ACIDS & BASES

PAST EXAM QUESTIONS

Section 1: Brønsted-Lowry theory, conjugate acids and bases

WACE 2016 Sample Exam Q5: ***Originally from WACE 3AB 2012***

Consider the following equation:

$$HS^-(aq) + CO_3^{2-}(aq) \rightleftharpoons S^{2-}(aq) + HCO_3^-(aq)$$

Which one of the following is **not** true of this equation?

(a) HCO_3^- is acting as a Brønsted-Lowry acid
(b) CO_3^{2-} is acting as a conjugate base
(c) HS^- is acting as a conjugate base
(d) S^{2-} is acting as a Brønsted-Lowry base

WACE 2016 Sample Exam Q4:

Which one of the following reactions does **not** represent the Brønsted-Lowry model?

(a) $HSO_4^-(aq) + H_2O(\ell) \rightarrow SO_4^{2-}(aq) + H_3O^+(aq)$
(b) $CH_3COOH(aq) + NaOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(\ell)$
(c) $HCO_3^-(aq) + H_2O(\ell) \rightarrow H_2CO_3(aq) + OH^-(aq)$
(d) $CaCO_3(s) + 2\ HC\ell(aq) \rightarrow CaC\ell_2(aq) + CO_2(g) + 2\ H_2O(\ell)$

WACE 2AB 2014 Q17:

In which one of the following equations is the underlined species acting as a Brønsted-Lowry base?

(a) $\underline{NH_3} + CH_3COOH \rightleftharpoons NH_4^+ + CH_3COO^-$
(b) $HCO_3^- + \underline{H_2O} \rightleftharpoons H_2CO_3 + OH^-$
(c) $HPO_4^{2-} + \underline{CH_3COOH} \rightleftharpoons H_2PO_4^- + CH_3COO^-$
(d) $\underline{NH_4^+} + OH^- \rightleftharpoons NH_3 + H_2O$

WACE 2AB 2012 Q21:

Which equation below **best** represents the $H_2PO_4^-$ ion acting as a Brønsted-Lowry acid?

(a) $H_2PO_4^-(aq) \rightarrow 2\ H^+(aq) + PO_4^{3-}(aq)$
(b) $H_2PO_4^-(aq) + OH^-(aq) \rightarrow H_3PO_4(aq) + H_2O(\ell)$
(c) $H_2PO_4^-(aq) + H_2O(\ell) \rightarrow H_3PO_4(aq) + OH^-(aq)$
(d) $H_2PO_4^-(aq) + H_2O(\ell) \rightarrow HPO_4^{2-}(aq) + H_3O^+(aq)$

Section 1: Brønsted-Lowry theory, conjugate acids and bases

WACE 2AB 2011 Q13:

Which of the following species is the conjugate acid of the CO_3^{2-} ion?

(a) HCO_3^-
(b) H_2CO_3
(c) H_3O^+
(d) CO_2

WACE 2AB 2010 Q16:

Which one of the following is the conjugate base of HS^-?

(a) S^{2-}
(b) OH^-
(c) H^+
(d) H_2S

WACE 3AB 2015 Q18:

The reaction equilibrium between hydrogencarbonate ion and dihydrogen sulfide is represented by the equation shown below.

$$HCO_3^-(aq) + H_2S(aq) \rightleftharpoons H_2CO_3(aq) + HS^-(aq)$$

According to the Brønsted-Lowry theory of acids and bases, which one of the following shows the two species acting as bases in this equilibrium system?

(a) HCO_3^- and H_2CO_3
(b) H_2S and HS^-
(c) H_2S and H_2CO_3
(d) HCO_3^- and HS^-

WACE 3AB 2014 Q15:

Consider the following reaction.

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

Which one of the following represents an acid-base conjugate pair for this reaction?

(a) OBr^- / H_2O
(b) $HOBr$ / OH^-
(c) OBr^- / OH^-
(d) H_2O / OH^-

Section 1: Brønsted-Lowry theory, conjugate acids and bases

WACE 3AB 2013 Q15:

Which one of the following substances can behave as a Brønsted-Lowry acid or base?

(a) H_2O_2
(b) NH_4^+
(c) CH_3NH_2
(d) $H_2PO_4^-$

WACE 3AB 2011 Q18:

In which one of the following is the reactant in **bold** reacting as an acid?

(a)	**2 Na(s)**	+	2 H_2O	→	2 NaOH	+ H_2
(b)	**NH_3**	+	H_2O	→	NH_4^+	+ OH^-
(c)	**$Fe(H_2O)_6^{3+}$**	**+**	**H_2O**	**→**	**$Fe(H_2O)_5(OH)^{2+}$**	**+ H_3O^+**
(d)	**CO_2**	+	H_2O	→	H_2CO_3	

WACE 3AB 2010 Q8:

Which one of the following species **cannot** act as a Brønsted-Lowry acid and a Brønsted-Lowry base?

(a) $H_2PO_4^-$
(b) CH_3COCH_3
(c) H_2O
(d) HCO_3^-

TEE 2009 Q15:

In which of the following is water acting as a base?

(a)	$H_2O(\ell)$	+	$NH_3(aq)$	→	$NH_4^+(aq)$	+ $OH^-(aq)$
(b)	$H_2O(\ell)$	+	$CO_2(g)$	→	$H_2CO_3(aq)$	
(c)	**$H_2O(\ell)$**	**+**	**$HSO_4^-(aq)$**	**→**	**$H_3O^+(aq)$**	**+ $SO_4^{2-}(aq)$**
(d)	2 $H_2O(\ell)$	+	2 Na(s)	→	2 NaOH(aq)	+ $H_2(g)$

TEE 2008 Q13:

In which of the following reactions is the underlined species acting as a base?

(a)	**$\underline{CH_3NH_2}$**	**+**	**CH_3CH_2COOH**	**⇌**	**$CH_3NH_3^+$**	**+ $CH_3CH_2COO^-$**
(b)	$\underline{NH_4^+}$	+	SO_4^{2-}	⇌	NH_3	+ HSO_4^-
(c)	$\underline{NH_3}$	+	H_2O	⇌	NH_2^-	+ H_3O^+
(d)	2 CrO_4^{2-}	+	2 $\underline{HSO_4^-}$	⇌	$Cr_2O_7^{2-}$	+ H_2O + SO_4^{2-}

Section 1: Brønsted-Lowry theory, conjugate acids and bases

TEE 2007 Q10:

Consider the following acid-base reaction:

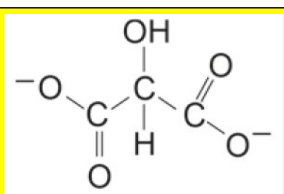
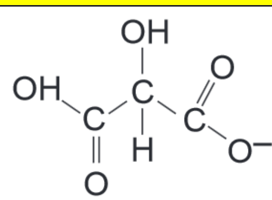
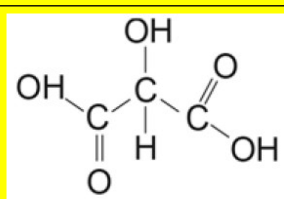
$$HSO_4^- \ + \ HS^- \ \rightleftharpoons \ SO_4^{2-} \ + \ H_2S$$

Which one of the following correctly identifies the acid-base conjugate pairs in this system?

	Acid	Conjugate base	Base	Conjugate acid
(a)	HSO_4^-	HS^-	SO_4^{2-}	H_2S
(b)	**HSO_4^-**	**SO_4^{2-}**	**HS^-**	**H_2S**
(c)	HSO_4^-	H_2S	HS^-	SO_4^{2-}
(d)	HS^-	HSO_4^-	H_2S	SO_4^{2-}

WACE 2016 Sample Exam Q27: ***Originally from WACE 2011***

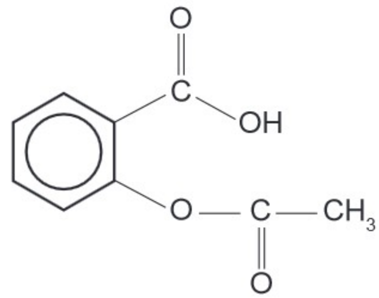
(a) Complete the table by writing the formula or drawing the structure for the conjugate base, species X or conjugate acid in each blank space as appropriate. Species X is the species that is able to form both a conjugate base and a conjugate acid. (6 marks)

Conjugate base	Species X	Conjugate acid
CH_3NH^-	**CH_3NH_2**	$CH_3NH_3^+$
$C_2O_4^{2-}$	**$HC_2O_4^-$**	**$H_2C_2O_4$**
		

Examiner's comments: The question was not particularly well done. Candidates apparently have difficulty drawing structures for conjugate acids/bases of what might be considered more 'complex' molecules, and identifying the acidic protons (hydrogens) in such structures. Candidates should be encouraged to draw structural diagrams where necessary.

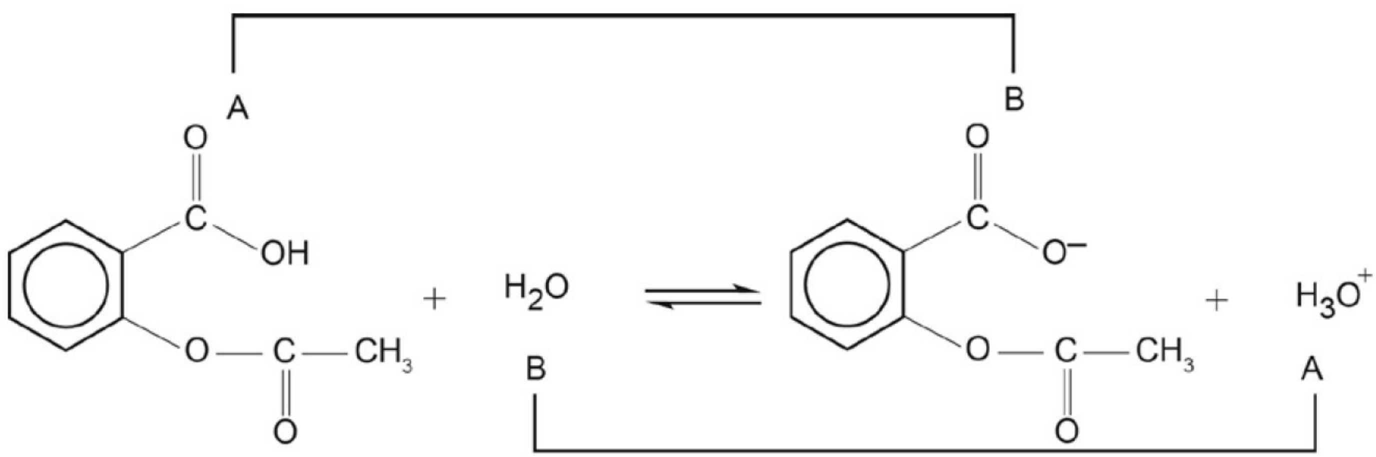
WACE 2016 Sample Exam Q28: ***Originally from WACE 2012***

The active ingredient in aspirin tablets (acetylsalicylic acid) has the structure shown below:



When acetylsalicylic acid is placed in water, some of it dissolves and ionises to form its conjugate base.

(a) Write the equation for the ionisation of acetylsalicylic acid in the space below, and identify the conjugate acid and base pairs in the reaction. Connect each acid-base pair with a line, and label the conjugate acid in the pair 'A', and the conjugate base 'B'. (3 marks)



1 mark: Correct ionisation reaction

1 mark: Correct connection of pairs

1 mark: Acids and bases labelled correctly

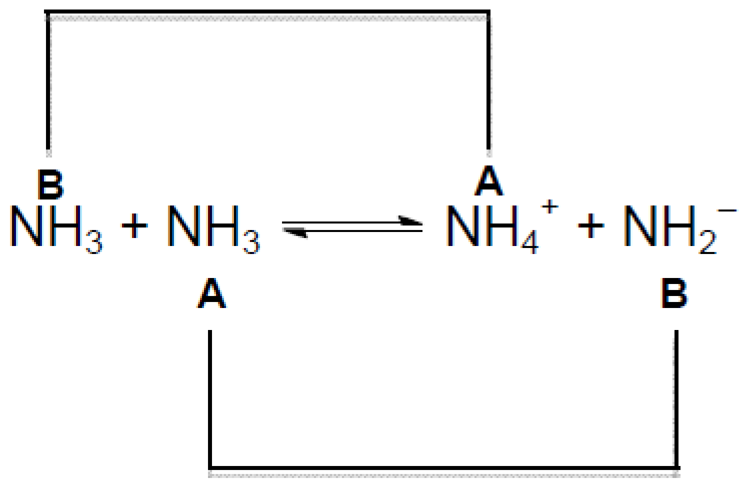
Examiner's comments: Most candidates were not able to rewrite correctly the complete structure of the aspirin (and were not penalized for this). Candidates should be discouraged from converting structures to molecular formulas as too many candidates did this incorrectly. A significant number of candidates also did not follow the instructions to label both acid-base pairs and connect them both with lines.

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.

$$NH_3(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + NH_2^-(aq)$$

(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)



Examiner’s comments: This question was well done. A common error was to not label *both* the conjugate base and acid in the pair, or to label the *process* as a conjugate acid or base. This suggests that candidates did not read the question carefully, ignored the instruction, or do not understand that a conjugate base must have a conjugate acid, and vice versa.

(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)

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.

1 mark: *forward reaction is favoured*
1 mark: *concentration of products increases (relative to reactants)*
1 mark: *K > 1×10^{-30} at temperatures greater than 0 °C*

Section 2: Polyprotic acids, acid strength, hydrolysis of salts

WACE 2AB 2014 Q20:

A few drops of water are added to one litre of pure nitric acid. Which one of the following **best** describes the resulting solution?

(a) A concentrated solution of a strong acid
(b) An acidic solution with a pH greater than 7
(c) A dilute solution of a weak acid
(d) A concentrated solution of a weak acid

WACE 2AB 2013 Q6:

Which one of the following lists the solutions, all 0.1 mol L^{-1}, in order of increasing electrical conductivity?

(a) $HNO_3(aq)$	$CH_3COOH(aq)$	$H_2SO_4(aq)$
(b) $CH_3COOH(aq)$	**$NaCℓ(aq)$**	**$MgCℓ_2(aq)$**
(c) $H_2SO_4(aq)$	$CH_3COOH(aq)$	$HNO_3(aq)$
(d) $MgF_2(aq)$	$KI(aq)$	$HCℓ(aq)$

WACE 2AB 2011 Q14:

Which one of the 0.02 mol L^{-1} solutions below will have the **highest** pH?

(a) $HCℓ$
(b) H_2SO_4
(c) HNO_3
(d) CH_3COOH

WACE 2AB 2010 Q15:

In which list are the following 1.00 mol L^{-1} solutions correctly arranged in order of decreasing pH:
calcium hydroxide, nitric acid, acetic (ethanoic) acid, sulfuric acid, ammonia and sodium chloride?

(a) $Ca(OH)_2$	**>**	**NH_3**	**>**	**$NaCℓ$**	**>**	**CH_3COOH**	**>**	**HNO_3**	**>**	**H_2SO_4**
(b) NH_3	>	$Ca(OH)_2$	>	$NaCℓ$	>	HNO_3	>	CH_3COOH	>	H_2SO_4
(c) HNO_3	>	H_2SO_4	>	CH_3COOH	>	$NaCℓ$	>	NH_3	>	$Ca(OH)_2$
(d) H_2SO_4	>	HNO_3	>	CH_3COOH	>	$NaCℓ$	>	NH_3	>	$Ca(OH)_2$

WACE 2016 Sample Exam Q7: ***Originally from WACE 3AB 2011***

Which one of the following describes the acidity/basicity of a solution of the following compounds when dissolved in distilled water?

	Ammonium chloride	**Potassium carbonate**	**Sodium nitrate**	**Sodium ethanoate**
(a)	**acidic**	**basic**	**neutral**	**basic**
(b)	acidic	basic	acidic	basic
(c)	basic	acidic	neutral	acidic
(d)	Basic	basic	basic	acidic

WACE 3AB 2015 Q19:

The following 1.00 mol L^{-1} solutions are diluted by the addition of water. In which solution will the pH **not** change but the electrical conductivity will decrease?

(a) sodium carbonate
(b) ammonium chloride
(c) sodium chloride
(d) ethanoic (acetic) acid

WACE 3AB 2014 Q14:

Which of the following 0.1 mol L^{-1} aqueous solutions has the highest pH?

(a) Ammonium hydrogensulfate
(b) Hydrochloric acid
(c) Potassium phosphate
(d) Sodium nitrate

WACE 3AB 2013 Q16:

A solution of hydrochloric acid conducts an electric current more readily than an equimolar solution of acetic acid. Which one of the following **best** explains this observation?

(a) Hydrochloric acid is a smaller molecule than acetic acid
(b) Hydrochloric acid is more soluble in water than acetic acid
(c) The equilibrium constant for the ionisation of hydrochloric acid is greater than that for acetic acid
(d) The pH of hydrochloric acid solution is always greater than that for acetic acid solution

WACE 3AB 2013 Q17:

Sodium hydrogensulfate was added to a swimming pool to reduce the pH of the water. Which one of the following equations **best** shows the reaction responsible for this?

(a) $Na^+(aq) + H_2O(\ell) \rightarrow NaOH(aq) + H^+(aq)$
(b) $HSO_4^-(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + SO_4^{2-}(aq)$
(c) $2\ HSO_4^-(aq) + H_2O(\ell) \rightarrow H_3O^+(aq) + H_2SO_4(aq) + SO_4^{2-}(aq)$
(d) $HSO_4^-(aq) + H_2O(\ell) \rightarrow OH^-(aq) + H_2SO_4(aq)$

WACE 3AB 2012 Q8:

Consider the following list of compounds:

(i) KNO_3
(ii) Na_3PO_4
(iii) Na_2S
(iv) $Ba(OH)_2$
(v) $Ca(NO_3)_2$

Which one of the above compounds will dissolve in water to give a basic solution?

(a) (i), (ii), (iii), (iv)
(b) (ii), (iii), (iv), (v)
(c) (ii), (iii), (iv)
(d) (ii), (iii)

WACE 3AB 2012 Q12:

Some solid magnesium carbonate is added to dilute hydrochloric acid. Which one of the following equations best represents the reaction that occurs?

(a) $MgCO_3 + 2\ HC\ell \rightarrow MgC\ell_2 + H_2O + CO_2$
(b) $MgCO_3 + 2\ H^+ \rightarrow Mg^{2+} + H_2O + CO_2$
(c) $CO_3^{2-} + 2\ HC\ell \rightarrow 2\ C\ell^- + H_2O + CO_2$
(d) $CO_3^{2-} + 2\ H^+ \rightarrow H_2O + CO_2$

WACE 3AB 2011 Q16:

Which one of the following is the strongest electrolyte?

(a) $NH_4C\ell$
(b) H_3PO_4
(c) H_2O
(d) CH_3COOH

Section 2: Polyprotic acids, acid strength, hydrolysis of salts

TEE 2008 Q11:

Which one of the following classifications is correct?

	KCℓ	**KCH_3COO**	**$NH_4C\ell$**	**$KHSO_4$**
(a)	**neutral**	**basic**	**acidic**	**acidic**
(b)	neutral	basic	acidic	neutral
(c)	acidic	acidic	basic	basic
(d)	neutral	acidic	basic	acidic

TEE 2006 Q12:

Which one of the following correctly identifies the acidity of the listed salts when dissolved in water?

	Potassium chloride	**Sodium nitrate**	**Ammonium sulfate**	**Sodium carbonate**
(a)	Neutral	Acidic	Acidic	Neutral
(b)	Acidic	Acidic	Basic	Acidic
(c)	**Neutral**	**Neutral**	**Acidic**	**Basic**
(d)	Acidic	Neutral	Neutral	Basic

TEE 2005 Q13:

Which one of the following correctly identifies the acidity, basicity or neutrality of each of the given solutions?

	Sodium hydrogensulfate	**Potassium phosphate**	**Ammonium chloride**	**Magnesium nitrate**
(a)	Acidic	Acidic	Acidic	Basic
(b)	Neutral	Basic	Neutral	Acidic
(c)	**Acidic**	**Basic**	**Acidic**	**Neutral**
(d)	Basic	Neutral	Basic	Neutral

WACE 2AB 2014 Q38:

(b) Write an ionic equation for the reaction between phosphoric acid and barium hydroxide solution. Include state symbols. (3 marks)

$3\ H_3PO_4(aq)\ +\ 3\ Ba^{2+}(aq)\ +\ 6\ OH^-(aq) \rightarrow Ba_3(PO_4)_2(s)\ +\ H_2O(\ell)$

- **1 mark:** Correct formulas
- **1 mark:** Balanced correctly
- **1 mark:** Correct 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' (4 marks)

- **1 mark:** A polyprotic acid is an acid that can done more than one hydrogen ion / proton per molecule
- **1 mark:** Containing three hydrogen atoms available for ionisation, phosphoric acid can undergo three successive ionisation reactions, each releasing one hydrogen ion / proton.
- **1 mark:** Successive ionisation equations:
  - o $H_3PO_4(aq)\ +\ H_2O(\ell) \rightleftharpoons H_2PO_4^-(aq)\ +\ H_3O^+(aq)$
  - o $H_2PO_4^-(aq)\ +\ H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq)\ +\ H_3O^+(aq)$
  - o $HPO_4^{2-}(aq)\ +\ H_2O(\ell) \rightleftharpoons PO_4^{3-}(aq)\ +\ H_3O^+$
- **1 mark:** Overall equation:
  - o $H_3PO_4(aq)\ +\ 3\ H_2O(\ell) \rightleftharpoons PO_4^{3-}(aq)\ +\ 3\ H_3O^+(aq)$

(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. (4 marks)

- **1 mark:** A strong electrolyte, like barium hydroxide, is a solute that completely or almost completely ionises or dissociates in a solution resulting in a completely ionic solution
- **1 mark:** $Ba(OH)_2(s) \rightarrow Ba^{2+}(aq)\ +\ 2\ OH^-(aq)$ 100% conversion
- **1 mark:** A weak electrolyte, like phosphoric acid, is a solute that only ionises or dissociates in a solution to a limited degree, creating a system of both ionic products and the original molecule.
- **1 mark:** $H_3PO_4(aq) \rightleftharpoons PO_4^{3-}(aq)\ +\ 3\ H^+(aq)$ limited conversion

Section 2: Polyprotic acids, acid strength, hydrolysis of salts

WACE 3AB 2011 Q29:

Write a relevant equation or equations to explain each of the observations shown in the table below.
(4 marks)

Observation	Explanatory equation/s
The pH of a $NaHSO_4$ solution is 5	**$HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$**
A solution of $Mg(OH)_2$ is basic	**$Mg(OH)_2(s) \rightarrow Mg^{2+}(aq) + 2\,OH^-(aq)$**
A solution of Na_2HPO_4 is basic, while a solution of KH_2PO_4 is acidic	**$HPO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons H_2PO_4^-(aq) + OH^-(aq)$** **$H_2PO_4^-(aq) + H_2O(\ell) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$**

Examiner's comments: This question was poorly done. Many candidates simply did not write a chemical equation, but rather a mathematical equation of some sort. Where chemical equations were given, they were very often not balanced, or were molecular rather than ionic. Candidates must be encouraged to take care in writing chemical equations through the examination; in this question, although the aim was not to examine directly whether a student can balance an equation, the examiners still expected that the equations be balanced. Full marks were not awarded for incorrectly balanced equations, or equations that were slightly incorrect in some way. Attention to detail is important.

TEE 2006 Q5:

Sodium hydrogencarbonate is often used to increase the pH in swimming pools. Explain, with the aid of suitable equations, how adding sodium hydrogencarbonate affects the pH of the water. (3 marks)

- **1 mark:** Hydrogencarbonate ion can act as a base by accepting protons
- **1 mark:** Undergoes hydrolysis with water to produce hydroxide ions, which raise the pH of the water in the pool
- **1 mark:** $HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

Examiner's comments: This was one of the questions requiring an explanation linking several key concepts, and many candidates struggled to do this.

Section 2: Polyprotic acids, acid strength, hydrolysis of salts

WACE 2AB 2011 Q30:

The poisonous compound oxalic acid ($H_2C_2O_4$) 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)

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

(b) Complete the table below by giving appropriate formulae. (2 marks)

Substance	Example
A polyprotic acid (other than oxalic acid)	**Any polyprotic acid. E.g. H_2SO_4, H_3PO_4**
A monoprotic acid	**Any monoprotic acid. E.g. $HC\ell$, HNO_3**

(c) Write the equations for the successive ionisation of oxalic acid. (2 marks)

- **1 mark:** $H_2C_2O_4 + H_2O \rightleftharpoons HC_2O_4^- + H_3O^+$
- **1 mark:** $HC_2O_4^- + H_2O \rightleftharpoons C_2O_4^{2-} + H_3O^+$

(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. (2 marks)

- **1 mark:** A strong acid will ionise completely, while a weak acid will not ionise completely
- **1 mark:** Given the same concentration of acids, the full ionisation from H_2SO_4 will result in a higher concentration of H^+ (/ H_3O^+) than the partial ionisation of oxalic acid, therefore H_2SO_4 will have a lower pH.

Examiner's comments: Some candidates answered this question very well but many struggled with the concepts of polyprotic acid, and strong and weak acids. Only the better candidates could write successive ionisation equations in part (c), while part (d), about the pH of equimolar strong and weak acid solutions, was poorly answered.

Section 2: Polyprotic acids, acid strength, hydrolysis of salts

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. (2 marks)

- **1 mark:** $2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(\ell)$
- **1 mark:** Reactants: sodium hydroxide, sulfuric acid

Examiner's comments: Answers here were reasonably good, although there were some candidates who failed to write a **balanced** equation, while others were unable to give the formula for the sulfate ion.

(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. (4 marks)

- **Ethanoate discussion (2 marks)**

  When dissolved, sodium ethanoate dissociates to form acetate ions and sodium ions.
  $$CH_3COONa(s) \rightarrow CH_3COO^-(aq) + Na^+(aq)$$

  The ethanoate ion acts as a Brønsted-Lowry base. It can undergo hydrolysis with water to produce hydroxide ions:
  $$CH_3COO^-(aq) + H_2O(\ell) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

  The production of OH^- ions results in a basic pH.

- **Sulfate discussion (2 marks)**

  When Na_2SO_4 is dissolved it also dissociates.
  $$Na_2SO_4(s) \rightarrow 2\ Na^+(aq) + SO_4^{2-}(aq)$$

  Sulfate ions can theoretically undergo hydrolysis with water to produce hydroxide ions:
  $$SO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons HSO_4^-(aq) + OH^-(aq)$$

  In practice, however, the sulfate ion is an extremely weak base, and the above equation has a very low equilibrium constant value. This results in an extremely small $[OH^-]$ being produced, and does not have a noticeable impact on pH.

Examiner's comments: This question was, on the whole, poorly answered. Some common errors included:
- concentrating on the ethanoate ion and ignoring the sulfate ion
- discussing the Brønsted-Lowry theory without explaining why sodium ethanoate is basic
- being unable to write an ionic equation
- failing to explain why sodium sulfate is neutral
- answering in general terms and not correctly identifying the acid/base species
- using word equations rather than ionic equations as specified in the question

Section 2: Polyprotic acids, acid strength, hydrolysis of salts

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. (3 marks)

- **1 mark:** The HPO_4^{2-} ion is a weak base, which will react with water as follows:
- **2 marks:** $HPO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons H_2PO_4^-(aq) + OH^-(aq)$
- The production of OH^- raises the pH of the solution

(b) A 0.1 mol L^{-1} solution of NH_4CH_3COO (ammonium acetate) has a pH of approximately 7. Explain this, using at least two equations. (3 marks)

- **1 mark:** NH_4^+ is a weak acid: $NH_4^+(aq) + H_2O(\ell) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$
- **1 mark:** CH_3COO^- is a weak base: $CH_3COO^-(aq) + H_2O(\ell) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$
- **1 mark:** The acidity of the NH_4^+ ion counters the basicity of the CH_3cOO^- ion (because the K_a of NH_4^+ and the K_b of CH_3COO^- are very similar)

WACE 3AB 2015 Q16:

An aqueous solution at 25.0 °C with a pH less than zero

(a) contains neither $H^+(aq)$ or $OH^-(aq)$ ions
(b) has a very high concentration of $H^+(aq)$ ions
(c) contains no $OH^-(aq)$ ions
(d) contains an equal concentration of $H^+(aq)$ and $OH^-(aq)$ ions

WACE 3AB 2014 Q16:

Consider the self-ionisation of water:

$$2\ H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq) \qquad \Delta H > 0.$$

Which of the following statements about aqueous solutions is true?

I. All aqueous solutions contain H_3O^+ and OH^- ions
II. In any neutral aqueous solution at any temperature, $[H_3O^+] = [OH^-]$
III. In aqueous solutions with pH greater than 7, $[H_3O^+] > [OH^-]$
IV. A neutral aqueous solution at 100 °C has a pH < 7

(a) I only
(b) I and II only
(c) I, II and III only
(d) I, II and IV only

TEE 2009 Q14:

Which one of the following statements best explains why water is classified as a weak electrolyte?

(a) A strong acid or strong base is required to ionise water molecules
(b) The rate of ionisation of water molecules is very slow
(c) When water ionises, the concentration of $OH^-(aq)$ is equal to the concentration of $H^+(aq)$
(d) A small proportion of the water molecules will form $H^+(aq)$ and $OH^-(aq)$

VCE 2015 Q22:

What is the pH of a 0.0500 mol L^{-1} solution of barium hydroxide, $Ba(OH)_2$?

(a) 1.00
(b) 1.30
(c) 12.7
(d) 13.0

HSC 2014 Q14:

What is the pH of a 0.018 mol L^{-1} solution of hydrochloric acid?

(a) 0.74
(b) 0.96
(c) 1.04
(d) 1.74

TEE 2009 Q17:

The pH of a solution formed by adding 200 mL of water to 20.0 mL of 2.00 mol L^{-1} hydrochloric acid is:

(a) 0.39
(b) 0.70
(c) 0.74
(d) 1.39

TEE 2006 Q15:

20.0 mL of a 0.0100 mol L^{-1} solution of NaOH is added to 20.0 mL of a 0.0300 mol L^{-1} solution of HCℓ. What is the pH of the resulting solution?

(a) 1.52
(b) 1.70
(c) 2.00
(d) 12.00

In modern exams they don't tend to have big calculation questions like this in multiple choice. You would be more likely to find it in short answer / extended answer, where it would be worth more marks.

VCE 2014 Q4:

If Solution X has a pH of 3 and Solution Y has a pH of 6, we can conclude that

(a) $[H^+]$ in Solution X is 1000 times that of $[H^+]$ in Solution Y
(b) $[H^+]$ in Solution X is half that of $[H^+]$ in Solution Y
(c) $[OH^-]$ in Solution Y is twice that of $[OH^-]$ in Solution X
(d) Solution Y must contain a stronger acid than Solution X

<u>Section 3: Self-ionisation of water, K_w, pH calculations</u>

Use the following information to answer TEE 2008 Questions 25 and 26:

A student has 20.0 mL of 0.15 mol L^{-1} $Ba(OH)_2$ solution and 30.0 mL of 0.223 mol L^{-1} HCℓ solution.

TEE 2008 Q25:

What is the pH of the $Ba(OH)_2$ solution?

(a) 0.52
(b) 2.52
(c) 13.18
(d) 13.48

TEE 2008 Q26:

If the two solutions are mixed, what is the pH of the resulting solution?

(a) 1.13
(b) 1.86
(c) 2.43
(d) 3.16

TEE 2004 Q10:

What is the concentration of a $Ba(OH)_2$ solution that has a pH of 9.30?

(a) 1.00 x 10^{-5} mol L^{-1}
(b) 2.00 x 10^{-5} mol L^{-1}
(c) 2.50 x 10^{-10} mol L^{-1}
(d) 5.01 x 10^{-10} mol L^{-1}

TEE 2004 Q12:

20.0 mL of a 0.0100 mol L^{-1} solution of NaOH is added to 20.0 mL of a 0.0300 mol L^{-1} solution of NaCℓ. What is the pH of the resulting solution?

(a) 2.00
(b) 7.00
(c) 11.70
(d) 12.00

Section 3: Self-ionisation of water, K_w, pH calculations

TEE 2002 Q26:

Which of the following statements **best** describes a neutral aqueous solution?

- **(a) The concentrations of H^+ and OH^- are equal**
- (b) The pH is 7
- (c) The solution contains no basic or acidic species
- (d) The solution may contain dissolved salts

VCE 2015 Q22:

The following table shows the value of the ionisation constant of pure water at various temperatures and at a constant pressure.

Temperature (°C)	0	25	50	75	100
K_w	1.1 x 10_{-15}	1.0 x 10^{-14}	5.5 x 10^{-14}	2.0 x 10^{-13}	5.6 x 10^{-13}

Given this data, which one of the following statements about pure water is correct?

- (a) The $[OH^-]$ will decrease with increasing temperature
- **(b) The $[H_3O^+]$ will increase with increasing temperature**
- (c) Its pH will increase with increasing temperature
- (d) Its pH will always be exactly 7 at any temperature

TEE 2001 Q30:

A chemistry measures the pH of four 1.0 x 10^{-2} mol L^{-1} acid solutions, and obtains the following results:

Experiment	Solution	pH
1	1.0 x 10^{-2} mol L^{-1} CH_3COOH	3.4
2	1.0 x 10^{-2} mol L^{-1} H_3PO_4	2.2
3	1.0 x 10^{-2} mol L^{-1} HNO_3	2.0
4	**1.0 x 10^{-2} mol L^{-1} H_2SO_4**	**1.4**

Which experiment result must be **incorrect**?

- (a) Experiment 1
- (b) Experiment 2
- (c) Experiment 3
- **(d) Experiment 4**

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 hydrochloric acid with the same pH as hydrochloric acid, a greater concentration of hydrofluoric acid is required. Explain why this is so. (3 marks)

- **1 mark:** Both acids will have to have the same $[H^+]$ to have the same pH
- **1 mark:** HF does not ionise to the same extent as HCℓ
- **1 mark:** Greater concentration of HF is needed to give the required $[H^+]$

Examiner's comments: Candidates often incorrectly used the terms 'dissociation' and 'ionisation' in part (b). Other candidates simply repeated what was given in the question. Many omitted to indicate that the hydrogen ion concentration needs to be the same to have the same pH.

(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)

- **1 mark:** The fluoride ion hydrolyses resulting the formation of hydroxide ions, ∴ results in a solution with a pH > 7
- **1 mark:** $F^-(aq) + H_2O(\ell) \rightleftharpoons HF(aq) + OH^-(aq)$
- **1 mark:** The chloride ion is the very weak conjugate base of a strong acid, and hence cannot hydrolyse and is a neutral ion.

Examiner's comments: For part (c) it appeared that many candidates used a memorized rubric to predict when a salt solution would be acidic or basic (e.g. NaCℓ comes from a strong acid and strong base so it must be neutral). Since the question gave the salts as being neutral and basic, and explanation was required, not the use of a predictive tool. Candidates were required to answer in terms of the hydrolysis of the anions, but few did so.

WACE 3AB 2015 Q29:

A 25.0 mL solution of nitric acid at 25.0 °C contains 8.50 x 10^{-3} moles of hydrogen ions.

(a) Calculate the hydrogen ion concentration and the pH of the solution. (2 marks)

Description	Marks
$c(H^+)$ = n/v = 8.50 x 10^{-3}/ 0.025 = 0.340 mol L^{-1}	1
pH = $-\log[H^+]$ = $-\log(0.340)$ = 0.469	1
Total	**2**

Examiner's comments: Part (a) was done well generally.

(b) Calculate the pH of the solution after 20.0 mL of 0.300 mol L^{-1} potassium hydroxide solution is added to the original 25.0 mL of nitric acid. (5 marks)

Description	Marks
$n(OH^-)$ = cv = 0.3 x 0.02 = 0.006 mol	1
Recognition that OH^- and H^+ react in 1:1 ratio	1
$n(H^+)$ in excess = 0.0085 – 0.006 = 2.5 x 10^{-3} mol	1
$c(H^+)$ = 2.5 x 10^{-3} / 0.045 = 0.0556 mol L^{-1}	1
pH = $-\log[H^+]$ = $-\log(0.0556)$ = 1.26	1
Total	**5**

Examiner's comments: In part (b), while candidates were required to show their reasoning throughout the calculation, not all did so. A significant number of candidates did not show the 1:1 relationship between hydroxide ions and hydrogen ions, so although the final answer was correct, full marks could not be awarded.

WACE 3AB 2013 Q41:

(d) Lead-acid storage batteries use Pb and PbO_2 electrodes. The overall equation is:

$$Pb(s) + PbO_2(s) + 4\,H^+(aq) + 2\,SO_4^{2-}(aq) \rightarrow 2\,PbSO_4(s) + 2\,H_2O(\ell)$$

i. Determine the number of moles of H^+(aq) in a lead-acid battery that contains 4.50 L of 3.55 mol L^{-1} sulfuric acid solution. Assume full ionisation. (1 mark)

n(H^+ initial) = 2 x n(H_2SO_4) = 2 x c x V = 2 x 3.55 x 4.50 = 32.0 mol

ii. Use the overall battery equation to determine the number of moles of H^+(aq) consumed when discharged of this battery forms 138.1 g of $PbSO_4$(s).
The molar mass of $PbSO_4$ is 303.26 g mol^{-1}. (2 marks)

n($PbSO_4$) = m / M = 138.1 / 303.26 = 0.455 mol

n(H^+ consumed) = 2 x n($PbSO_4$) = 2 x 0.455 = 0.911 mol

iii. Use your answers to (i) and (ii) to determine the concentration of H^+(aq) in the electrolyte in the discharged battery. Assume that the electrolyte volume remains constant, and ignore any changes due to the formation of water. (2 marks)

n(H^+ remaining) = n(H^+ initial) – n(H^+ consumed) = 31.04 mol

c(H^+) = n/V = 31.04 / 4.5 = 6.9 mol L^{-1}

iv. Use your answers to (i) and (iii) to show that when this battery discharges as described above, the change in pH of the electrolyte solution is negligible. Note that in any acid solution whose H^+(aq) concentration is greater than 1 mol L^{-1}, the pH is negative. (3 marks)

pH original = -log(2 x 3.55) = -0.851

pH final = -log(6.898) = -0.839

∴ Very small (difference = 0.012 pH units)

Examiner's comments: Parts (d)(i), (ii) and (iii) were generally well done. The main errors were in assuming only one ionisable hydrogen ion in (d)(i), and the addition of the answers from (i) and (ii) rather than subtraction when answering (d)(iii). Part (d)(iv) produced a mix of responses; about one-third of the candidates did not attempt this question. Some candidates gave a descriptive answer instead of calculating the pH values; others calculated pH based on the number of moles of H^+ rather than the concentration of H^+.

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. (1 mark)

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

OR

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

Examiner’s comments: Generally this question was done well. Candidates should be encouraged to use double arrows.

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

$[H^+] > [OH^-]$ $[H^+] < [OH^-]$ **$[H^+] = [OH^-]$**

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

$[H^+] > [OH^-]$ $[H^+] < [OH^-]$ **$[H^+] = [OH^-]$**

Examiner’s comments: The majority of candidates identified correctly the relative concentrations of H^+ and OH^- at 25 °C, however, many candidates did not realise water was neutral at 10 °C.

(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)

- **1 mark:** The K_w value for 25 °C is greater than K_w at 10 °C, indicating that the formation of products is favoured by an increase in temperature
- **1 mark:** Le Chatelier’s principle predicts that increases in temperature will favour the endothermic reaction
- **1 mark:** Therefore, self-ionisation of water must be an endothermic process.

Examiner’s comments: Many candidates were able to recognize the self-ionisation of water was endothermic although many found it difficult to explain why with reference to changes in K value

Section 3: Self-ionisation of water, K_w, pH calculations

WACE 3AB 2011 Q39:

A student was given three bottles, A, B and C. Each bottle was labelled with its contents as shown in the table below.

Bottle	Contents
A	46.5 mL of 0.010 mol L^{-1} HCℓ
B	65.7 mL of 0.0555 mol L^{-1} HNO_3
C	20.9 mL of 0.4161 mol L^{-1} NaOH

(a) Calculate the pH of the NaOH solution. (2 marks)

Description	Marks
$[H^+] = \dfrac{1\times10^{-14}}{[OH^-]} = \dfrac{1\times10^{-14}}{0.4161} = 2.403\times10^{-14}\ \text{mol L}^{-1}$	1
pH = $-\log^{10}[H^+] = -\log 2.403 \times 10^{-14} = 13.6\ (13.619)$	1

(b) The contents of all three bottles are placed in one beaker and mixed thoroughly. Calculate the pH of the final mixture. (10 marks)

Description	Marks
$n(H^+)$ from HCl = $c \times V = 0.010 \times 0.0465 = 4.65 \times 10^{-4}$ mol	1
$n(H^+)$ from HNO_3 = $c \times V = 0.0555 \times 0.0657 = 3.646 \times 10^{-3}$ mol	1
$n(H^+)_{total} = 4.65 \times 10^{-4} + 3.646 \times 10^{-3} = 4.111 \times 10^{-3}$ mol	1
$n(OH^-) = c \times V = 0.4161 \times 0.0209 = 8.696 \times 10^{-3}$ mol	1
Recognition that 1 mole of H^+ reacts with 1 mole of OH^-; this may be by showing the mole relationship $n(H^+) = n(OH^-)$ or giving the balanced equation	1
$n(OH^-)$ reacted = $n(H^+)_{total} = 4.111 \times 10^{-3}$ mol	1
$n(OH^-)$ excess = $8.696 \times 10^{-3} - 4.111 \times 10^{-3} = 4.585 \times 10^{-3}$ mol	1
$c(OH^-) = \dfrac{n}{V} = \dfrac{4.585\times10^{-3}}{(0.0465+0.0657+0.0209)} = \dfrac{4.585\times10^{-3}}{0.1331} = 3.444\times10^{-2}\ \text{mol L}^{-1}$	1
$[H^+] = \dfrac{1\times10^{-14}}{[OH^-]} = \dfrac{1\times10^{-14}}{3.444\times10^{-2}} = 2.903\times10^{-13}\ \text{mol L}^{-1}$	1
pH = $-\log_{10}[H^+] = -\log 2.903 \times 10^{-13} = 12.5\ (12.537)$	1
Question incorrectly answered	0
Question not attempted	–
Total	**10**

N.B.: Steps may be amalgamated

Students may also calculate pH using:
pOH = $-\log^{10}[OH^-] = -\log 0.03444 = 1.46$ (2 marks)
pH = 14 – 1.46 = 12.54 (1 mark)

Examiner's comments: Part (b) presented challenges for a large number of candidates. Many did not recognise the chemical reaction that would take place when the solutions are mixed, and therefore did not pay attention to the stoichiometric aspects of the problem. Those who recognise the neutralisation reaction and the need to calculate excess reactants performed well in answering the question. A common error was simply to calculate the $[H^+]$ of all solutions (including NaOH) and sum the hydrogen ion concentrations.

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.

(4 marks)

In HCℓ(aq), all of the molecules ionise, so $[H^+]_{HCl}$ = 0.200 mol L^{-1}

For H_2SO_4, the first ionisation step goes to completion…

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$$

The second ionisation step does not…

$$HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$$

If both steps involved full ionisation then $[H^+]$ for the H_2SO_4 would be equal to 2 x 0.100 = 0.200 mol L^{-1}, but the partial ionisation in step 2 means that $[H^+]$ < 0.200 mol L^{-1}. This gives the H_2SO_4 a higher pH than the HCℓ solution.

- **1 mark: equation(s)**
- **1 mark: explanation of HCℓ**
- **2 marks: explanation of H_2SO_4**

TEE 2003 Q8:

The pH of a 0.0010 mol L^{-1} solution of HCℓ is 3. The pH of a 1.0 mol L^{-1} solution of CH_3COOH is also about 3. Explain these observations using equations where appropriate. (4 marks)

To have the same pH both solutions must have the same $[H^+]$.

HCℓ fully ionises,

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

so $[H^+]$ in HCℓ = 0.0010 mol L^{-1}.
pH = -log$[H^+]$ = -log(0.0010) = 3, matching the information provided in the question.

CH_3COOH is a weak acid, so only a small percentage ionises.

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

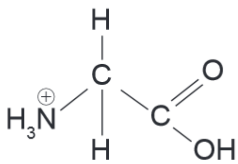
This means that $[H^+]$ << $[CH_3COOH]$. The 1.0 mol L^{-1} solution of CH_3COOH must only produce ~0.001 mol L^{-1} of H^+ ions in solution.

Section 4: Equilibrium, acid dissociation constants, buffers

WACE 2016 Sample Exam Q6: ***Originally from WACE 3AB 2012***

Consider the list below:

i. PO_4^{3-}

ii. 

iii. $NH_2CH_2COO^-$

iv. Na_2HPO_4

Which **two** of the above species, when mixed together in water, form a buffer solution?

(a) i and ii
(b) iii and iv
(c) i and iv
(d) ii and iii

WACE 2016 Sample Exam Q8: ***Originally from WACE 3AB 2010***

Hydrochloric acid (HCℓ) is a stronger acid than the ammonium ion (NH_4^+). Which one of the statements below is **true**?

(a) The equilibrium constant for the hydrolysis of HCℓ is smaller than that for NH_4^+
(b) $C\ell^-$(aq) is a weaker base than NH_3(aq)
(c) Solutions of HCℓ will always have more hydrogen ions than solutions of NH_3
(d) The pH of a 0.1 mol L^{-1} solution of HCℓ will be greater than the pH of a 0.1 mol L^{-1} solution of NH_3

WACE 3AB 2014 Q13:

Which of the following is the strongest acid?

	Acid	Acid dissociation (equilibrium) constant
(a)	CH_3COOH	1.8×10^{-5}
(b)	HCO_3^-	5.6×10^{-11}
(c)	HF	6.8×10^{-4}
(d)	**$H_2C_2O_4$**	**5.4×10^{-2}**

WACE 3AB 2013 Q18:

A buffer solution is prepared by mixing equal moles of sodium acetate (ethanoate) and acetic acid in water. Which of the following statements applies to the buffer?

(a) Addition of a few drops of concentrated nitric acid will produce more acetic acid molecules
(b) The sodium ions play a significant role in the buffering action
(c) Addition of water to the buffer will reduce its buffering capacity
(d) Most of the hydrogen ions will be supplied by water

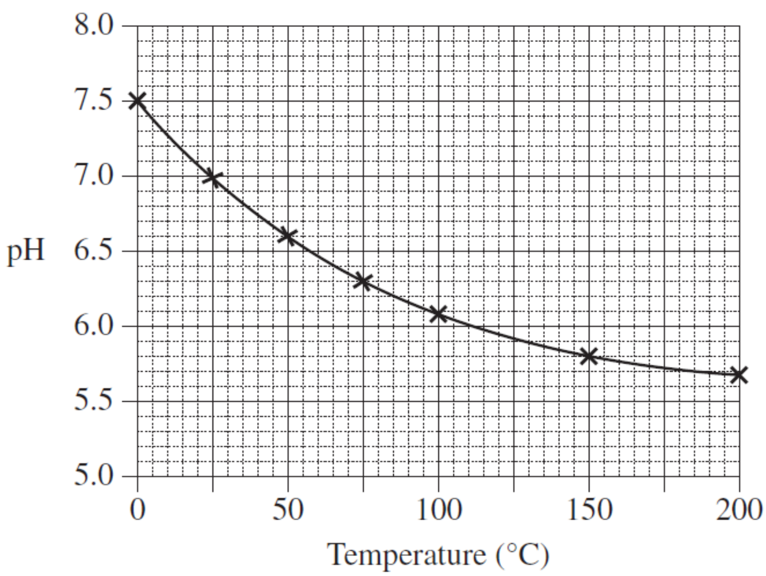
WACE 3AB 2010 Q6:

Which one of the following pairs of substances forms a buffer in aqueous solution?

(a) $HC\ell$ and $NaC\ell$
(b) H_2SO_4 and Na_2SO_4
(c) $NH_4C\ell$ and $NaNH_2$
(d) $NaHCO_3$ and Na_2CO_3

HSC 2014 Q8:

The graph shows the pH of a solution of a weak acid, HA, as a function of temperature.



What happens as the temperature decreases?

(a) HA becomes less ionised and the H^+ concentration increases
(b) HA becomes less ionised and the H^+ concentration decreases
(c) HA becomes more ionised and the H^+ concentration increases
(d) HA becomes more ionised and the H^+ concentration decreases

WACE 3AB 2014 Q19:

A buffer solution is made by dissolving ammonium chloride in a dilute solution of ammonia. The following equilibrium exists in the prepared solution:

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

A small amount of a strong acid is added to the buffer solution. Once the equilibrium has been re-established, the effect would be:

(a) an overall decrease in the H^+ ion concentration
(b) that the equilibrium has shifted to the left
(c) an overall increase in the NH_4^+ ion concentration
(d) an overall increase in the OH^- ion concentration

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} $HC\ell(aq)$ to this buffer does not significantly change its pH. Explain this observation, including any relevant equation(s). (3 marks)

$$CH_3CH_2COOH + H_2O \rightleftharpoons CH_3CH_2COO^-(aq) + H_3O^+(aq)$$

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

Examiner's comments: More than 14% of candidates did not attempt part (d)(i) and those that did so performed poorly. It might be inferred that the concept of buffer has not been understood clearly by the candidates. Alternately, the references to volumes and concentrations in the question might have discouraged candidates as only a qualitative understanding of buffers is required by the syllabus. On careful reading of the question, candidates should have realise that this information was only given to provide the conditions necessary to answer the question; no calculation was required.

(ii) State **two** conditions required to ensure that this system has a high buffering capacity. (3 marks)

One: **Equal concentrations of acid and conjugate base**

Two: **High concentrations of acid and conjugate base**

Examiner's comments: Despite part (d)(ii) being a simple recall question of the conditions required for high buffering capacity many candidates failed to get more than one out of the two available marks. Candidates need to take care to use the appropriate terms such as 'concentration' rather than 'amount' or 'moles'.

WACE 3AB 2013 Q31:

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)

- **1 mark:** $HC_2O_4^-(aq) + H_2O(\ell) \rightleftharpoons C_2O_4^{2-}(aq) + H_3O^+(aq)$
- **1 mark:** $HC_2O_4^-(aq) + H_2O(\ell) \rightleftharpoons H_2C_2O_4(aq) + OH^-(aq)$
- **1 mark:** Use of double arrows for equilibria

Examiner's comments: Part (a) was generally well done

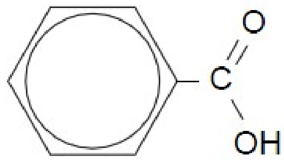
(b) The pH of the solution was measured and found to be less than seven. Based on this observation, state which of the hydrolysis equations has the higher equilibrium constant. Use your understanding of equilibrium concepts to explain your choice fully. (4 marks)

- **1 mark:** If solution has a pH < 7, the concentration of $[H_3O^+] > [OH^-]$
- **1 mark:** K is the ratio of products to reactants
- **1 mark:** H_3O^+ producing equation has the higher K value
- **1 mark:** Thus, H_3O^+ producing equation moves forward to a greater extent than the OH^- producing equation

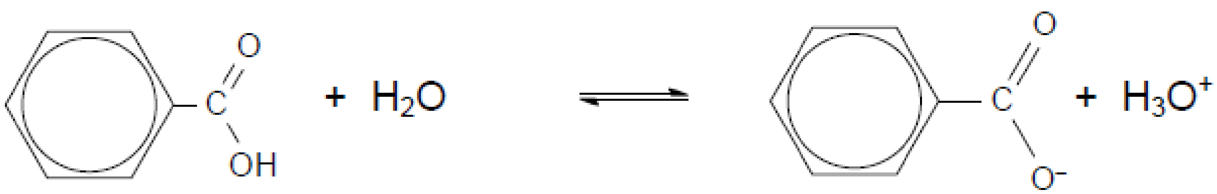
Examiner's comments: In part (b), candidates were able to correctly identify the reaction that produces an acid solution but many did not refer to the second reaction in their response and comment on the relative quantities of H^+ to OH^-.

WACE 3AB 2010 Q29:

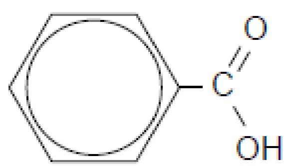
Benzoic acid is found in many berries and some other fruits, and is used as a food preservative. The structure of benzoic acid is shown below. In an aqueous environment, benzoic acid ionises and exists in equilibrium with the benzoate ion.



(a) Write the equation for the reaction between benzoic acid and water. (1 mark)



(b) Draw the structure (either benzoic acid or the benzoate ion) that would predominate in the acidic environment of the stomach. (1 mark)



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. (3 marks)

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

$$C_6H_5COO^-(aq) + H_3O^+(aq) \rightleftharpoons C_6H_5COOH(aq) + H_2O(\ell)$$

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

$$C_6H_5COOH(aq) + OH^-(aq) \rightleftharpoons C_6H_5COO^-(aq) + H_2O(\ell)$$

VCE 2012 Part 2 Q3:

The following weak acids are used in the food industry.

[see table in question booklet]

(a) What does the term 'weak acid' mean? (1 mark)

- **A weak acid is an acid that does not completely ionise in water**

Examiner's comments: Many students defined a weak acid as having a higher pH than a strong acid. This would only apply in aqueous solutions, of the same concentration, of both acids. A 0.0001 mol L^{-1} solution of HCℓ, a strong acid, has a higher pH (4.0) than a 1 mol L^{-1} solution of CH_3COOH, a weak acid (2.4).

Statements such as 'a weak acid does not completely ionise' without mentioning water (or an aqueous solution of the acid) were too vague, given that in the presence of a stoichiometric quantity of a strong base, a weak acid may be expected to fully ionise.

(b) (i) Why are two K_a values listed for malic acid? (1 mark)

- **Malic acid is diprotic. It is able to donate two protons (H^+) so can undergo two ionisation reactions.**

Examiner's comments: Most students did not relate the two K_a values to the ability to donate two protons – that is, that the acid is diprotic. Students should be aware that the COOH groups in carboxylic acids is the source of their acidity, more responses were expected to refer to the presence of two COOH groups on malic acid molecules.

(ii) The equation related to the first K_a value of malic acid is:

$$C_4H_6O_5(aq) + H_2O(\ell) \rightleftharpoons C_4H_5O_5^-(aq) + H_3O^+(aq)$$

Write an appropriate chemical equation that relates to the second K_a of malic acid. (1 mark)

$$\mathbf{C_4H_5O_5^-(aq) + H_2O(\ell) \rightleftharpoons C_4H_4O_5^{2-}(aq) + H_3O^+(aq)}$$

Examiner's comments: Some students who recognized the intent of the question did not provide an 'appropriate' chemical equation. Many students attempted to write the equation starting from a malic acid molecule rather than the ion formed in the equation for the first K_a value. Others wrote an equilibrium law equation, which was often accurate, for the second ionisation; however, the question clearly asked for a chemical equation.

When writing equations for equilibrium reactions, students must remember to include equilibrium arrows.

(c) Sorbic acid, $CH_3(CH)_4COOH$, has antimicrobial properties that are used to inhibit yeast and mould growth. However, its solubility yis very low. The more soluble potassium sorbate is used instead. The antimicrobial activity is retained because an equilibrium exists according to the equation

$$\underset{\text{sorbate ion}}{CH_3(CH)_4COO^-(aq)} + H_2O(\ell) \rightleftharpoons \underset{\text{sorbic acid}}{CH_3(CH)_4COOH(aq)} + OH^-(aq)$$

How would the addition of a small amount of 1.0 mol L^{-1} hydrochloric acid affect the concentration of sorbic acid in solution? Justify your answer in terms of equilibrium principles. (2 marks)

- **The concentration of sorbic acid would increase because the added HCℓ reacts with OH^-, reducing $[OH^-]$, thus causing the reaction to shift to the right to partially oppose the change.**

Examiner's comments: Although there were many excellent responses to this question, overall it was not handled as well as might have been expected. What the sorbate/sorbic acid equilibrium may be unfamiliar to students, it was a relatively fundamental application of Le Chatelier's principle.

Some students argued that the added acid reaction with OH^- to produce H_2O and the increased concentration of water was the reason the equilibrium shifted to the right. However, the concentration of water in an aqueous solution is effectively constant.

HSC 2015 Q24:

(a) Explain why the salt, sodium acetate, forms a basic solution when dissolved in water. Include an equation in your answer. (2 marks)

$CH_3COO^-(aq) + H_2O(\ell) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$
The presence of OH^- ions produced by the hydrolysis of CH_3COO^- increases the pH of the solution and results in a basic pH.

- **1 mark:** Recognises that the salt produces OH^- ions and relates this to the formation of a basic solution
- **1 mark:** Includes a relevant equation

(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. (3 marks)

$CH_3COO^-(aq) + H_3O^+(aq) \rightleftharpoons CH_3COOH(aq) + H_2O(\ell)$
The addition of OH^- ions will cause the reaction with H_3O^+ ions, reducing their concentration in the equilibrium mixture. This will force the reaction to the left to increase the $[H_3O^+]$, thus minimising the change in pH.

- **1 mark:** Explains how an increase in $[OH^-]$ will affect the reaction
- **1 mark:** Relates to minimal change in pH
- **1 mark:** Includes a relevant equation

WACE 3AB 2012 Q43:

Soaps function because their molecules dissolve in both grease and water. Water containing significant quantities of calcium and magnesium ions will not later properly with soap, and will form an insoluble ‘scum’ according to the reaction below. Water that does not later effectively is referred to as ‘hard’ water, and calcium ions are the primary cause of water hardness.

There are a number of methods that may be used to soften hard water. One of these involves the addition of $Ca(OH)_2$ to the water in the process known as liming.

In the liming process, the pH of water is raised when $Ca(OH)_2(s)$ is added.

(c) Calculate the pH of 1.05 x 10^3 L of water solution to which 125 mg of $Ca(OH)_2$ have been added. Assume all added $Ca(OH)_2$ dissolves. (3 marks)

Description	Marks
$n(Ca(OH)_2) = \dfrac{0.125}{74.096} = 1.687000678 \times 10^{-3} \text{mol}$	1
$[Ca(OH)_2] = \dfrac{1.687000678 \times 10^{-3} \text{mol}}{1050 \text{ L}} = 1.606667284 \times 10^{-6} \text{ mol L}^{-1}$ $\therefore [OH^-] = 3.213334567 \times 10^{-6} \text{ mol L}^{-1}$	1
$pOH = -\log(3.213334567 \times 10^{-6}) = 5.49$ $\therefore pH = 14 - 5.49 = 8.51$	1
Total	**3**

Examiner’s comments: A significant number of candidates did not attempt this question. Of those who did, the most common error was to not determine the $[OH^-]$ correctly from the $[Ca(OH)_2]$.

The increase in pH (i.e., addition of OH^-) of the water shifts the equilibria of the carbonate species in the water so that first HCO_3^- predominates, and as the pH is raised further, CO_3^{2-} predominates.

Hard water containing HCO_3^- has significant 'buffering capacity'.

(d) Explain what is meant by the term 'buffering capacity'. (1 mark)

The extent to which a solution can resist changes in pH
or
The extent to which a solution can resist the effects of added H^+ or OH^-

Examiner's comments: Many candidates confused buffer capacity with the definition of a buffer

(e) Write two equations that demonstrate the buffering capacity of hard water containing HCO_3^-. (2 marks)

Description	Marks
$HCO_3^- + H_3O^+ \rightleftharpoons H_2CO_3 + H_2O$	1
$HCO_3^- + OH^- \rightleftharpoons CO_3^{2-} + H_2O$	1
Total	**2**
Or accept as well:	
$HCO_3^- + H_2O \rightleftharpoons CO_3^{2-} + H_3O^+$	1
$HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$	1
Total	**2**

Examiner's comments: A significant number of candidates did not attempt this question.

(f) Write equations to show how the addition of OH^- shifts the equilibria of the carbonate species in the water. (2 marks)

Description	Marks
$H_2CO_3 + OH^- \rightleftharpoons HCO_3^- + H_2O$	1
$HCO_3^- + OH^- \rightleftharpoons CO_3^{2-} + H_2O$	1
Total	**2**

Examiner's comments: This was the question that produced the most non-attempts. Most candidates did not appear to realise that 'cabonate species' did not just mean CO_3^{2-}, so were unable to derive the correct equations.

Section 4: Equilibrium, acid dissociation constants, buffers

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*. (4 marks)

An indicator needs to change colour in different pH conditions. The solution would have the equilibrium:

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

Red Green

When a base is present, the $[H_3O^+]$ will be reduced. Le Chatelier's principle predicts the equilibrium will shift to the right increasing the ionisation of the indicator. This shift causes the green colour to dominate.

Alternatively, when an acid is present the increased concentration of H_3O^+ will shift the equilibrium left and the red colour will dominate.

VCE 2014 Q5:

A 2% solution of glycolic acid, $CH_2(OH)COOH$, is used in some skincare products.

The equation for the ionisation of glycolic acid is:

$$CH_2(OH)COOH(aq) + H_2O(\ell) \rightleftharpoons CH_2(OH)COO^-(aq) + H_3O^+(aq) \qquad K_a = 1.48 \times 10^{-4}$$

Sodium glycolate, $CH_2(OH)COONa$ is a soluble salt of glycolic acid.

(b) How does the pH of glycolic acid change when some solid sodium glycolate is dissolved in the solution? Justify your answer. (2 marks)

The pH increases because the equilibrium moves to the left to partially compensate for the addition of gylcolate ions. This causes the $[H_3O^+]$ to decrease and so the pH increases.

Examiner's comments: Students may not have realised that $CH_2(OH)COONa$ would release $CH_3(OH)COO^-$ into the solution and that subsequent changes should be explained via Le Chatelier's principle. Statements such as 'because glycolic acid is acidic and sodium glycolate is basic they will neutralize each other' and 'a base was added therefore pH increases' overlooked the equilibrium provided. Students were expected to relate change sin pH to changes in the $[H_3O^+]$ and explain why, in the context provided, the $[H_3O^+]$ changes.

HSC 1995 Q21:

The ionisation of any weak acid, HA, in water may be represented as

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

Acid dissociation constants for three weak acids are given below.

Acid	**K_a**
HX	2.3×10^{-4}
HY	7.1×10^{-5}
HZ	5.2×10^{-4}

(a) Arrange these three acids in order of decreasing acid strength. Explain your answer. (2 marks)

- **HZ > HX > HY**
- **Ordered by decreasing K_a values. A higher K_a value indicates reaction more strongly favours the formation of products, ∴ is a stronger acid.**

Examiner's comments: A number of students were confused by the scientific notation. The concept of K_a and decreasing acid strength was not well understood by most candidates.

(b) If all three acids had the same concentration, which would best conduct electricity? Explain your answer. (2 marks)

- **HZ**
- **Given that it has the largest K_a value, it would have the highest concentration of ions (H_3O^+ and A^-) at equilibrium, and ∴ would have the greatest electrical conductivity**

Examiner's comments: In answering this part a number of students confused *ions* with *electrons*.

WACE 2016 Sample Exam Q10:

Over the last 200 years, the pH of oceans has dropped from 8.2 to 8.1. A drop of 0.1 pH units represents an

(a) approximate 20% increase in the concentration of hydrogen ions
(b) increase of the hydrogen ion concentration by a factor of 10
(c) approximate 20% increase in pH
(d) insignificant change in hydrogen ion concentration, due to the large volume of the ocean

WACE 2AB 2010 Q13:

According to Arrhenius theory, what is produced when sodium hydroxide is dissolved in water?

(a) Hydroxide ions
(b) Electrons
(c) Water molecules
(d) Hydrogen ions

HSC 2014 Q3:

Which row of the table correctly matches the scientist(s) with their theory of acids?

	Scientist(s)	Theory
(a)	Arrhenius	Acids contain oxygen
(b)	**Brønsted and Lowry**	**Acids and proton donors**
(c)	Davy	Acids are able to produce hydrogen ions in water
(d)	Lavoisier	Acids contain hydrogen

HSC 2010 Q8:

In a research report a student wrote, ‘Acids are compounds that contain hydrogen and can dissolve in water to release hydrogen ions into solution.’

Who originally stated this theory of acids?

(a) Arrhenius
(b) Brønsted-Lowry
(c) Davy
(d) Lavoisier

Section 5: Ocean acidification, models of acids and bases

HSC 2006 Q11:

In 1884, Svante Arrhenius proposed a definition for acids. His definition was soon accepted as superior to that put forward by earlier chemists.

Why was Arrhenius' definition seen as a major improvement?

(a) It explained why some acids do not contain oxygen
(b) It showed how the solvent can affect the strength of an acid
(c) It showed the relationship between pH and the concentration of H^+ ions
(d) It could be used to explain why some acids are strong and others are weak

HSC 2004 Q5:

Which statement best represents Davy's definition of an acid?

(a) Acids contain oxygen
(b) Acids are proton donors
(c) Acids contain replaceable hydrogens
(d) Acids ionize in solution to form hydrogen ions

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. (2 marks)

$CO_2(aq) + H_2O(\ell) \rightleftharpoons H_2CO_3(aq)$

1 mark: *Balanced equation*
1 mark: *Double arrows for 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. (1 mark)

$H_2CO_3(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$

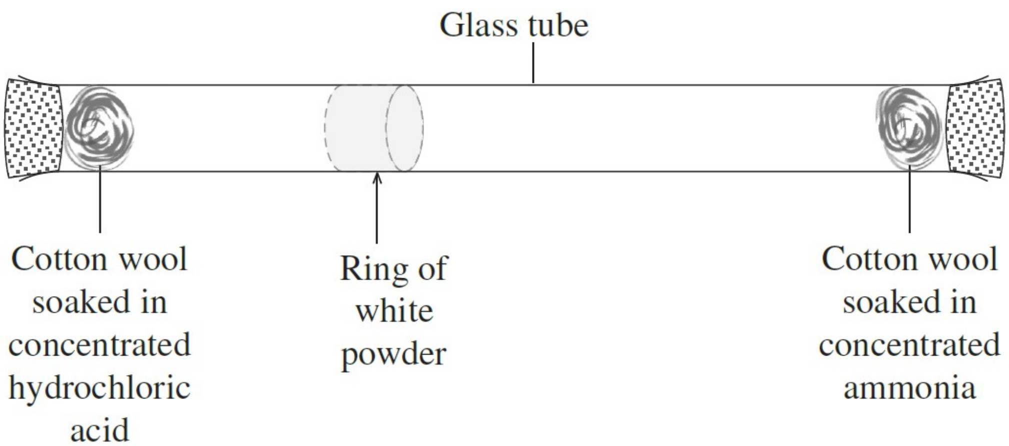
1 mark: *Balanced equation*

(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. (3 marks)

- **CO_3^{2-} and HCO_3^- exist in equilibrium. $HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$**
- **Low pH conditions cause above equation to shift to the left, decreasing $[CO_3^{2-}]$**
- **Reduced $[CO_3^{2-}]$ makes it more difficult to marine organisms such as shellfish and coral to develop calcium carbonate structures (e.g. shells, exoskeletons)**

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? (1 mark)

- **The reaction does not occur in aqueous solution**

(b) Explain why this would be considered a Brønsted-Lowry acid-base reaction. Include an equation in your answer. (2 marks)

- **Reaction involves proton donor ($HC\ell$) and proton acceptor (NH_3)**
- **$HC\ell(g) + NH_3(g) \rightarrow NH_4^+ + C\ell^-$**