, and hence the red colour dominates.

When an acid is present, $[H_3O^+]$ is increased. This causes the equilibrium to shift left, increasing $[nd^-]$, and hence the green colour dominates.

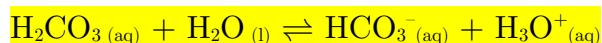
WACE 2016 Sample Exam Q34:

Ocean acidification results from carbon dioxide dissolving in water and an equilibrium being established between the water and carbon dioxide to produce carbonic acid (H_2CO_3).

(a) Write a balanced equation for this equilibrium.



(b) The formation of carbonic acid leads to an increase in the hydronium ion (H_3O^+) concentration in the water. Show the equilibrium that results in the formation of hydronium ions when carbonic acid reacts with water.



(c) State **one** problem that ocean acidification is causing for marine organisms.

Explain how this problem arises and support your answer with an appropriate balanced equation.

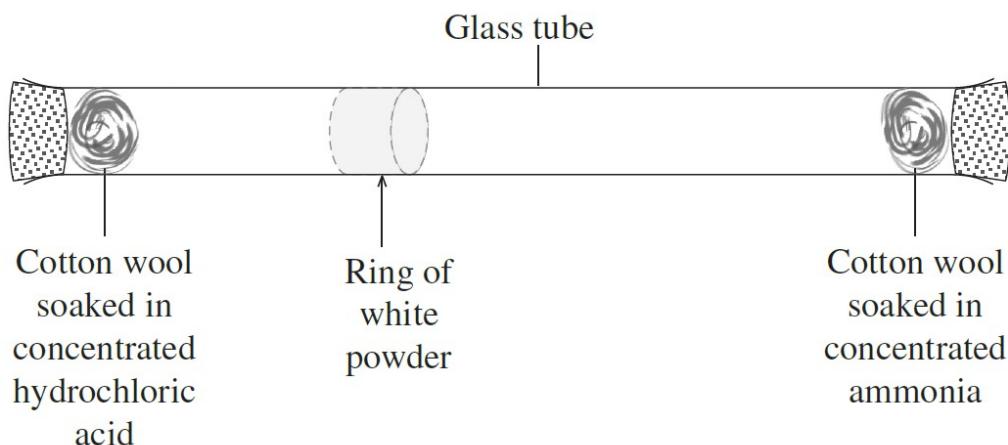


An increase in $[\text{CO}_2]$ causes the equilibrium in (a) to shift right, causing $[\text{H}_2\text{CO}_3]$ to increase. This causes the equilibrium in (b) to shift right, increasing $[\text{H}_3\text{O}^+]$. This causes the equilibrium in (c) to shift left, decreasing $[\text{CO}_3^{2-}]$.

A decrease in $[\text{CO}_3^{2-}]$ makes it difficult for marine organisms to develop CaCO_3 structures e.g., shells, exoskeletons, etc.

HSC 2015 Q28:

The equipment shown is set up. After some time a ring of white powder is seen to form on the inside of the glass tube.



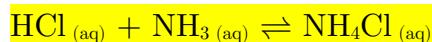
(a) Why would this NOT be an acid-base reaction according to Arrhenius?

The reaction doesn't occur in an aqueous solution.

(b) Explain why this would be considered a Brønsted-Lowry acid-base reaction.

Include an equation in your answer.

The reaction involves a Brønsted-Lowry acid (HCl) and a Brønsted-Lowry base (NH_3).



2021 Creelman:

Question 32 **(6 marks)**

- (a) A buffer of carbonic acid (H_2CO_3)/hydrogencarbonate (HCO_3^-) is present in blood plasma to maintain a pH between 7.35 and 7.45. Write an equation to show the relevant species present in a carbonic acid/hydrogencarbonate buffer solution. (2 marks)

- (b) Explain why 300.0 mL of 1.00 mol L⁻¹ carbonic acid/hydrogencarbonate buffer does **not** change in pH significantly when 3 drops of 1.00 mol L⁻¹ HCl are added to it, yet when 3 drops of 1.00 mol L⁻¹ HCl are added to 300.0 mL of distilled water there is a significant change in pH? (4 marks)

Question 32**(6 marks)**

- (a) A buffer of carbonic acid (H_2CO_3)/hydrogencarbonate (HCO_3^-) is present in blood plasma to maintain a pH between 7.35 and 7.45. Write an equation to show the relevant species present in a carbonic acid/hydrogencarbonate buffer solution. (2 marks)

Description	Marks
) (or reverse)	2
One error or omission	1
Total	2

- (b) Explain why 300.0 mL of 1.00 mol L⁻¹ carbonic acid/hydrogencarbonate buffer does **not** change in pH significantly when 3 drops of 1.00 mol L⁻¹ HCl are added to it, yet when 3 drops of 1.00 mol L⁻¹ HCl are added to 300.0 mL of distilled water there is a significant change in pH? (4 marks)

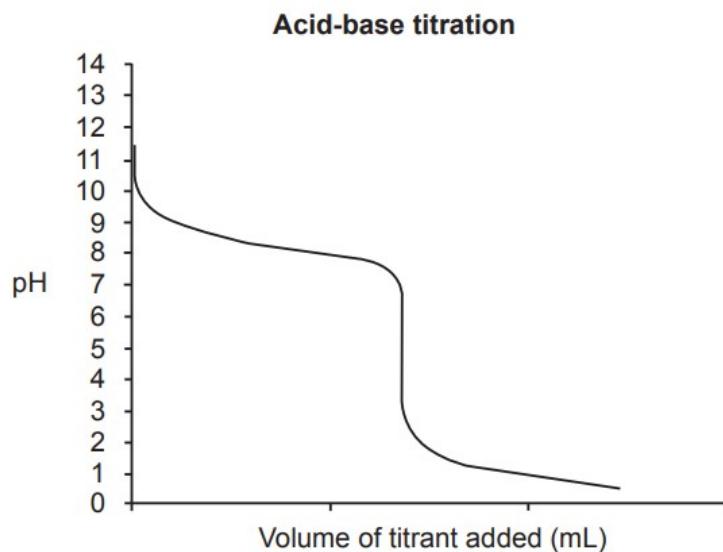
Description	Marks
Addition of HCl increases the concentration of H_3O^+	1
In the buffer there is sufficient HCO_3^- (or species appropriate to the student's equation above) to react and reduce the amount of added H_3O^+	
or	
The additional H_3O^+ reacts with the HCO_3^- (or species appropriate to the student's equation above) and the equilibrium shifts to the left	1
or	
The rate of the reverse reaction increases due to the additional H_3O^+ reacting with HCO_3^- more than the forward reaction resulting in a reduction in the amount of H_3O^+ present when equilibrium is re-established	
As the concentration of H_3O^+ is reduced to close to before the addition of HCl, there is a minimal change in pH	1
(While an equilibrium system, water cannot act as a buffer system because) in water there is insufficient OH^- to react with added H_3O^+ so all of the added H_3O^+ contributes to the significant change in pH	1
Total	4
Note: It is not adequate to simply state that water is not a buffer to earn the mark.	

Adding acid to the buffer solution increases $[\text{H}_3\text{O}^+]$, decreasing the distance between product particles. This increases the proportion of collisions on the RHS, increasing the rate of the reverse reaction relative to the forward reaction. This decreases $[\text{H}_3\text{O}^+]$. Equilibrium is then re-established. Since $[\text{H}_3\text{O}^+]$ is decreased to close to the original value, pH is maintained.

Question 34**(6 marks)**

The data below were collected from an acid-base titration.

- (a) Label the equivalence point on the titration curve below using an arrow and record the pH value at this point. (2 marks)



pH value at equivalence point: _____

- (b) Select an indicator from the table below that would be **best** for this titration and justify your choice. (4 marks)

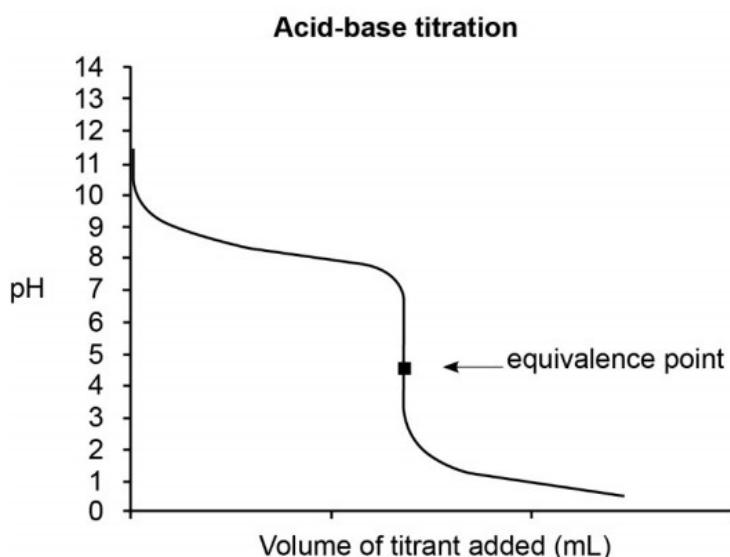
Indicator	Low pH colour	Transition pH range	High pH colour
Methyl Yellow	red	2.1 – 3.3	yellow
Bromocresol Green	yellow	3.8 – 5.4	blue
Bromothymol Blue	yellow	6.0 – 7.6	blue
Phenolphthalein	colourless	8.3 – 10.0	pink
Alzarine Yellow R	yellow	10.2 – 12.0	red

Indicator: _____

Justification: _____

Question 34**(6 marks)**

- (a) Label the equivalence point on the titration curve below using an arrow and record the pH value at this point. (2 marks)



Description	Marks
Label on Titration curve pointing to correct equivalence point pH 4.8 ± 0.5	1
pH stated within range of 4.8 ± 0.5 (pH 4.3 to 5.3)	1
Total	2

- (b) Select an indicator from the table below that would be **best** for this titration and justify your choice. (4 marks)

Description	Marks
Bromocresol Green	1
Justification may include up to, but not be limited to, three of the following recognitions: <ul style="list-style-type: none">That a change in colour of the indicator (end point) indicates that the equivalence point has been reachedThe (appropriate) indicator changes colour in the range of the equivalence pointBromocresol Green is the only indicator that changes colour in the range of the equivalence pointOther indicators will change colour either before (Phenolphthalein, Alizarine Yellow R) or after the equivalence point (Methyl Yellow, Bromocresol Blue)The use of indicators other than Bromocresol Green will produce an incorrect titre (too large or too small)The use of indicators other than Bromocresol Green will result in a systematic error or the use of Bromocresol Green will minimize experimental systematic errorThis indicator provides an end point that is similar in pH to the equivalence point.	1-3
Total	4

Bromocresol green.

A change in colour of the indicator (end point) indicates that the equivalence point has been reached. The appropriate indicator changes colour in the pH range of the equivalence point. Bromocresol green is the only indicator that changes colour in the pH range of the equivalence point. Other indicators will change colour before or after the equivalence point, resulting in an incorrect titre.

Question 27**(12 marks)**

Phosphoric acid, H_3PO_4 (aq), is a weak, triprotic acid.

- (a) Write the ionisation equation for phosphoric acid in water which shows the **second** proton of the acid being released into solution. (2 marks)

Magnesium carbonate, MgCO_3 (s), is an ingredient of a commonly-used antacid.

- (b) Other than water, list **three** species (elements, compounds, ions) that would be found in the reacting vessel open to the atmosphere at the completion of the reaction between excess solid magnesium carbonate and an aqueous solution of phosphoric acid. (3 marks)

One: _____

Two: _____

Three: _____

Sodium hydroxide solution, NaOH (aq), was used in a titration to determine the concentration of phosphoric acid.

- (c) Other than it having too low a molar mass, state **two** reasons why the concentration of the sodium hydroxide solution cannot be reliably determined by weighing out an amount of solid sodium hydroxide and dissolving it in a known volume of distilled water. (2 marks)

One: _____

Two: _____

Question 27**(12 marks)**

- (a) Write the ionisation equation for phosphoric acid in water which shows the **second** proton of the acid being released into solution. (2 marks)

Description	Marks
$\text{H}_2\text{PO}_4^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{HPO}_4^{2-}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$ (one minor error)	2 (1)
	Total 2
Note: State symbols not required Note: Minor errors could include no double arrow, one charge missing	

- (b) Other than water, list **three** species (elements, compounds, ions) that would be found in the reacting vessel open to the atmosphere at the completion of the reaction between excess solid magnesium carbonate and an aqueous solution of phosphoric acid. (3 marks)

Description	Marks
Only the first three species listed are marked.	
1. magnesium carbonate, $\text{MgCO}_3(\text{s})$ 2. magnesium phosphate, $\text{Mg}_3(\text{PO}_4)_2(\text{s})$ 3. $\text{H}_2\text{CO}_3(\text{aq})$ (from the reaction of CO_2 with H_2O) 4. $\text{HCO}_3^-(\text{aq})$ (from the reaction of CO_2 with H_2O) 5. $\text{CO}_3^{2-}(\text{aq})$ (from the reaction of CO_2 with H_2O) 6. $\text{H}_3\text{O}^+(\text{aq})$ ions (from the self-ionisation of water) 7. OH^- (from the self-ionisation of water) 8. CO_2 9. $\text{Mg}^{2+}/\text{PO}_4^{3-}/\text{HPO}_4^{2-}/\text{H}_2\text{PO}_4^-/\text{O}_2$	1–3
	Total 3
Note: the equation for the reaction is: $3 \text{MgCO}_3(\text{s}) + 2 \text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{Mg}_3(\text{PO}_4)_2(\text{s}) + 3 \text{H}_2\text{O}(\ell) + 3 \text{CO}_2(\text{g})$ MgCO_3 is in excess so there is no H_3PO_4 left A small amount of carbon dioxide gas, $\text{CO}_2(\text{g})$, will dissolve and react with water, H_2O , to form H_2CO_3 (which will form $\text{HCO}_3^-(\text{aq})$ and $\text{CO}_3^{2-}(\text{aq})$ ions). Note: Only mark the first three listed	

- (c) Other than it having too low a molar mass, state **two** reasons why the concentration of the sodium hydroxide solution cannot be reliably determined by weighing out an amount of solid sodium hydroxide and dissolving it in a known volume of distilled water?
(2 marks)

Description	Marks
Any two of the following: <ul style="list-style-type: none"> • hygroscopic – it absorbs atmospheric moisture which changes the mass of the NaOH sample. • deliquescent – it dissolves in the water it absorbs, so no mass of solid can be weighed • absorbs atmospheric carbon dioxide gas, CO₂(g) to form sodium carbonate, Na₂CO₃(s), and water, H₂O(l). • Not known purity (% purity not known) 	1–2
Total	2
Note: Do not accept 'not available in pure form'	

- (d) Select the acid-base indicator from the table above that would be most suitable for the titration between phosphoric acid, H₃PO₄(aq), and sodium hydroxide solution, NaOH(aq). Justify your choice of indicator, including **one** relevant equation.
(5 marks)

Description	Marks
Phenolphthalein	
Recognition that PO ₄ ³⁻ present in solution at equivalence point. $(3 \text{ OH}^-(\text{aq}) + \text{H}_3\text{PO}_4(\text{aq}) \rightarrow \text{PO}_4^{3-}(\text{aq}) + 3 \text{H}_2\text{O}(l))$	
The solution at the equivalence point will be (slightly) basic (with a pH of approximately 9) due to the excess of hydroxide ions	
The phosphate ion undergoes hydrolysis to form hydroxide ions. $\text{PO}_4^{3-}(\text{aq}) + \text{H}_2\text{O}(l) \rightleftharpoons \text{HPO}_4^{2-}(\text{aq}) + \text{OH}^-(\text{aq})$	1–5
The pH at which the indicator changes colour approximates the pH of the equivalence point	
The Na ⁺ ion does not hydrolyse/has no effect on the pH of the solution/is a neutral ion	
Total	5
Note: candidate must give five valid points for full marks, and must include an equation.	
Note: no hydrolysis equation, maximum of four marks	
Note: If phenol red with explanation above, four maximum	
Alternative responses that candidate may provide: Methyl Orange The pH of the first equivalence point is around 4.7. If candidate identify this and supplied appropriate logic with equation, up to full marks may be awarded.	
If a candidate recognises that the third equivalence point is beyond the end point of phenolphthalein and explains why none of the indicators would be appropriate with sufficient reasoning, up to full marks may be awarded.	

Question 30**(7 marks)**

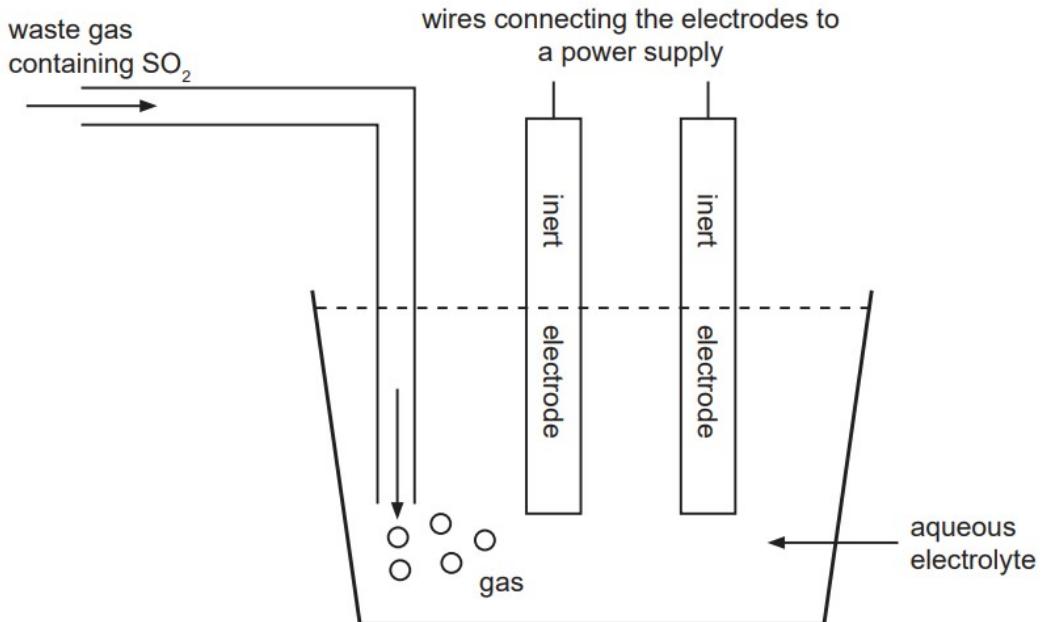
Sulfur dioxide must be removed from waste industrial gases before they are released into the atmosphere. One method of doing this is the electrolytic conversion of sulfur dioxide into dithionite ($\text{S}_2\text{O}_6^{2-}$):



- (a) Identify the atom that is oxidised and the atom that is reduced in this reaction. (2 marks)

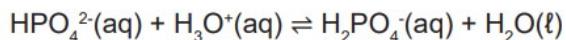
Atom that is oxidised	
Atom that is reduced	

An electrolytic cell, similar to the simplified one shown below, can be used for the above process.



A chemist, who was investigating this process, used 1.00 mol L^{-1} sodium perchlorate (NaClO_4) solution as the electrolyte. The chemist found that the pH of this electrolyte steadily decreased as more SO_2 -containing waste gas was treated. The final pH was 2.42.

The observed pH change prompted the chemist to change the electrolyte to a mixture of potassium hydrogen phosphate (K_2HPO_4) and potassium dihydrogenphosphate (KH_2PO_4), in which the following equilibrium occurred:



No significant pH changes occurred when this new electrolyte was used.

- (b) Explain how the $\text{HPO}_4^{2-}/\text{H}_2\text{PO}_4^-$ prevented any significant pH change when the SO_2 was bubbled into the solution. (5 marks)

Question 30 (7 marks)

- (a) Identify the atom that is oxidised and the atom that is reduced in this reaction. (2 marks)

Description	Marks
The atom that is oxidised is sulfur (or S).	1
The atom that is reduced is hydrogen (or H).	1
Total	2

Note:

- Must have the actual atom. No marks allocated for SO_2 being oxidised or H_2O being reduced.
- H^+ and S^{2-} are not acceptable answers.

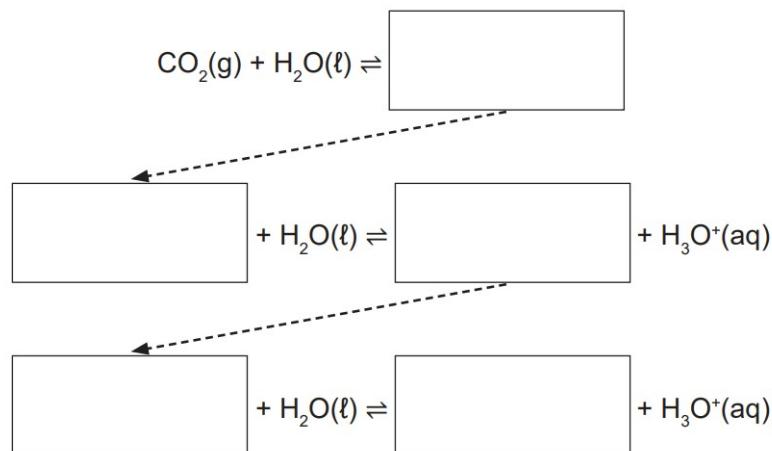
- (b) Explain how the $\text{HPO}_4^{2-}/\text{H}_2\text{PO}_4^-$ prevented any significant pH change when the SO_2 was bubbled into the solution. (5 marks)

Description	Marks
Recognition that SO_2 reaction results in increase in the $[\text{H}^+]$	1
Recognition $\text{HPO}_4^{2-}/\text{H}_2\text{PO}_4^-$ is a buffer (because it is a weak base/weak acid combination)	1
Produced H^+ reacts with the HPO_4^{2-} base in the buffer.	1
Recognition that this consumes the majority of the added H^+ in the solution (therefore overall minimal increase in $[\text{H}^+]$)	1
Recognition of how $[\text{H}^+]$ links to pH	1
Total	5

Question 31**(9 marks)**

The amount of carbon dioxide in the Earth's atmosphere is increasing, leading to more carbon dioxide dissolving in the oceans and hence ocean acidification.

- (a) Complete the following sequence of equations to show what happens to carbon dioxide when it dissolves in water. (3 marks)



- (b) Other than death, state **two** consequences of the above sequence of equations on marine organisms with shells. (2 marks)

One: _____

Two: _____

- (c) Use Le Châtelier's Principle and the sequence of equations in part (a) to predict what might happen, in relation to ocean acidification, if the United Nations Kyoto Protocol is discarded. Explain your reasoning. (4 marks)

Question 31**(9 marks)**

- (a) Complete the following sequence of equations to show what happens to carbon dioxide when it dissolves in water. (3 marks)

Description	Marks
$\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell) \rightleftharpoons$ H₂CO₃(aq)	1
$\text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons$ HCO₃⁻(aq) + H ₃ O ⁺ (aq)	1
$\text{HCO}_3^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons$ CO₃²⁻(aq) + H ₃ O ⁺ (aq)	1
Total	3

Note:

- Only allocate marks for the species in the boxes shown above.
- State symbols are desirable but not essential.

- (b) Other than death, state **two** consequences of the above sequence of equations on marine organisms with shells. (2 marks)

Description	Marks
States a plausible consequence	1
States another (different) plausible consequence	1
Total	2
Answers could include:	
<ul style="list-style-type: none"> • building shells becomes harder • weaker/thinner shells/poor quality shells • makes organisms more vulnerable to predators • less likely to grow large • less likely to reproduce. 	
Note:	
<ul style="list-style-type: none"> • Do not accept answers that refer to coral reefs or habitats. 	

- (c) Use Le Châtelier's Principle and the sequence of equations in part (a) to predict what might happen, in relation to ocean acidification, if the United Nations Kyoto Protocol is discarded. Explain your reasoning. (4 marks)

Description	Marks
If the Kyoto Protocol is discarded then it is likely that there will be increased CO ₂ emissions into the atmosphere that lead to increased amounts of CO ₂ dissolved in the oceans.	1
Therefore, the forward reactions in the above sequence of reactions will be favoured.	1
The result is an increase in the concentration of H ₃ O ⁺ ions.	1
The prediction is, therefore, that the pH of oceans will decrease even more.	1
Total	4

Note:

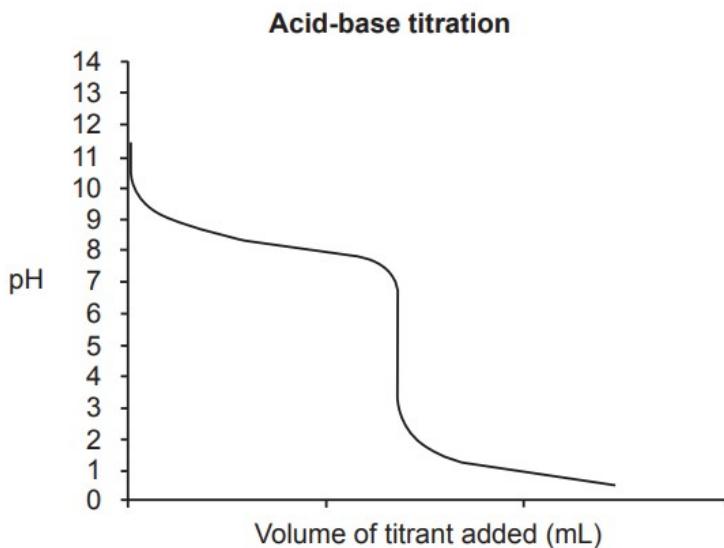
- Accept oceans become more acidic.

2017 Creelman:

Question 34**(6 marks)**

The data below were collected from an acid-base titration.

- (a) Label the equivalence point on the titration curve below using an arrow and record the pH value at this point. (2 marks)



pH value at equivalence point: _____

- (b) Select an indicator from the table below that would be **best** for this titration and justify your choice. (4 marks)

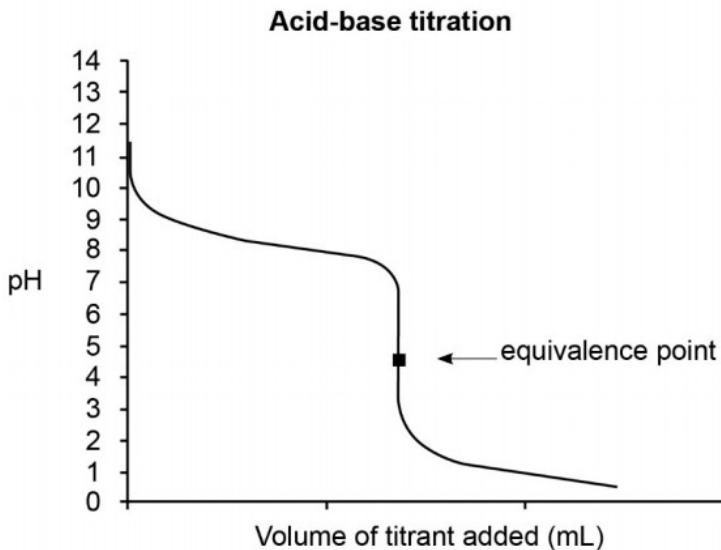
Indicator	Low pH colour	Transition pH range	High pH colour
Methyl Yellow	red	2.1 – 3.3	yellow
Bromocresol Green	yellow	3.8 – 5.4	blue
Bromothymol Blue	yellow	6.0 – 7.6	blue
Phenolphthalein	colourless	8.3 – 10.0	pink
Alzarine Yellow R	yellow	10.2 – 12.0	red

Indicator: _____

Justification: _____

Question 34**(6 marks)**

- (a) Label the equivalence point on the titration curve below using an arrow and record the pH value at his point. (2 marks)



Description	Marks
Label on Titration curve pointing to correct equivalence point pH 4.8 ± 0.5	1
pH stated within range of 4.8 ± 0.5 (pH 4.3 to 5.3)	1
Total	2

- (b) Select an indicator from the table below that would be **best** for this titration and justify your choice. (4 marks)

Description	Marks
Bromocresol Green	1
Justification may include up to, but not be limited to, three of the following recognitions:	
<ul style="list-style-type: none"> That a change in colour of the indicator (end point) indicates that the equivalence point has been reached The (appropriate) indicator changes colour in the range of the equivalence point Bromocresol Green is the only indicator that changes colour in the range of the equivalence point Other indicators will change colour either before (Phenolphthalein, Alizarine Yellow R) or after the equivalence point (Methyl Yellow, Bromocresol Blue) The use of indicators other than Bromocresol Green will produce an incorrect titre (too large or too small) The use of indicators other than Bromocresol Green will result in a systematic error or the use of Bromocresol Green will minimize experimental systematic error This indicator provides an end point that is similar in pH to the equivalence point. 	1-3
Total	4

Question 40**(15 marks)**

Hydrogen fluoride, HF, is a highly dangerous and corrosive liquid that boils at near room temperature. It readily forms hydrofluoric acid in the presence of water and is an ingredient used to produce many important compounds, including medicines and polymers.

- (a) (i) Name the electrostatic attractive force that holds the hydrogen and fluorine atoms together **within** hydrogen fluoride molecules. (1 mark)

- (ii) Name the electrostatic attractive force **between** the hydrogen fluoride molecules. (1 mark)

- (iii) Explain the origin of the attractive force **between** the hydrogen fluoride molecules. (2 marks)

- (b) The equilibrium constant (K) for the dissociation of hydrofluoric acid is 6.8×10^{-4} , and for hydrochloric acid K is very large. To make a solution of hydrofluoric acid with the same pH as hydrochloric acid, a greater concentration of hydrofluoric acid is required. Explain why this is so. (3 marks)

- (c) The salts, sodium chloride and sodium fluoride, readily dissolve in water. At 25.0 °C the pH of the sodium chloride solution is equal to 7 whereas the pH of the sodium fluoride solution is greater than 7. Explain this difference in pH. Include any relevant equation(s) to support your answer. (3 marks)

Propanoic acid, $\text{CH}_3\text{CH}_2\text{COOH}$, is also a weak monoprotic acid. When 0.500 mol of sodium propanoate is dissolved in 1.00 L of 0.500 mol L^{-1} propanoic acid at 25.0 °C a buffer solution is formed.

- (d) (i) Addition of 10.0 mL of 1.00 mol L^{-1} HCl(aq) to this buffer does not significantly change its pH. Explain this observation, including any relevant equation(s). (3 marks)

- (ii) State **two** conditions required to ensure that this system has a high buffering capacity. (2 marks)

One: _____

Two: _____

- (b) The equilibrium constant (K) for the dissociation of hydrofluoric acid is 6.8×10^{-4} , and for hydrochloric acid K is very large. To make a solution of hydrofluoric acid with the same pH as hydrochloric acid, a greater concentration of hydrofluoric acid is required. Explain why this is so. (3 marks)

Description	Marks
Both acids will have to have the same $[H^+]$ to have the same pH	1
HF does not ionise to the same extent as HCl	1
Greater concentration of HF is needed to give the required $[H^+]$	1
Total	3

- (c) The salts, sodium chloride and sodium fluoride, readily dissolve in water. At 25.0 °C the pH of the sodium chloride solution is equal to 7 whereas the pH of the sodium fluoride solution is greater than 7. Explain this difference in pH. Include any relevant equation(s) to support your answer. (3 marks)

Description	Marks
Recognition that fluorine hydrolyses resulting in the formation of hydroxide ions results in a solution with a pH > 7	1
$F^-(aq) + H_2O(l) \rightleftharpoons HF(aq) + OH^-(aq)$	1
The chloride ion (is the very weak conjugate base of a strong acid and hence) cannot hydrolyse and/or is a neutral ion	1
Total	3

Note: A comment regarding Na^+ is not required for the mark.

- (d) (i) Addition of 10.00 mL of 1.00 mol L⁻¹ HCl(aq) to this buffer does not significantly change its pH. Explain this observation, including any relevant equation(s). (3 marks)

Description	Marks
Equation for the hydrolysis of propanoic acid: $CH_3CH_2COOH + H_2O \rightleftharpoons CH_3CH_2COO^-(aq) + H_3O^+(aq)$	1
As hydrogen ions are added to the buffer equilibrium, the equilibrium will shift left to use up the added H^+ ions	1
Due to the shift in equilibrium, the overall concentration of hydrogen ion is minimised (and so the pH change is insignificant)	1
Total	3

- (ii) State **two** conditions required to ensure this system has a high buffering capacity. (2 marks)

Description	Marks
Equal concentrations (of acid and conjugate base)	1
High concentrations (of acid and conjugate base)	1
Total	2

Question 31

(7 marks)

An aqueous solution is prepared that contains 0.1 mol L^{-1} Na^+ and 0.1 mol L^{-1} HC_2O_4^- .

- (a) Write the **two** possible reactions for the hydrolysis of the HC_2O_4^- ion. (3 marks)

One: _____

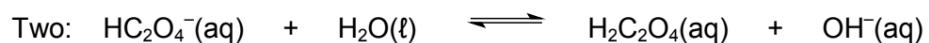
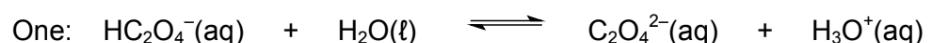
Two: _____

- (b) The pH of the solution was measured and found to be less than seven. Based on this observation, state which of the hydrolysis equations has the higher equilibrium constant. Use your understanding of equilibrium concepts to explain your choice fully. (4 marks)

Question 31**(7 marks)**

An aqueous solution is prepared that contains $0.1 \text{ mol L}^{-1} \text{ Na}^+$ and $0.1 \text{ mol L}^{-1} \text{ HC}_2\text{O}_4^-$.

- (a) Write the **two** possible reactions for the hydrolysis of the HC_2O_4^- ion. (3 marks)



Description	Marks
$\text{HC}_2\text{O}_4^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{C}_2\text{O}_4^{2-}(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$	1
$\text{HC}_2\text{O}_4^-(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_2\text{C}_2\text{O}_4(\text{aq}) + \text{OH}^-(\text{aq})$	1
Use of double arrows for equilibria	1
Total	3

N.B. State symbols not required for full marks.

Award 1 mark if double arrows only used once.

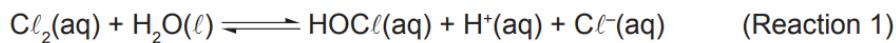
- (b) The pH of the solution was measured and found to be less than 7. Based on this observation, state which of the hydrolysis equations has the higher equilibrium constant. Use your understanding of equilibrium concepts to explain your choice fully. (4 marks)

Description	Marks
If solution has $\text{pH} < 7$, the concentration of $[\text{H}_3\text{O}^+] > [\text{OH}^-]$	1
K is ratio of products to reactants	1
H_3O^+ producing equation has the higher K value	1
Thus, H_3O^+ producing equation moves forward to a greater extent than the OH^- producing equation	1
Total	4

Question 42

(12 marks)

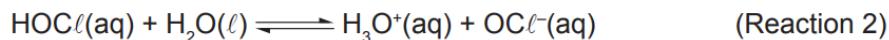
Large public swimming pools are often chlorinated using chlorine gas. The gas is bubbled through the water forming the equilibrium reaction shown below:



The equilibrium constant for this reaction at 25.0°C is 3.94×10^4 .

- (a) Compare the relative amounts of chlorine and hypochlorous acid at equilibrium at 25°C. (1 mark)

The hypochlorous acid can dissociate as shown in the equilibrium below to give hypochlorite ion.



- (b) The pH of swimming pools is kept at approximately 7.5. A reason for this is to maximise the concentration of hypochlorous acid, the most effective disinfectant form of chlorine in water. Explain, using the appropriate chemistry concepts, why a pH of about 7.5 will maximise hypochlorous acid concentration. Your explanation should consider equilibrium Reactions 1 and 2. (3 marks)

Question 42**(12 marks)**

Large public swimming pools are often chlorinated using chlorine gas. The gas is bubbled through the water forming the equilibrium reaction shown below.

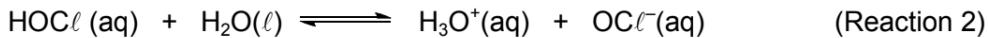


The equilibrium constant for this reaction, at 25°C, is 3.94×10^4 .

- (a) What can be said about the relative amounts of chlorine and hypochlorous acid at equilibrium at 25°C? (1 mark)

Description	Marks
The quantity of hypochlorous acid is much greater than chlorine	1
Total	1

The hypochlorous acid can dissociate as shown in the equilibrium below to give hypochlorite ion.



- (b) The pH of swimming pools is kept at approximately 7.5. A reason for this is to maximise the concentration of the hypochlorous acid, most effective disinfectant form of chlorine in water. Explain, using the appropriate chemistry concepts, why a pH of about 7.5 will maximise hypochlorous acid concentration. Your explanation should consider equilibrium Reactions 1 and 2. (3 marks)

Description	Marks
At low pH (less than 7), in equilibrium (1) the reverse reaction is favoured which results in less hypochlorous acid.	1
For equilibrium (2) a pH less than 7 will favour the reverse reaction which leads to more hypochlorous acid	1
Thus as a compromise a pH slightly above 7 is used. At this pH, equilibrium (1) moves forward whilst for equilibrium (2) the forward reaction will not be strongly favoured thus maximising production of hypochlorous acid	1
Total	3

For equilibrium (1) at a high pH the forward reaction is favoured resulting in more hypochlorous acid. (1 mark)

Question 36

(6 marks)

Water is able to react with itself in the process known as 'self-ionisation' or 'auto-ionisation'.

- (a) Write the equation for the self-ionisation of water. (1 mark)

For more information about the study, please contact Dr. [REDACTED] at [REDACTED].

- (b) At 25°C, the value of K_w is approximately 1.0×10^{-14} . At 10°C, the value of K_w is approximately 2.9×10^{-15} . (2 marks)

What are the relative concentrations of H^+ and OH^- ions in a neutral water solution at 25°C? Circle the correct answer.

$$[\text{H}^+] > [\text{OH}^-]$$

$$[\text{H}^+] < [\text{OH}^-]$$

$$[\text{H}^+] = [\text{OH}^-]$$

What are the relative concentrations of H^+ and OH^- ions in a neutral water solution at 10°C? Circle the correct answer.

$$[\text{H}^+] > [\text{OH}^-]$$

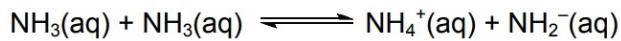
$$[\text{H}^+] < [\text{OH}^-]$$

$$[\text{H}^+] = [\text{OH}^-]$$

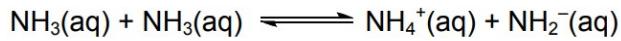
- (c) Consider the values of K_w at 10°C and 25°C, and state whether the self-ionisation of water is an endothermic or exothermic process. Give a reason to support your answer. (3 marks)

Question 28**(4 marks)**

Like water, ammonia is able to react with itself, in the process known as 'self-ionisation'. The equation for the self-ionisation of ammonia is below.



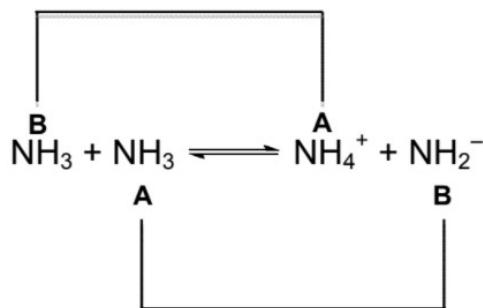
- (a) Identify the conjugate acid and base pairs in the reaction. Join each pair with a line, and label the conjugate acid and base of each pair appropriately. (1 mark)



- (b) At standard temperature and pressure, the equilibrium constant, K, for this reaction is about 1×10^{-30} . The self-ionisation of ammonia is an endothermic process. Will the value of K be less than or greater than 1×10^{-30} at temperatures greater than 0°C? Explain. (3 marks)

Question 28**(4 marks)**

- (a) Identify the conjugate acid and base pairs in the reaction. Join each pair with a line, and label the conjugate acid and base of each pair appropriately. (1 mark)



A = conjugate acid
B = conjugate base

Description	Marks
Correct answer	1
Question incorrectly answered or not attempted.	0
Total	1

Students can use the term conjugate acid/base or CA/CB

- (b) At standard temperature and pressure, the equilibrium constant, K, for this reaction is about 1×10^{-30} . The self-ionisation of ammonia is an endothermic process. Will the value of K be less than or greater than 1×10^{-30} at temperatures greater than 0°C? Explain. (3 marks)

The reaction is endothermic. At T > 0 °C (the temperature for which the equilibrium constant is given), the forward reaction will be favoured. This will increase the concentration of products relative to reactants, meaning that the value of K will increase. i.e. the value of K will be greater than 1×10^{-30} at temperatures greater than 0 °C.

Description	Marks
Answer as above. Important points are (i) forward reaction is favoured , meaning (ii) concentration of products increased [relative to reactants] , therefore (iii) K > 1×10^{-30} or greater or larger	1-3
Question incorrectly answered or not attempted.	0
Total	3