

**TEACHER EDITION**

**HSC CHEMISTRY  
STUDENT WORKBOOK**

**CORE MODULE 2:  
THE ACIDIC ENVIRONMENT**

**COLLATED BY:**

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# **THE ACIDIC ENVIRONMENT**

## **How to Use this Workbook**

**This workbook is designed for use both by students and teachers.**

- **It provides a guided pathway through the syllabus content for Core Module 2 of the NSW Chemistry Syllabus.**
- **Worksheets direct students and teachers to the most significant parts of each section of the syllabus, including the first-hand investigations and research topics.**
- **The workbook is structured so that it can form the basis of a teaching sequence and of lesson plans for teachers.**
- **Revision questions provide practice for students at the completion of each syllabus section and allow students to test their understanding of each section.**
- **Questions similar to those asked in the HSC are included in each section to allow students to gain experience in the answering techniques needed for success in the final HSC examination.**
- **The Notes on the Syllabus allow students and teachers to recognise the context of this Core Module and to relate the learning and teaching to the relevant issues for society and the environment.**
- **Students and teachers can use the checklist provided to ensure that they have covered all syllabus “dot-points” in their learning or teaching.**
- **A Periodic Table and Data Sheet, including a Table of Reduction Potentials, are provided.**

## **Notes on the Syllabus**

It is essential that students be thoroughly familiar with the syllabus and they should download it from the NSW Board of Studies website ([www.boardofstudies.nsw.edu.au](http://www.boardofstudies.nsw.edu.au)).

Students should recognise the context of this Core Module. HSC questions are often framed around the following points:

- Acidic and basic species can be found in the living and non-living environments.
- In humans, the slightly acidic skin surface assists in disease control.
- Digestion occurs in both acidic and basic environments to assist in the breakdown of the biopolymers which make up food.
- Micro-organisms found in the digestive system are well adapted to acidic or basic environments.
- Acidic and basic compounds are used widely in industry and in the home. An awareness of the properties of acids and bases is important for safe handling of materials.
- There are concerns about the increased release of acidic and basic substances into the environment and the impact of these substances on the environment and on living things.
- Students should develop an understanding of the history, nature and practice of chemistry and the applications, uses and implications of chemistry for society and the environment.

### **Student Summaries**

Students should prepare their own summaries for each of the syllabus “dot-points”, remembering that recent HSC examination papers have included questions closely following the syllabus wording. These questions may be worth up to 7 or 8 marks. In particular, students should ensure that their summaries include analysis, assessment, evaluation etc when it is specified in the syllabus.

Following is a checklist to allow you to prepare your summaries and to tick them off when completed.

<b>Syllabus statement</b>	<b>Summary completed</b>
Indicators were identified with the observation that the colour of some flowers depends on soil composition	
Perform a first-hand investigation to prepare and test a natural indicator	
Classify common substances as acidic, basic or neutral	
Identify that indicators such as litmus, phenolphthalein, methyl orange and bromothymol blue can be used to determine the acidic or basic nature of a material over a range, and that the range is identified by change in indicator colour	
Identify and describe some everyday uses of indicators including the testing of soil acidity/basicity	
Identify data and choose resources to gather information about the colour changes of a range of indicators	
Solve problems by applying information about the colour changes of indicators to classify some household substances as acidic, neutral or basic	
While we usually think of the air around us as neutral, the atmosphere naturally contains acidic oxides of carbon, nitrogen and sulfur. The concentrations of these acidic oxides have been increasing since the Industrial Revolution	
Identify oxides of non-metals which act as acids and describe the conditions under which they act as acids	
Analyse the position of these non-metals in the Periodic Table and outline the relationship between position of elements in the Periodic Table and acidity/basicity of oxides	
Define Le Chatelier’s Principle	
Identify factors which can affect the equilibrium in a reversible reaction	
Describe the solubility of carbon dioxide in water under various conditions as an equilibrium process and explain in terms of Le Chatelier’s Principle	
Identify data, plan and perform a first-hand investigation to decarbonate soft drink and gather data to measure the mass changes involved and calculate the volume of gas released at 25°C and 100 kPa	
Identify natural and industrial sources of sulfur dioxide and oxides of nitrogen	
Analyse information from secondary sources to summarise the industrial origins of the above gases and evaluate reasons for concern about their release into the environment	
Describe, using equations, examples of chemical reactions which release sulfur dioxide and chemical reactions which release oxides of nitrogen	
Assess the evidence which indicates increases in atmospheric concentration of oxides of sulfur and nitrogen	

Syllabus statement (cont.)	Summary completed
Calculate volumes of gases given masses of some substances in reactions, and calculate masses of substances given gaseous volumes, in reactions involving gases at 0°C and 100 kPa or 25°C and 100 kPa	
Explain the formation and effects of acid rain	
Acids occur in many foods, drinks and even within our stomachs	
Define acids as proton donors and describe the ionisation of acids in water	
Identify acids including acetic (ethanoic), citric (2-hydroxypropane-1,2,3-tricarboxylic), hydrochloric and sulfuric acid	
Describe the use of the pH scale in comparing acids and bases	
Describe acids and their solutions with the appropriate use of the terms strong, weak, concentrated and dilute	
Identify pH as $-\log_{10} [\text{H}^+]$ and explain that a change in pH of 1 means a ten-fold change in $[\text{H}^+]$	
Compare the relative strengths of equal concentrations of citric, acetic and hydrochloric acids and explain in terms of the degree of ionisation of their molecules	
Solve problems and perform a first-hand investigation to use pH meters/probes and indicators to distinguish between acidic, basic and neutral chemicals	
Plan and perform a first-hand investigation to measure the pH of identical concentrations of strong and weak acids	
Gather and process information from secondary sources to write ionic equations to represent the ionisation of acids	
Use available evidence to model the molecular nature of acids and simulate the ionisation of strong and weak acids	
Gather and process information from secondary sources to explain the use of acids as food additives	
Identify data, gather and process information from secondary sources to identify examples of naturally occurring acids and bases and their chemical composition	
Process information from secondary sources to calculate pH of strong acids, given appropriate hydrogen ion concentrations	
Describe the difference between a strong and a weak acid in terms of an equilibrium between the intact molecule and its ions	
Because of the prevalence and importance of acids, they have been used and studied for hundreds of years. Over time, the definitions of acid and base have been refined	
Outline the historical development of ideas about acids including those of: <ul style="list-style-type: none"> <li>- Lavoisier</li> <li>- Davy</li> <li>- Arrhenius</li> </ul>	
Outline the Brønsted-Lowry theory of acids and bases	
Gather and process information from secondary sources to trace developments in understanding and describing acid/base reactions	
Describe the relationship between an acid and its conjugate base and a base and its conjugate acid	

Syllabus statement (cont.)	Summary completed
Identify a range of salts which form acidic, basic or neutral solutions and explain their acidic, neutral or basic nature	
Identify conjugate acid/base pairs	
Identify amphiprotic substances and construct equations to describe their behaviour in acidic and basic solutions	
Identify neutralisation as a proton transfer reaction which is exothermic	
Choose equipment and perform a first-hand investigation to identify the pH of a range of salt solutions	
Describe the correct technique for conducting titrations and preparation of standard solutions	
Perform a first-hand investigation and solve problems using titrations and including the preparation of standard solutions. Use available evidence to quantitatively and qualitatively describe the reaction between selected acids and bases	
Perform a first-hand investigation to determine the concentration of a domestic acidic substance	
Analyse information from secondary sources to assess the use of neutralisation reactions as a safety measure or to minimise damage in accidents or chemical spills	
Qualitatively describe the effect of buffers with reference to a specific example in a natural system	
Esterification is a naturally occurring process which can be performed in the laboratory	
Describe the differences between the alkanol and alkanoic acid functional groups in carbon compounds	
Identify the IUPAC nomenclature for describing the esters produced by reactions of straight-chained alkanoic acids from C1 to C8 and straight-chained primary alkanols from C1 to C8	
Explain the difference in melting point and boiling point caused by straight-chained alkanoic acid and straight-chained primary alkanol structures	
Identify esterification as the reaction between an acid and an alkanol and describe, using equations, examples of esterification	
Describe the purpose of using acid in esterification for catalysis	
Explain the need for refluxing during esterification	
Identify data, plan, select equipment and perform a first-hand investigation to prepare an ester using reflux	
Outline some examples of the occurrence, production and uses of esters	
Process information from secondary sources to identify and describe the uses of esters as flavours and perfumes in processed foods and cosmetics	

## Section 1

### INDICATORS

1. Litmus, phenolphthalein, methyl orange and bromothymol blue can be used to determine the acidic or basic nature of a material over a range that is identified by a change in colour of the indicator.

Indicator	Colours				
	Highly acidic	Slightly acidic	Neutral	Slightly basic (alkaline)	Highly basic (alkaline)
Methyl orange	red	yellow	yellow	yellow	yellow
Bromothymol blue	yellow	yellow	green	blue	blue
Litmus	red	red	purple	blue	blue
Phenolphthalein	colourless	colourless	colourless	colourless	purple/pink

Solutions X, Y and Z were tested with the indicators methyl orange and phenolphthalein. The colours of the solutions are shown below.

Solution	Methyl orange	Phenolphthalein
X	red	colourless
Y	yellow	purple
Z	yellow	colourless

What can you conclude about the acidity/alkalinity of X, Y and Z?

X is strongly acidic

Y is strongly alkaline

Z is not highly acidic nor highly alkaline but could be neutral, slightly acidic or slightly alkaline.

2. Identify and describe at least 3 everyday examples of the use of indicators.

Testing the pH of soil samples.

Testing the pH of pool samples.

Testing the pH of natural water supplies or of rain (to determine level of acid rain).

Answer should include some description as well as identification.

3. Classify the following everyday substances as acidic, basic or neutral.

Drain cleaner	Basic
Vinegar	Acidic
Washing soda	Basic
Oranges	Acidic
Lemons	Acidic
Car battery fluid	Acidic
Stomach fluid	Acidic
“Cloudy ammonia” floor cleaner	Basic
Green apples	Acidic
Oven cleaner	Basic
Household salt (NaCl)	Neutral

4. Describe the pH scale and how it is commonly used to categorise the acidity/basicity/neutrality of various substances.

The pH scale normally ranges from 0 to 14 (highly acidic to highly basic). Neutral substances have a pH of 7 at 25°C.

pH values are determined by the expression  $\text{pH} = -\log_{10} [\text{H}^+]$ .

The greater the pH, the more alkaline the solution. The lower the pH, the more acidic the solution.

5. Describe the pH range over which litmus is useful as an indicator. What are some of the limitations of the use of this indicator?

Litmus changes colour at pH 7, where it is purple. If  $\text{pH} < 7$ , litmus is red; if  $\text{pH} > 7$ , litmus is blue. Litmus is limited in that it cannot differentiate between strong and weak acids or strong and weak bases. Colour change from pink to purple to blue not sharp.

6. A water sample is thought to be slightly acidic. You have been asked to test the water with the following indicators. What colours would you expect the indicators to show if the water sample is slightly acidic?

Indicator	Colour when the sample is slightly acidic
Phenolphthalein	colourless
Methyl orange	yellow
Litmus	red
Bromothymol blue	yellow



7. A sample of pond water was tested and gave a purple colour with phenolphthalein. What does this tell you about the water?  
The water is highly alkaline.
8. You collected a soil sample from your garden and tested it with litmus paper. The litmus stayed a red colour. What conclusions can you draw from this test concerning your garden's pH?  
The soil is either acidic or neutral. It is NOT alkaline.
9. A sample of your swimming pool water sends litmus blue but phenolphthalein remains colourless. What does this indicate about your pool's pH and what can you do to bring the water back to a neutral pH?  
The water is slightly alkaline, so pH is between 7 and 8. Add acid to bring the pool water back to neutral.
10. What is the origin of most indicators?  
Coloured plant materials (e.g. flowers, leaves)
11. You have been given a handful of variegated coloured leaves. Describe a method to extract the fluid to be used as an indicator.  
  
Crush and grind up the leaves using a mortar and pestle. Use water or an ethanol/water mixture to extract the pigment.
12. Before the above indicator can be used, you will need to determine the acid/base range of the fluid. Describe a method to determine the exact range.  
  
Use samples of known pH to determine the colour of the indicator at different pH values.  
  
Once a rough estimate of the pH range over which colour changes occur is achieved, dilute the solutions of known pH to obtain more standards for testing.

13. Outline an example of how two indicators (chosen from the 4 specified in question 6 above) can be used to determine if a sample is slightly acidic (pH of 5-6) or not.

Test the sample with litmus (or bromothymol blue) to show that the sample is acidic. Litmus should turn red; bromothymol blue should turn yellow.

Then test the sample with methyl orange. The indicator should be yellow. Hence the sample is NOT highly acidic but only slightly acidic.

14. Fill in the indicators that would be best for determining the pH of the following common substances. You should choose from any combination of bromothymol blue, litmus, phenolphthalein and methyl orange.

Common substance	pH	Best indicator(s)
oven cleaner	strongly alkaline	phenolphthalein
citric juice	slightly acidic	litmus and methyl orange
washing soda	slightly alkaline	litmus and phenolphthalein
stomach acid	strongly acidic	methyl orange

### Other Questions relating to Section 1

(Note: these questions are not necessarily of HSC standard but will be useful to test your knowledge and understanding of this section)

1. An indicator
- (A) always changes to red in acids
  - (B) always changes to either blue or green in bases
  - (C) only changes colour in acids or bases if water is present or if they are in solution
  - (D) changes colour in acids and bases whether they are wet or dry
2. A particular chemical was extracted and sent to a laboratory for analysis. One of the first tests carried out was with indicators. The following results were obtained:

Indicator	Colour
methyl orange	yellow
phenolphthalein	colourless
bromothymol blue	blue

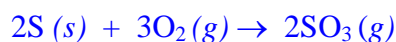
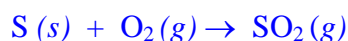
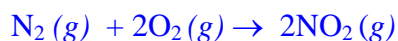
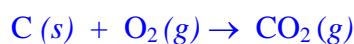
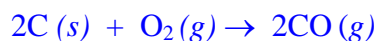
The analysis would conclude the chemical is

- (A) strongly acidic
- (B) slightly acidic
- (C) neutral
- (D) slightly alkaline

## Section 2

### ACIDIC AND BASIC OXIDES

1. The non-metals carbon, nitrogen and sulphur react with oxygen to form oxides. Write equations for the formation of carbon monoxide, carbon dioxide, nitrogen monoxide, nitrogen dioxide, sulphur dioxide and sulphur trioxide.



2. Four of these oxides will react with water to form acids. Identify these acidic oxides and write equations for the formation of acids. Name the acids formed.

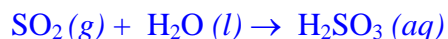
Carbon dioxide, nitrogen dioxide, sulfur dioxide and sulfur trioxide are acidic and react with water to form acids.



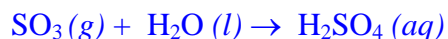
Carbonic acid is formed.



Nitric and nitrous acids are formed.



Sulfurous acid is formed.



Sulfuric acid is formed.

3. Identify at least 2 other elements which are able to form acidic oxides and write the formulae of the acids formed from these acidic oxides.

Phosphorus forms  $\text{P}_2\text{O}_5$ , diphosphorus pentoxide. This forms phosphoric acid,  $\text{H}_3\text{PO}_4$

Chlorine forms  $\text{Cl}_2\text{O}$ , dichlorine monoxide. This forms hypochlorous acid,  $\text{HClO}$ .

4. The metals sodium, magnesium, potassium, calcium and barium react with oxygen to form basic oxides. Write equations for these reactions and identify the products formed when these oxides of sodium, potassium, calcium and barium react with water.
- $4\text{Na (s)} + \text{O}_2\text{ (g)} \rightarrow 2\text{Na}_2\text{O (s)}$  Sodium hydroxide forms when sodium oxide reacts with water.
- $2\text{Mg (s)} + \text{O}_2\text{ (g)} \rightarrow 2\text{MgO (s)}$
- $4\text{K (s)} + \text{O}_2\text{ (g)} \rightarrow 2\text{K}_2\text{O (s)}$  Potassium hydroxide forms when potassium oxide reacts with water.
- $2\text{Ca (s)} + \text{O}_2\text{ (g)} \rightarrow 2\text{CaO (s)}$  Calcium hydroxide forms when calcium oxide reacts with water.
- $2\text{Ba (s)} + \text{O}_2\text{ (g)} \rightarrow 2\text{BaO (s)}$  Barium hydroxide forms when barium oxide reacts with water.
5. Why are the oxides of sodium, potassium, magnesium and calcium described as basic oxides? Because they produce hydroxide ions when they react with water. They also react with acids to form a salt and water (neutralisation).
6. The hydroxides of sodium, potassium, calcium and barium are known as alkalis. These hydroxides dissolve in water to form ions. Identify the ions formed in each of these reactions. What ions are always present in alkaline solutions?
- $\text{Na}^+$  and  $\text{OH}^-$  - sodium ions and hydroxide ions  
 $\text{K}^+$  and  $\text{OH}^-$  - potassium ions and hydroxide ions  
 $\text{Ca}^{2+}$  and  $\text{OH}^-$  - calcium ions and hydroxide ions  
 $\text{Ba}^{2+}$  and  $\text{OH}^-$  - barium ions and hydroxide ions
- $\text{OH}^-$  hydroxide ions are always present in alkaline solutions.
7. Acidic oxides react with bases (such as sodium hydroxide) to form salts and water. Basic oxides react with acids (such as hydrochloric acid) to form salts and water. What do we call this type of reaction? Neutralisation
- Write 4 different equations to form salts and water.
- $\text{CO}_2\text{ (g)} + 2\text{NaOH (aq)} \rightarrow \text{Na}_2\text{CO}_3\text{ (aq)} + \text{H}_2\text{O (l)}$   
 $\text{SO}_2\text{ (g)} + 2\text{NaOH (aq)} \rightarrow \text{Na}_2\text{SO}_3\text{ (aq)} + \text{H}_2\text{O (l)}$
- $\text{Na}_2\text{O (s)} + 2\text{HCl (aq)} \rightarrow 2\text{NaCl (aq)} + \text{H}_2\text{O (l)}$   
 $\text{CaO (s)} + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2\text{ (aq)} + \text{H}_2\text{O (l)}$
8. Magnesium oxide does not readily dissolve or react in water. However, we still describe magnesium oxide as a basic oxide. Suggest a reason. Write an equation for the reaction of magnesium oxide with hydrochloric acid. Magnesium oxide reacts with acids to form a salt and water - neutralisation - as in the reactions of sodium oxide and calcium oxide above.
- $\text{MgO (s)} + 2\text{HCl (aq)} \rightarrow \text{MgCl}_2\text{ (aq)} + \text{H}_2\text{O (l)}$

9. Is CuO an acidic or basic oxide? Write an equation which would show the acidic or basic nature of this oxide.  
 A basic oxide, as it reacts with acids to give a salt and water.  

$$\text{CuO (s)} + 2\text{HCl (aq)} \rightarrow \text{CuCl}_2\text{(aq)} + \text{H}_2\text{O (l)}$$
10. Where in the Periodic Table would you find elements which form basic oxides?  
 Metals form basic oxides. They are found to the left hand side of the Periodic Table.
11. Where in the Periodic Table would you find elements which form acidic oxides?  
 Non-metals form acidic oxides. They are found to the right hand side of the Periodic Table.
12. Would you expect oxides of the following elements to be classified as basic or acidic?
- |            |        |
|------------|--------|
| lithium    | basic  |
| phosphorus | acidic |
| chlorine   | acidic |
| iron       | basic  |
| strontium  | basic  |
13. Oxides of some elements can react with both acids and with sodium hydroxide. A salt and water is formed during both reactions. These oxides are classified as amphoteric because they can react with both acids and bases such as sodium hydroxide. Examples are aluminium oxide, zinc oxide, lead II oxide, tin II oxide and chromium III oxide.
- Write 2 equations to show that zinc oxide is amphoteric (the zincate ion,  $\text{ZnO}_2^{2-}$  is formed when zinc oxide reacts with sodium hydroxide).  

$$\text{ZnO (s)} + 2\text{HCl (aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{H}_2\text{O (l)}$$

$$\text{ZnO (s)} + 2\text{OH}^- \text{(aq)} \rightarrow \text{ZnO}_2^{2-} \text{(aq)} + \text{H}_2\text{O (l)}$$
  - Write 2 equations to show that aluminium oxide is amphoteric (the aluminate ion,  $\text{AlO}_2^-$  is formed when aluminium oxide reacts with sodium hydroxide).  

$$\text{Al}_2\text{O}_3 \text{(s)} + 6\text{HCl (aq)} \rightarrow 2\text{AlCl}_3\text{(aq)} + 3\text{H}_2\text{O (l)}$$

$$\text{Al}_2\text{O}_3 \text{(s)} + 2\text{OH}^- \text{(aq)} \rightarrow 2\text{AlO}_2^- \text{(aq)} + \text{H}_2\text{O (l)}$$

14. Analyse the position of the elements which form oxides (i.e. those mentioned in questions 1 to 13 above) in the Periodic Table and outline the relationship between the position of elements in the Periodic Table and acidity/basicity of oxides. Make sure you have included amphoteric oxides, such as aluminium oxide, in your analysis.

Elements which are metals are found towards the LHS of the Periodic Table. These include sodium, potassium, lithium from Group 1 and magnesium, calcium, barium, strontium from Group 2. Elements such as copper are transition metals. All these metals form positive ions and ionic oxides. The oxides are basic in that they either produce hydroxide ions by reaction with water or react with acids to form a salt and water.

Elements which are non-metals are found towards the RHS of the Periodic Table. These include nitrogen and phosphorus from Group 5, sulfur from Group 6 and chlorine from Group 7. All these non-metals form molecular oxides. The elements listed all form one or more oxides which are acidic in that either they produce acids and hence hydronium ions by reaction with water or they react with bases (such as sodium hydroxide) to form a salt and water.

Elements such as aluminium, zinc, chromium, lead form amphoteric oxides. Amphoteric oxides react to form a salt and water with both acids and alkalis. These elements tend to be close to the diagonal line separating metals and non-metals.

Note: In the HSC, this answer should include some of the equations required in questions 1-13 above.

15. Some oxides are neutral. Explain (with examples) the meaning of a neutral oxide.

Neutral oxides do not react with water to form either acids or alkalis and do not undergo neutralisation reactions with acids or bases.

Carbon monoxide (CO), nitrogen monoxide (NO), dinitrogen monoxide (N<sub>2</sub>O) are neutral oxides.

16. Explain the everyday use of 2 acidic oxides (eg. SO<sub>2</sub> and CO<sub>2</sub>).

SO<sub>2</sub> = sulfur dioxide. It is formed by the reaction of sulfur with oxygen, or by the reaction of metal sulfides with air. It is manufactured as a stage in the production of the important industrial chemical, sulfuric acid. Sulfur dioxide is oxidised to sulfur trioxide, using vanadium (V) oxide supported on silica as a catalyst.



Sulfur dioxide is used as a bleach (especially in the paper industry), as a disinfectant and as a preservative. It is a preservative because it slows down the oxidation by bacteria of dried fruits, the fruit content of wines and fruit juices.

Carbon dioxide is used in the production of soft drinks (soda water), sparkling wines, for fire extinguishers, as dry ice, in the manufacture of urea, an important fertiliser. Its roles in photosynthesis and in the “greenhouse effect” are of significance to society and the environment.

17. Research the internet to determine which oxides of the noble gases have been prepared. Write a summary of your research findings.

Excerpt from [http://en.wikipedia.org/wiki/Noble\\_gas\\_compound](http://en.wikipedia.org/wiki/Noble_gas_compound).

In recent years, several compounds of noble gases, particularly xenon, have been prepared. Among these are the xenon fluorides (xenon difluoride ( $\text{XeF}_2$ ), xenon tetrafluoride ( $\text{XeF}_4$ ), xenon hexafluoride ( $\text{XeF}_6$ ), oxyfluorides ( $\text{XeOF}_2$ ,  $\text{XeOF}_4$ ,  $\text{XeO}_2\text{F}_2$ ,  $\text{XeO}_3\text{F}_2$ ,  $\text{XeO}_2\text{F}_4$ ) and oxides (xenon trioxide ( $\text{XeO}_3$ ), xenon tetroxide ( $\text{XeO}_4$ )). Xenon difluoride can be produced by the simple exposure of Xe and  $\text{F}_2$  gases to sunlight; while the mixing of the two gases had been tried over 50 years before in an attempt to produce a reaction, nobody had thought to simply expose the mixture to sunlight.

18. Under what conditions do these oxides of the noble gases exist?

These compounds are very unstable and only exist at very low temperatures (close to  $-273^\circ\text{C}$ ).

19. Would you expect these oxides to form acidic or basic solutions?

Acidic solutions. The acids would be oxyacids, similar to those formed by chlorine and fluorine. In recent years thousands of compounds of xenon of the type  $\text{XeO}_x\text{Y}_2$  (where  $x$  is 1, 2 or 3 and  $Y$  is any electronegative group) have been made. Perxenic acid has been isolated and identified.

*Note: The HSC examiners, in their report of a recent HSC examination question, criticised students (and presumably their teachers and text book writers!) for not recognising the existence of compounds of the noble gases. It would be wise to include information about the existence and acid-base properties of the oxides of some noble gases in your answers to questions relating to the Periodic Table.*

## Oxides of Sulfur and Nitrogen

1. Identify some natural and industrial sources of sulphur dioxide and oxides of nitrogen.  
 $\text{SO}_2$  - natural sources – geothermal hot springs and volcanoes  
- industrial – processing or burning of fossil fuels (mainly northern hemisphere) and extracting metals from sulfide ores  
  
 $\text{NO}_x$  - natural sources – NO, from lightning (combination of nitrogen and oxygen in air). NO in turn reacts to form  $\text{NO}_2$ .  
- industrial – high temperature combustion in air, forms NO and then  $\text{NO}_2$ .
2. Describe, using equations, examples of chemical reactions which release sulphur dioxide and examples of chemical reactions which release oxides of nitrogen.  
 $\text{SO}_2$  - Burning of sulfur or compounds containing sulfur in oxygen.  
 $\text{S}(s) + \text{O}_2(g) \rightarrow \text{SO}_2(g)$   
 $\text{S}(\text{in compounds}) + \text{O}_2(g) \rightarrow \text{SO}_2(g)$   
  
Heating metal sulfides is a step used to extract metals from their ores.  
 $2\text{ZnS}(s) + 3\text{O}_2(g) \rightarrow 2\text{SO}_2(g) + 2\text{ZnO}(s)$   
  
 $\text{NO}_x$  - High temperature combustion of oxygen and nitrogen in air (can be brought about by lightning). Oxidation of NO to  $\text{NO}_2$ .  
 $\text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g)$   
 $\text{N}_2(g) + 2\text{O}_2(g) \rightarrow 2\text{NO}_2(g)$
3. Assess the evidence which indicates increases in atmospheric concentration of oxides of sulphur and nitrogen and evaluate reasons for concern about their release into the environment.  
Increases in oxides of nitrogen and sulfur lead to increase in acid rain. Change (decrease) in pH of soils and natural waterways can be measured as evidence for increase in concentration of these gases. The effects of acid rain include impact on plant growth, breakdown of carbonate minerals in rocks and in buildings (limestone). Air quality due to pollution by sulfur dioxide was measured by the characteristic unpleasant odour of sulfur compounds. Atmospheric pollution due to sulfur dioxide increased from the 1800s to recently (legislation now prohibits release of sulfur compounds). The number of reports of “smelly” air and health issues can be used as evidence of the concentration of oxides of sulfur.  
Nitrogen oxides cause breathing difficulties and contribute to photochemical smog (visually unattractive and a health hazard). Evidence of the number of days of smog and of respiratory problems can be used as a measure of concentration of oxides of nitrogen.  
Reasons for concern: discussed above. Health issues, plant growth issues, impact of acid rain, visual pollution.
4. Explain the formation and effects of acid rain.  
Acid rain is formed by carbon dioxide, sulfur dioxide, sulfur trioxide, nitrogen dioxide. These react with water to form acids (carbonic, sulphurous, sulfuric, nitric and nitrous). These acids change the pH of soils and natural water supplies and impact on plant and animal metabolism and cause breakdown of carbonate minerals in rocks and in buildings (limestone).



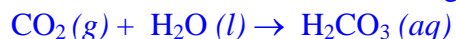
## Other Questions relating to Acidic and Basic Oxides (including Oxides of Sulfur and Nitrogen)

(Note: these questions are not necessarily of HSC standard but will be useful to test your knowledge and understanding of this topic)

1. An oxide occurs
  - (A) when a metal or non-metal reacts with oxygen
  - (B) only after oxygen reacts with a metal or non-metal in solution
  - (C) only when metals react with oxygen
  - (D) only when non-metals react with oxygen
2. Which of the following lists contains only acidic oxides?
  - (A) CO, Na<sub>2</sub>O, SO<sub>2</sub>
  - (B) P<sub>2</sub>O<sub>5</sub>, CO<sub>2</sub>, CO
  - (C) P<sub>2</sub>O<sub>5</sub>, CO<sub>2</sub>, SO<sub>2</sub>
  - (D) MgO, Na<sub>2</sub>O, CaO
3. An indication that an oxide is an “acidic oxide” is when it
  - (A) reacts with water to form a base or with a salt to form an acid
  - (B) reacts with water to form an acid or with a base to form a salt
  - (C) reacts with a base to form an acid
  - (D) turns red litmus to a blue colour
4. An example of an acidic oxide is
  - (A) SO<sub>3</sub>
  - (B) K<sub>2</sub>O
  - (C) CaO
  - (D) CuO
5. Which of the following contains only basic oxides?
  - (A) CuO, MgO
  - (B) SO<sub>2</sub>, CO<sub>2</sub>
  - (C) CaO, CO<sub>2</sub>
  - (D) Al<sub>2</sub>O<sub>3</sub>, P<sub>2</sub>O<sub>3</sub>
6. Which of the following groups contains only elements that can form amphoteric oxides?
  - (A) Na, Ca, Mg, K
  - (B) C, N, S, P
  - (C) Zn, Al, Pb, Sn
  - (D) Pb, Al, Cu, Zn
7. When aluminium oxide reacts with sodium hydroxide, the products formed are
  - (A) sodium aluminate and hydrogen
  - (B) sodium aluminate and water
  - (C) sodium, water and oxygen
  - (D) sodium, water and hydrogen

8. When  $\text{SO}_2$  is released, it reacts with water in the atmosphere over a period of time to produce
- (A)  $\text{SO}_3$
  - (B)  $\text{H}_2\text{SO}_3$
  - (C)  $\text{H}_2\text{CO}_3$
  - (D)  $\text{H}_2\text{SO}_4$
9. Sulfides of metals are used to extract the metal by heating in air. When  $\text{PbS}$  is heated in air, the products will be
- (A)  $\text{PbO} + \text{SO}_2$
  - (B)  $\text{PbS} + \text{O}_2$
  - (C)  $\text{PbSO}_2 + \text{S}$
  - (D)  $\text{PbO} + \text{H}_2\text{O} + \text{S}$
10. Which of the following activities is not a direct source of the  $\text{CO}_2$  gas in the atmosphere?
- (A) use of internal combustion engines
  - (B) photosynthesis
  - (C) rotting cow manure
  - (D) burning plastics
11. Which of the following would be a correct chemical reaction pathway for the production of nitrogen dioxide ( $\text{NO}_2$ )?
- (A) Lightning produces nitrogen oxide, followed by a slow reaction with oxygen to produce nitrogen dioxide.
  - (B) Lightning produces nitric acid, which in turn reacts with oxygen and sunlight to produce nitrogen dioxide.
  - (C) Combustion in cars produces carbon monoxide and this then reacts with nitrogen in the atmosphere to produce nitrogen dioxide.
  - (D) Bacteria from fertilised soil produce nitrogen dioxide directly.
12. Outline how the gas emissions from a manufacturing city can cause the pH to drop in lakes near that city. Use equations to assist your explanation.

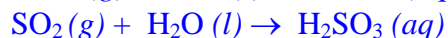
Gas emissions from manufacturing include those which cause acid rain.



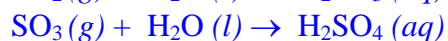
Carbonic acid is formed.



Nitric and nitrous acids are formed.



Sulfurous acid is formed.



Sulfuric acid is formed.

If the concentrations of these acids increase, then the  $[\text{H}_3\text{O}^+]$  increases in natural waterways. Since  $\text{pH} = -\log_{10}[\text{H}^+]$ , if the  $[\text{H}_3\text{O}^+]$  increases the pH decreases.

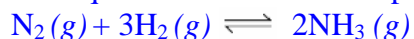
## Equilibrium

1. Explain what is meant by an “equilibrium” reaction. Use the formation of ammonia gas from nitrogen and hydrogen gas in your explanation.

An equilibrium reaction consists of a reaction and its reverse, occurring at the same time and at the same rate.



The equilibrium reaction is represented:



What are the common features of all equilibrium reactions?

The forward and backward reactions are occurring at the same time and at the same rate.

The macroscopic properties (pressure, colour, concentration) remain the same. At a microscopic level, chemical reactions are proceeding but at the same rate in the opposite direction, so the observer cannot see any overall change. Neither matter nor energy enters or leaves the system.

2. Define Le Chatelier’s Principle.

When a system is at equilibrium and a change (concentration, pressure, volume, temperature) is imposed on the system, the equilibrium will shift in a direction to compensate or counteract that imposed change.

3. Identify factors which can affect the equilibrium position in a reversible reaction.

Changes in temperature, concentration and, if gases involved, volume and pressure.

4. Given that the formation of ammonia from its elements nitrogen and hydrogen is an exothermic reaction, predict the equilibrium reaction conditions which would achieve the highest yield of ammonia from a given amount of nitrogen and hydrogen.

The highest yield of ammonia will be achieved at the highest pressure and the lowest temperature.

5. Describe the solubility of carbon dioxide in water under various conditions as an equilibrium process and relate this equilibrium to Le Chatelier's Principle.



When the temperature is increased, the backward reaction (the endothermic reaction) is favoured according to Le Chatelier's Principle (as it takes in the added heat). Hence the carbon dioxide is less soluble at high temperature and more soluble at low temperature.

When the pressure is increased, according to Le Chatelier's Principle, the equilibrium favours the side of the reaction with the smaller number of gas molecules (here the RHS). Thus the forward reaction is favoured by a pressure increase and more  $\text{CO}_2(g)$  dissolves.

### Other Questions relating to Equilibrium

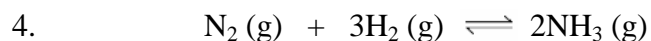
(Note: these questions are not necessarily of HSC standard but will be useful to test your knowledge and understanding of this topic)

- Le Chatelier's Principle states that
  - a system will remain in equilibrium under all changes to conditions except for temperature changes
  - an equilibrium will be achieved faster if a catalyst is present
  - if a system in equilibrium is disturbed, the system will change to maximise the effect of the disturbance
  - if a system in equilibrium is disturbed, then it will change to minimise the effect of the disturbance
- A reaction is in equilibrium when the
  - solid present is all used up
  - pressure of gases is at normal atmospheric pressure
  - forward and reverse reaction rates are equal
  - number of moles on the left hand side equals the number of moles on the right hand side of the reaction



This is the equilibrium reaction which exists in sealed soft drink bottles. What will happen to the soft drink if the lid is left off for several hours?

- The  $\text{CO}_2$  volume decreases.
- The concentration of  $\text{H}_2\text{CO}_3$  decreases.
- The  $\text{H}_2\text{O}$  concentration decreases.
- The concentration of  $\text{CO}_2$  increases in the solution.

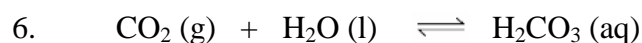


If the overall pressure is increased in the above equilibrium, then

- (A) the concentrations of the  $\text{N}_2(\text{g})$  and  $\text{NH}_3(\text{g})$  will both decrease
- (B) the concentration of the  $\text{NH}_3(\text{g})$  will decrease
- (C) the equilibrium position will move to the right
- (D) there will be no effect on the equilibrium, because the temperature has not changed

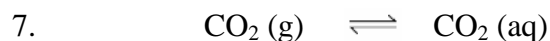
5. Solubility

- (A) is the amount of solid remaining undissolved in a liquid
- (B) is rarely affected by temperature changes
- (C) is the maximum amount of a substance which can dissolve in a given amount of solution
- (D) of gases becomes higher as the pressure on the liquid decreases



As  $\text{CO}_2$  dissolves in the water, heat is given out. If the above reaction were to be cooled down, then according to Le Chatelier's Principle:

- (A) more  $\text{CO}_2$  would be released
- (B) more water would be produced
- (C) the  $\text{H}_2\text{CO}_3$  concentration would decrease
- (D) the right hand direction of the equilibrium would be favoured



The above equilibrium occurs in soft drink bottles and in bodies of natural water (eg. sea water).

- Explain, in terms of Le Chatelier's Principle, what happens to the above equilibrium if the pressure on the equilibrium is increased.

When the pressure is increased, according to Le Chatelier's Principle, the equilibrium favours the side of the reaction with the smaller number of gas molecules (here the RHS). Thus the forward reaction is favoured by a pressure increase and more  $\text{CO}_2(\text{g})$  dissolves.

- Explain, in terms of Le Chatelier's Principle, what happens to the above equilibrium if the temperature of the equilibrium is decreased.

When the temperature is increased, the backward reaction (the endothermic reaction) is favoured according to Le Chatelier's Principle (as it takes in the added heat). Hence the carbon dioxide is less soluble at high temperature and more soluble at low temperature. The equilibrium shifts to the right as the temperature is decreased.

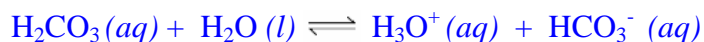
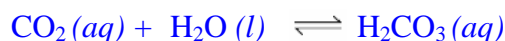
8. Why does the mass of a bottle of carbonated soft drink change if the lid is left off for a few hours?

The equilibrium between  $\text{CO}_2(g)$  and  $\text{CO}_2(aq)$  is destroyed as carbon dioxide gas moves into the atmosphere. The reaction:  $\text{CO}_2(aq) \rightarrow \text{CO}_2(g)$  continues until the pressure of the carbon dioxide gas as it leaves the surface of the soft drink is the same as the partial pressure of carbon dioxide gas in the surrounding atmosphere.

Hence the mass of the soft drink plus bottle will drop as the carbon dioxide moves out of solution and into the atmosphere. It will then remain constant.

9. When carbon dioxide reacts with water, carbonic acid is formed. Carbonic acid is diprotic and ionises partially in water. Both reactions involve equilibria.

- Write 2 equations for the 2 reactions described above.

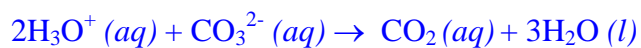


- Use these equations and Le Chatelier's Principle to predict what happens when the pH of the carbonic acid is decreased by addition of a small amount of a strong acid.

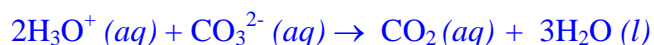
When a small amount of a strong acid is added, the  $[\text{H}_3\text{O}^+]$  increases, hence the second equation above shifts to the left, to make more carbonic acid,  $\text{H}_2\text{CO}_3(aq)$ . As a result, the first equilibrium shifts to the left to compensate for the increase in carbonic acid. The combination of the 2 equilibrium reactions means that the amount of carbon dioxide gas dissolved increases and this in turn shifts the equilibrium in question 7 (above) to the left, releasing more carbon dioxide gas.

10. Write an equation for the reaction which occurs when hydrochloric acid is added to a sodium carbonate solution.

$2\text{HCl}(aq) + \text{Na}_2\text{CO}_3(aq) \rightarrow 2\text{NaCl}(aq) + \text{CO}_2(aq) + \text{H}_2\text{O}(l)$  or, removing the spectator ions, and recognising HCl ionises completely:



11. What happens when any acid is added to any solution containing carbonate ions? Relate your answer to your answer in question 10 above.



This reaction happens whenever any acid is added to any solution containing carbonate ions.

## Molar Volume of Gases

1. The Data Sheet gives the following information:  
Volume of 1 mole ideal gas: at 100 kPa and  
at 0°C (273.15 K) ..... 22.71 L  
at 25°C (298.15 K) ..... 24.79 L

Calculate the volume occupied by :

2 mol of hydrogen gas at 100 kPa and at 0°C

Volume of hydrogen =  $2 \times 22.71 \text{ L} = 45.42 \text{ L} = 45 \text{ L}$  (to 2 s.f. – too many as 2 mol has only 1 s.f.)

0.15 mol of carbon dioxide at 100 kPa and at 0°C

Volume of carbon dioxide =  $0.15 \times 22.71 \text{ L} = 3.4 \text{ L}$  (to 2 s.f.)

2.2 g of carbon dioxide at 100 kPa and at 25°C

Volume of carbon dioxide =  $2.2/44 \times 24.79 \text{ L} = 1.2 \text{ L}$  (to 2 s.f.)

2.  $\text{S (s)} + \text{O}_2 \text{ (g)} \rightarrow \text{SO}_2 \text{ (g)}$

What volume of  $\text{SO}_2$  is produced when 6.4 g of  $\text{O}_2$  is combusted with excess sulfur at 100 kPa and 25 °C?

1 mole  $\text{O}_2 \text{ (g)}$  forms 1 mole  $\text{SO}_2 \text{ (g)}$

$\therefore 6.4/32 \text{ mol O}_2 \text{ (g)}$  forms  $6.4/32 \text{ mol SO}_2 \text{ (g)}$

Volume of  $6.4/32 \text{ mol}$  any gas at 100 kPa and 25 °C =  $6.4/32 \times 24.79 \text{ L} = 5.0 \text{ L}$

3. Calculate the mass of carbon dioxide formed when 95.6 L of oxygen gas at 100 kPa and 25 °C is used to completely combust a sample of ethanol ( $\text{CH}_3\text{CH}_2\text{OH}$ ).  
(Hint: write a balanced chemical equation first.)



95.6 L oxygen gas is  $95.6/24.79$  moles

Since 3 moles  $\text{O}_2$  forms 2 moles  $\text{CO}_2$ , then moles  $\text{CO}_2 = 2/3 \times 95.6/24.79 = 2.57 \text{ mol}$

Mass of carbon dioxide = moles  $\times$  molar mass =  $2.57 \times 44.1 = 113 \text{ g}$

What volume of carbon dioxide will be formed at 100 kPa and 25 °C?

Volume carbon dioxide = moles  $\times 24.79 \text{ L} = 2.57 \times 24.79 = 63.7 \text{ L}$  (3 s.f.)





## Section 3

### ACIDS

Your syllabus focuses on the following 4 acids:

hydrochloric acid (HCl)

acetic acid (ethanoic acid, CH<sub>3</sub>COOH)

citric acid (2-hydroxypropane-1,2,3-tricarboxylic acid, C<sub>6</sub>H<sub>8</sub>O<sub>7</sub>)

sulfuric acid (H<sub>2</sub>SO<sub>4</sub>)

Other significant acids include:

nitric acid (HNO<sub>3</sub>)

nitrous acid (HNO<sub>2</sub>)

hydrobromic acid (HBr)

sulfurous acid (H<sub>2</sub>SO<sub>3</sub>)

carbonic acid (H<sub>2</sub>CO<sub>3</sub>)

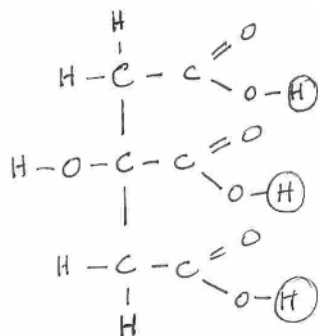
vitamin C (ascorbic acid, C<sub>6</sub>H<sub>8</sub>O<sub>6</sub>)

phosphoric acid (H<sub>3</sub>PO<sub>4</sub>)

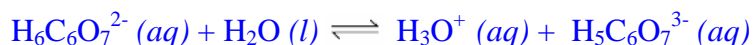
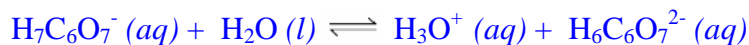
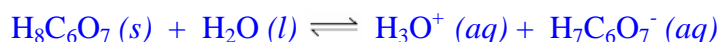
oxalic acid (H<sub>2</sub>C<sub>2</sub>O<sub>4</sub>·2H<sub>2</sub>O)

1. Classify all the above acids as naturally occurring or as synthetic.  
All occur naturally except for hydrobromic and phosphoric acids. Nitric and nitrous acids are formed from the reaction of nitrogen dioxide with water; sulfurous acid is formed from the reaction of sulfur dioxide (from volcanoes) with water, carbonic acid from carbon dioxide in water, vitamin C in citrus fruits and oxalic acid in vegetables (rhubarb), nuts, berries.
2. Classify all the above acids as monoprotic, diprotic or polyprotic.  
Monoprotic: nitric acid (HNO<sub>3</sub>), nitrous acid (HNO<sub>2</sub>), hydrobromic acid (HBr)  
Diprotic: sulfurous acid, carbonic acid, ascorbic acid, oxalic acid  
Triprotic: phosphoric acid
3. Both hydrochloric and sulfuric acid are classified as strong acids.  
Write equations to show the ionisation of each of these acids in water.  
$$\text{HCl (aq)} + \text{H}_2\text{O (l)} \rightarrow \text{H}_3\text{O}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$$
$$\text{H}_2\text{SO}_4 \text{ (aq)} + \text{H}_2\text{O (l)} \rightarrow \text{H}_3\text{O}^+ \text{ (aq)} + \text{HSO}_4^- \text{ (aq)}$$
4. Acetic acid is a weak acid. Write an equation to show the partial ionisation of this acid in water.  
$$\text{CH}_3\text{COOH (aq)} + \text{H}_2\text{O (l)} \rightleftharpoons \text{H}_3\text{O}^+ \text{ (aq)} + \text{CH}_3\text{COO}^- \text{ (aq)}$$

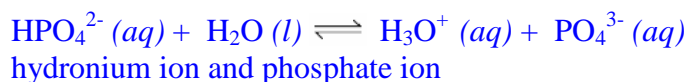
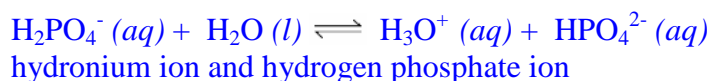
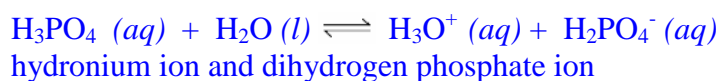
5. Draw a structural formula to represent citric acid. This acid is triprotic. Circle the 3 “acidic” hydrogens on your formula. (Acidic hydrogens are those which can be donated by acids when they ionise or react with bases.)



6. A triprotic acid ionises in 3 steps. Write 3 equations to show the 3 steps in the ionisation of citric acid in water.



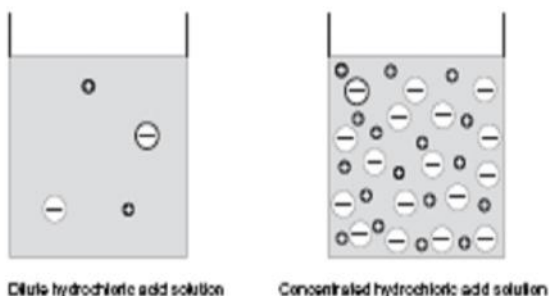
7. Phosphoric acid is triprotic. Write 3 equations to show the 3 steps in the ionisation of phosphoric acid in water and name the species formed in each reaction.



8. A dilute solution contains a small amount of solute in a given amount of solution. Stomach acid could be described as a dilute solution of hydrochloric acid, averaging about 0.01 mol/L.

A concentrated solution contains a large amount of solute in a given amount of solution. Industrially manufactured hydrochloric acid is concentrated averaging 10 mol/L.

The following diagrams represent dilute and concentrated hydrochloric acid.

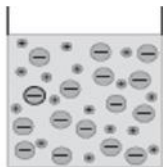


What do the symbols on these diagrams represent?

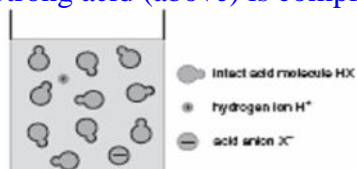
The large circles (with the negative charge) represent chloride ions.

The small circles represent hydrogen (hydronium) ions.

9. The following diagrams represent solutions of strong and weak acids. How do these diagrams show the difference between these 2 types of acids?



Strong acid (above) is completely ionised. No molecules remain.

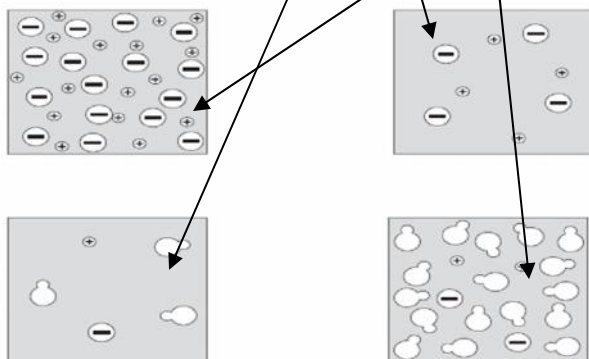


Weak acid (above) is only partially ionised. Intact molecules remain. Only a few ions form.

10. Give an example of:

a dilute solution of a strong acid	0.001 mol/L hydrochloric acid
a concentrated solution of a weak acid	4 mol/L acetic acid
a concentrated solution of a diprotic strong acid	10 mol/L sulfuric acid
a dilute solution of a triprotic weak acid	0.05 mol/L citric acid

11. The 4 diagrams below show a:
- dilute weak acid solution/
  - concentrated weak acid solution
  - dilute strong acid solution
  - concentrated strong acid solution
- Label each of the diagrams appropriately.



12. Strong acids are close to 100% ionised. Give an example. [Hydrochloric acid](#)  
 Moderately strong acids, like phosphoric acid and sulphurous acid, are typically 10 to 20% ionised.  
 Citric acid is about 8% ionised.  
 Weak acids, like acetic acid, are typically only about 1% ionised.

Use this information to explain why 0.1 mol/L solutions of these acids would have different pH values.

[pH measures the concentration of hydrogen ions in solution. The greater the concentration of hydrogen ions, the lower the pH.](#)

[A 0.1 mol/L solution of hydrochloric acid is completely ionised. Hence the concentration of hydrogen ions is also 0.1 mol/L and the pH would be 1.](#)

[Moderately strong acids, if say 10% ionised, would give only 0.01 mol/L of hydrogen ions, and a pH of 2, from a 0.1 mol/L solution of the acid.](#)

[Citric acid would have a slightly lower concentration of hydrogen ions, and a slightly higher pH. The pH of a 0.1 mol/L solution would be about 2.1](#)

[0.1 mol/L acetic acid \(1% ionised\) would have a lower hydrogen ion concentration and a pH of about 3.](#)

13. Pure acids are molecular. They ionise (form ions) when they react with water. Compare the proportion of molecules and ions in strong and weak acids of the same concentration.  
[Strong acids have no molecules and are totally ionised \(i.e.100% ions\).](#)  
[Weak acids have a high proportion of molecules and few ions.](#)

## Other Questions relating to Acids

1. An acid can best be defined as
- (A) a proton donor
  - (B) a proton acceptor
  - (C) an active chemical
  - (D) a chemical that can neutralize organic substances like methane

2. Consider the following acids:

- 1 ethanoic acid
- 2 citric acid (2-hydroxypropane-1,2,3-tricarboxylic acid)
- 3 vitamin C
- 4 HBr
- 5 HCl
- 6 H<sub>2</sub>SO<sub>4</sub>

From the above list, which of the following only includes acids not normally found naturally?

- (A) 4
  - (B) 1, 2, 4, 5 and 6
  - (C) 4 and 6
  - (D) they are all found naturally
3.  $\text{HCl}(g) + \text{H}_2\text{O}(l) \rightarrow \text{Cl}^-(aq) + \text{X}(aq)$

In this equation, hydrogen chloride is reacting with water. Substance X will be

- (A) NaCl
  - (B) OH<sup>-</sup>
  - (C) H<sub>2</sub>O
  - (D) H<sub>3</sub>O<sup>+</sup>
4. Consider the following common substances, which are spread across the pH scale:
- 1 Laundry detergent
  - 2 Wine
  - 3 Vinegar
  - 4 Sea water
  - 5 Saliva
  - 6 Oven cleaner
  - 7 Car battery fluid

Which of the following lists places them in their correct order of decreasing pH?

- (A) 2 5 4 1
- (B) 5 4 1 6
- (C) 5 2 3 7
- (D) 6 2 1

## pH

1. Describe the use of the pH scale in comparing acids and bases.

At 25°C, acids have a pH <7; bases have a pH >7. The lower the pH, the greater the concentration of the hydrogen/hydronium ion present. The higher the pH, the greater the concentration of the hydroxide ion present. In aqueous solutions the pH range lies (practically) between 0 and 14. Theoretically, solutions of pH <0 and >14 are possible.

2. Identify pH as  $-\log_{10}[\text{H}^+]$  and show (mathematically) that a change of 1 in pH means a tenfold change in  $[\text{H}^+]$ .

pH is defined as  $-\log_{10}[\text{H}^+]$ .

If  $[\text{H}^+] = 0.1 \text{ mol/L}$                       pH =  $-\log_{10} [\text{H}^+] = 1.0$

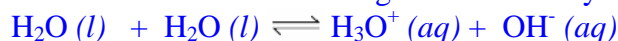
If  $[\text{H}^+] = 0.01 \text{ mol/L}$                       pH = 2.0

If  $[\text{H}^+] = 0.001 \text{ mol/L}$                       pH = 3.0

Hence a change of 1 pH unit means a tenfold change in  $[\text{H}^+]$ .

3. Explain the self-ionisation of water and write the equilibrium equation for this ionisation reaction.

Water is capable of both gaining and donating a proton. As a result, it is amphoteric and 2 water molecules can react together to form hydronium and hydroxide ions.



The molecules are in equilibrium with their ions. The equilibrium lies to the left, with only a very small fraction of the water molecules ionised at any time.

4. Use Le Chatelier's Principle to explain why the concentration of hydroxide ion decreases as acid is added to water.

Using the equilibrium equation in 3 above, if the hydronium ion concentration in water is increased by addition of acid, the equilibrium shifts to the left to compensate. As a result, the concentration of hydroxide ions decreases.

5. Predict what happens to the hydrogen ion concentration if a base (like sodium hydroxide) is added to water.

Using the equilibrium equation in 3 above, if the hydroxide ion concentration in water is increased by addition of base, the equilibrium shifts to the left to compensate. As a result, the concentration of hydronium ions decreases.

6. In aqueous solutions, at 25°C, the product of the concentrations of the hydrogen and hydroxide ions always equals  $10^{-14}$ .

Write this statement as a mathematic expression.



7. Identify pOH as  $-\log_{10}[\text{OH}^-]$  and show (mathematically) that a change of 1 in pOH means a tenfold change in  $[\text{OH}^-]$ .

pOH is defined as  $-\log_{10}[\text{OH}^-]$

If pOH = 4  $[\text{OH}^-] = 10^{-4} = 0.0001 \text{ mol/L}$

If pOH = 3  $[\text{OH}^-] = 10^{-3} = 0.001 \text{ mol/L}$

If pOH = 2  $[\text{OH}^-] = 10^{-2} = 0.01 \text{ mol/L}$

Hence, a change of 1 unit in pOH means a tenfold change in  $[\text{OH}^-]$ .

8. Construct a diagram showing the relationship between  $[\text{H}^+]$ ,  $[\text{OH}^-]$ , pH, pOH, acidic solution, alkaline solution, neutral solution at 25°C.

$[\text{H}^+]$ (mol/L)	$[\text{OH}^-]$ (mol/L)	pH	pOH	Acidic/alkaline/neutral
1.0	$1.0 \times 10^{-14}$	0	14	Acidic
$1.0 \times 10^{-1}$	$1.0 \times 10^{-13}$	1	13	Acidic
$1.0 \times 10^{-2}$	$1.0 \times 10^{-12}$	2	12	Acidic
$1.0 \times 10^{-3}$	$1.0 \times 10^{-11}$	3	11	Acidic
$1.0 \times 10^{-4}$	$1.0 \times 10^{-10}$	4	10	Acidic
$1.0 \times 10^{-5}$	$1.0 \times 10^{-9}$	5	9	Acidic
$1.0 \times 10^{-6}$	$1.0 \times 10^{-8}$	6	8	Acidic
$1.0 \times 10^{-7}$	$1.0 \times 10^{-7}$	7	7	Neutral
$1.0 \times 10^{-8}$	$1.0 \times 10^{-6}$	8	6	Alkaline
$1.0 \times 10^{-9}$	$1.0 \times 10^{-5}$	9	5	Alkaline
$1.0 \times 10^{-10}$	$1.0 \times 10^{-4}$	10	4	Alkaline
$1.0 \times 10^{-11}$	$1.0 \times 10^{-3}$	11	3	Alkaline
$1.0 \times 10^{-12}$	$1.0 \times 10^{-2}$	12	2	Alkaline
$1.0 \times 10^{-13}$	$1.0 \times 10^{-1}$	13	1	Alkaline
$1.0 \times 10^{-14}$	1.0	14	0	Alkaline

## pH CALCULATIONS

When calculating pH values, ensure that you are careful to use the correct number of significant figures. The rule is:

***Determine the number of significant figures of the  $\text{H}_3\text{O}^+$  concentration. Express your pH value to this same number of decimal points; e.g. if a concentration is given to 3 significant figures, the pH would be reported to 3 decimal places.***

1. Calculate the pH of a 0.001 mol/L HCl solution.

$$\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10} 0.001 = -\log_{10} (1 \times 10^{-3}) = 3.0 \text{ (1 d.p.)}$$

2. Calculate the pH of a 0.001 mol/L  $\text{H}_2\text{SO}_4$  solution.

Assume complete ionisation.

$$[\text{H}^+] = 0.002 \text{ mol/L}$$

$$\text{pH} = -\log_{10} (2 \times 10^{-3}) = 2.7 \text{ (1 d.p.)}$$

3. Calculate the  $\text{H}_3\text{O}^+$  concentration of a solution if the pH is 4.16.

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-4.16} = 6.9 \times 10^{-5} \text{ mol/L (2 s.f.)}$$

4. What is the pH of a solution of pure water? What is the  $[\text{H}_3\text{O}^+]$  of pure water at 25°C?

$$\text{pH} = 7.0$$

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-7} \text{ mol/L}$$

5. Calculate the  $\text{OH}^-$  concentration and pOH of a 0.1 mol/L solution of NaOH. Hence determine the pH of this solution.

$$\text{If } [\text{NaOH}] = 0.1 \text{ mol/L, the } [\text{OH}^-] = 0.1 \text{ mol/L and } [\text{H}^+] = 1 \times 10^{-14}/0.1 = 1 \times 10^{-13} \text{ mol/L}$$

$$\text{pOH} = 1.0$$

$$\text{pH} = 14 - 1 = 13.0$$

6. Calculate the  $\text{OH}^-$  concentration and pH of a 0.05 mol/L solution of  $\text{Ba}(\text{OH})_2$ .

Since 1 mole  $\text{Ba}(\text{OH})_2$  forms 2 moles  $\text{OH}^-$

$$[\text{OH}^-] = 0.1 \text{ mol/L}$$

$$\text{pOH} = 1.0$$

$$\text{pH} = 13.0$$

7. Calculate the  $\text{H}_3\text{O}^+$ ,  $\text{Cl}^-$  and  $\text{OH}^-$  concentrations in a  $1.0 \times 10^{-3}$  mol/L solution of HCl.

Since HCl is completely ionised

$$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-3} \text{ mol/L}$$

$$[\text{Cl}^-] = 1.0 \times 10^{-3} \text{ mol/L}$$

$$[\text{OH}^-] = 1.0 \times 10^{-11} \text{ mol/L}$$



8. Calculate the  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$  concentrations in a sample of orange juice with a pH of 3.67.

If  $\text{pH} = 3.67$ , then  $[\text{H}_3\text{O}^+] = 10^{-3.67} = 2.1 \times 10^{-4} \text{ mol/L}$  (2 s.f.)

$\text{pOH} = 14.00 - 3.67 = 10.33$

$[\text{OH}^-] = 10^{-10.33} \text{ mol/L} = 4.7 \times 10^{-11} \text{ mol/L}$  (2 s.f.)

9. Phenolphthalein is colourless when the pH of solution is 8.5 or lower and changes to purple-pink over the pH range 8.5 – 10.0. By what factor does the hydrogen ion concentration of the solution decrease over this range?

$[\text{H}^+]$  at a pH of 8.5 =  $10^{-8.5} \text{ mol/L}$

$[\text{H}^+]$  at a pH of 10.0 =  $10^{-10.0} \text{ mol/L}$

$10^{-8.5}/10^{-10.0} = 10^{1.5} = 31.6$

Answer should be expressed to 1 significant figure (as only 1 decimal point in pH).

The hydrogen ion concentration decreases by a factor of 30.

10. Calculate the pH of a solution containing a mixture of 25.0 mL of 1.25 mol/L  $\text{HNO}_3$  and 25.0 mL of 0.65 mol/L  $\text{KOH}$ .

Moles  $\text{HNO}_3 = 1.25 \times 25.0/1000 = 3.13 \times 10^{-2} \text{ mol}$

Moles  $\text{KOH} = 0.65 \times 25.0/1000 = 1.63 \times 10^{-2} \text{ mol}$

$\text{HNO}_3$  is in excess. All  $\text{KOH}$  will be neutralised by  $1.63 \times 10^{-2}$  moles of  $\text{HNO}_3$  and  $1.50 \times 10^{-2}$  moles of  $\text{HNO}_3$  will remain, in 50.0 mL of solution.

Final  $[\text{H}^+] = 1.50 \times 10^{-2} / 50.0 \times 10^{-3} = 0.30 \text{ mol/L}$  (only 2 s.f. justified)

Final  $\text{pH} = -\log_{10} 0.300 = 0.52$  (2 d.p.)

11. Hydrochloric acid ionises completely. Calculate the pH of a 0.1 mol/L  $\text{HCl}$  solution.

$\text{pH} = -\log_{10} 0.1 = 1.0$

12. A 0.10 mol/L acetic acid solution is 1.5% ionised (at 25°C).

- Are there more molecules of acetic acid or acetate ions in the resulting equilibrium mixture?

Acetic acid

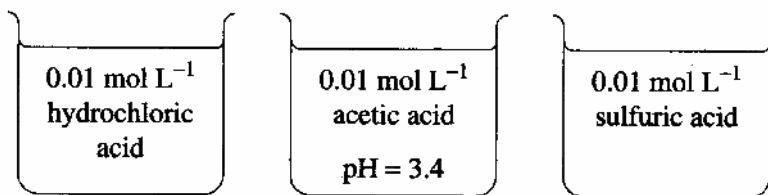
- Calculate the hydrogen (hydronium) and acetate ion concentrations in this solution.

$[\text{H}_3\text{O}^+] = [\text{CH}_3\text{COO}^-] = 1.5/100 \times 0.1 = 1.5 \times 10^{-3} \text{ mol/L}$  (really should only be expressed to 1 s.f., so I changed the question to 2 s.f.)

- Calculate the pH of a 0.1 mol/L solution of acetic acid.

$\text{pH} = -\log_{10} (1.5 \times 10^{-3}) = 2.82$  (2 d.p.)

13. Solutions of hydrochloric acid, acetic acid and sulphuric acid were prepared. Each of the solutions had the same concentration ( $0.01 \text{ mol/L}$ ). The pH of the acetic acid solution was 3.4.



- Calculate the pH of the hydrochloric acid solution.

$[\text{H}^+] = 0.01 \text{ mol/L}$  pH = 2

- Compare the pH of the sulfuric acid solution to the pH of the hydrochloric acid solution. Justify your answer (no calculations necessary).

The pH of the sulfuric acid solution will be lower than the pH for hydrochloric acid, as sulfuric acid is diprotic and at the same concentration.

- Explain why the acetic acid solution has a higher pH than the hydrochloric acid solution.

The acetic acid has a higher pH as it has a lower hydrogen ion concentration than the strong acids. It is only partially ionised, so will have fewer hydrogen ions.

### Other Questions relating to pH calculations

(Note: these questions are not necessarily of HSC standard but will be useful to test your knowledge and understanding of this section)

- If a student mixed  $20 \text{ mL}$  of  $1 \text{ mol L}^{-1} \text{ HCl}$  with  $20 \text{ mL}$  of  $1 \text{ mol L}^{-1} \text{ NaOH}$ , what would be the approximate pH of the final solution?  
(A) 7  
(B) 5  
(C) 11  
(D) 1
- A substance with a pH of 1 is  
(A) nearly neutral  
(B) weakly acidic  
(C) strongly basic  
(D) strongly acidic
- When  $100 \text{ mL}$  of water is added to  $100 \text{ mL}$  of  $1 \text{ mol L}^{-1} \text{ HCl}$  solution, then the new solution  
(A) is more concentrated  
(B) stays at the same pH  
(C) will have a greater concentration of protons because of the extra water  
(D) is more dilute

4. HCl is a strong acid and CH<sub>3</sub>COOH is a weak acid because
- (A) CH<sub>3</sub>COOH<sub>(aq)</sub> always has a lower concentration than HCl<sub>(aq)</sub>
  - (B) CH<sub>3</sub>COOH<sub>(aq)</sub> always absorbs more water
  - (C) HCl<sub>(aq)</sub> has stronger ions in solution
  - (D) CH<sub>3</sub>COOH<sub>(aq)</sub> does not ionise to the same extent as HCl<sub>(aq)</sub> in water
5. A solution of HCl is changed from a pH of 1 to a pH of 2. This means that the concentration of the H<sub>3</sub>O<sup>+</sup> ions (H<sup>+</sup> ions) has
- (A) decreased to 1/10
  - (B) increased by 100
  - (C) decreased to 1/100
  - (D) increased by 10
6. HCl is a strong acid. A 0.1 mol L<sup>-1</sup> M solution of HCl produces a pH of 1. What proportion of the HCl molecules has been left unconverted to ions?
- (A) 0 %
  - (B) 10 %
  - (C) 50 %
  - (D) 90 %
7. Acetic acid is a weak acid. A 0.1 mol L<sup>-1</sup> solution produces a pH of 2.9. What proportion of the acetic acid molecules remains unconverted to ions?
- (A) 0 %
  - (B) 5 %
  - (C) 60 %
  - (D) 99 %
8. Describe a method you could use to determine the pH difference between a 0.5 mol/L of HCl and 0.5 mol/L of CH<sub>3</sub>COOH.  
 Test the pH of the solution. The HCl should have a pH of 0.3, whereas the acetic acid will be higher (close to 3).
9. A 196 g sample of H<sub>2</sub>SO<sub>4</sub> is dissolved in 400 mL of water. Write an ionic equation and determine the H<sub>3</sub>O<sup>+</sup> concentration.
- $\text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \rightarrow 2\text{H}_3\text{O}^+(aq) + \text{SO}_4^{2-}(aq)$   
 Moles H<sub>2</sub>SO<sub>4</sub> = 196/98 = 2      [H<sub>2</sub>SO<sub>4</sub>] = 2/.400 = 5 mol/L  
 [H<sub>3</sub>O<sup>+</sup>] = 10 mol/L (given H<sub>2</sub>SO<sub>4</sub> is diprotic and strong and assume completely ionised).
10. Explain how it is possible to have a solution of a strong acid, like sulfuric acid, which has a pH of 5.  
 It can be a very dilute solution. If pH = 5, the [H<sup>+</sup>] = 10<sup>-5</sup>. Thus (assuming the acid is diprotic) the acid concentration is 0.5 x 10<sup>-4</sup> mol/L.



## Section 4

### THEORIES OF ACIDS AND BASES

1. Outline the historical development of ideas about acids as proposed by:

- Antoine Lavoisier (1780s)

Acids believed to be substances containing oxygen. Lavoisier had been experimenting with oxygen, so the compounds he was studying all contained oxygen. As a vast number of acids (sulfuric, nitric, carbonic, phosphoric, acetic, etc) all contain hydrogen atoms attached to an oxygen atom, the majority of our “current” acids (by Lowry-Bronsted theory) would have been classified as acids by Lavoisier. However, acids like HCl and HBr would not have been classified as acids in 1780.

- Humphry Davy (1815)

Davy said that acids contain replaceable hydrogen. The hydrogen could be replaced by another element, such as a metal. His work was involved with electrolysis, so he tested substances and found that many produced hydrogen at the cathode. He classified these as acids.

- Svante Arrhenius (1884)

Arrhenius defined acids and bases in terms of their reactions in aqueous solutions. Acids gave hydrogen ions as the only positive ions in solution; bases were alkalis – they gave hydroxide ions as the only negative ions in solution. This theory did not take equilibrium into account. Arrhenius would not have considered ammonium chloride (salt) as an acid, as it contained a positive ion other than hydrogen ion, the ammonium ion. Acid-base reactions were neutralisation reactions, with a salt and water being formed.

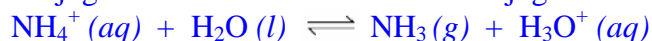
2. Outline the Lowry-Bronsted theory of acids and bases (1923). In your answer include definitions of acids and bases according to this theory and examples of acids and bases.

According to Lowry and Bronsted, acids were proton donors and bases proton acceptors. Acid-base reactions did NOT need to occur in water. The definition was based on the structure of the compound. Acids contain a hydrogen atom attached to an electronegative atom. The hydrogen can be lost (without electrons) from the bond as  $H^+$  ion. Bases all contain a non-bonding (lone) pair of electrons, able to accept the  $H^+$  (by formation of a co-ordinate covalent bond).

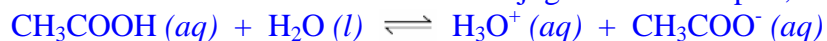
Acids include  $HCl(g)$  and bases include  $NH_3(g)$ , which can react in the gaseous state by transfer of a proton. Ammonium chloride salt is acidic in water, since the ammonium ion can react with the water. Sodium ethanoate is alkaline, as the ethanoate ion is a base and reacts with water to produce hydroxide ions.

3. Describe the relationship between an acid and its conjugate base and a base and its conjugate acid. Show at least two examples of reactions that demonstrate conjugate acid/base pairs.

An acid loses a proton, to become its conjugate base. A base gains a proton to become its conjugate acid. The members of a conjugate acid-base pair only differ by ONE proton.



Ammonium ion and ammonia are a conjugate acid-base pair, as are water and hydronium ion.



Acetic acid and acetate ion; water and hydronium ion – 2 conjugate acid-base pairs.

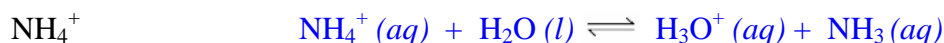
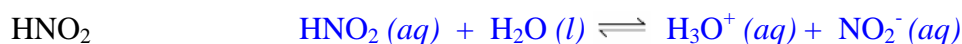
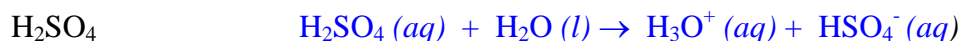
4. List the conjugate bases of the following acids:

HCl	$\text{Cl}^-$
$\text{CH}_3\text{COOH}$	$\text{CH}_3\text{COO}^-$
$\text{H}_2\text{SO}_4$	$\text{HSO}_4^-$
$\text{HNO}_3$	$\text{NO}_3^-$
$\text{HNO}_2$	$\text{NO}_2^-$
$\text{H}_2\text{SO}_3$	$\text{HSO}_3^-$
$\text{H}_2\text{CO}_3$	$\text{HCO}_3^-$
$\text{H}_3\text{PO}_4$	$\text{H}_2\text{PO}_4^-$
$\text{HSO}_4^-$	$\text{SO}_4^{2-}$
$\text{HCO}_3^-$	$\text{CO}_3^{2-}$
$\text{H}_3\text{O}^+$	$\text{H}_2\text{O}$
$\text{H}_2\text{O}$	$\text{OH}^-$
$\text{OH}^-$	$\text{O}^{2-}$
$\text{NH}_4^+$	$\text{NH}_3$
$\text{NH}_3$	$\text{NH}_2^-$

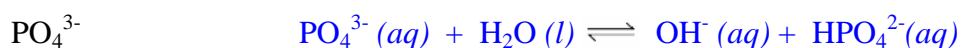
5. List the conjugate acids of the following bases:

$\text{Br}^-$	$\text{HBr}$
$\text{CO}_3^{2-}$	$\text{HCO}_3^-$
$\text{NH}_3$	$\text{NH}_4^+$
$\text{H}_2\text{O}$	$\text{H}_3\text{O}^+$
$\text{OH}^-$	$\text{H}_2\text{O}$
$\text{HSO}_4^-$	$\text{H}_2\text{SO}_4$
$\text{HCO}_3^-$	$\text{H}_2\text{CO}_3$
$\text{H}_2\text{PO}_4^-$	$\text{H}_3\text{PO}_4$
$\text{HPO}_4^{2-}$	$\text{H}_2\text{PO}_4^-$
$\text{PO}_4^{3-}$	$\text{HPO}_4^{2-}$
$\text{CH}_3\text{COO}^-$	$\text{CH}_3\text{COOH}$

6. Write equations to show the reactions of the following acids in water. Only show the transfer of 1 proton to water. Take care as to whether you use single or reversible arrows.



7. Write equations to show the reactions of the following bases in water. Only show the transfer of 1 proton from water. They are all reversible reactions so reversible arrows must be used.



8. Identify species from the lists in questions 4 and 5 above which can act as both acids and bases according to the Lowry-Bronsted theory.

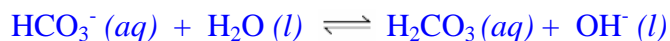
$\text{HSO}_4^-$  (hydrogen sulfate ion),  $\text{HCO}_3^{2-}$  (hydrogen carbonate ion),  $\text{H}_2\text{PO}_4^-$  (dihydrogen phosphate ion),  $\text{HPO}_4^-$  (hydrogen phosphate ion),  $\text{H}_2\text{O}$  (water),  $\text{OH}^-$  (hydroxide ion),  $\text{NH}_3$  (ammonia).

9. Define “amphiprotic” species and give 2 equations to demonstrate that the hydrogen carbonate ion is amphiprotic. You should react the hydrogen carbonate ion with a strong acid and a strong base.

An amphiprotic species is one that can either donate a proton (with a stronger base) or accept a proton (with a stronger acid).



10. Since both hydrogen carbonate ion and water are amphiprotic, you cannot easily predict how the 2 species will react together. It depends on which of the species is the stronger acid. Given that the resulting solution is alkaline when tested with an indicator, write the equation for the reaction which must occur between the 2 species.

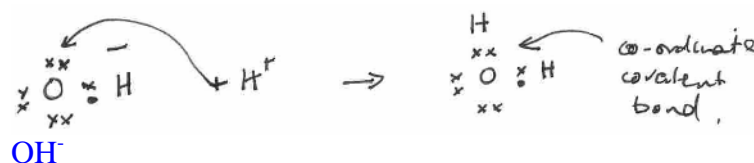
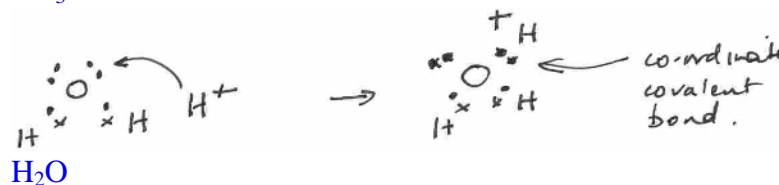
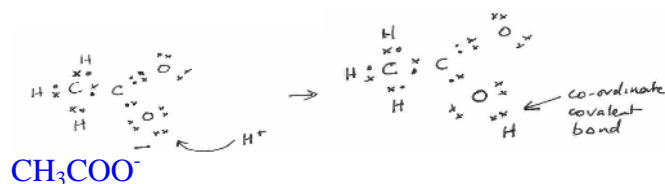
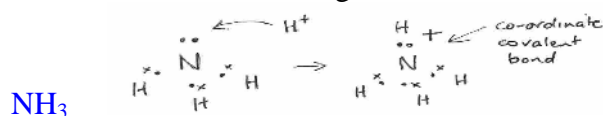


11. All acids contain H atoms which are capable of being lost as  $\text{H}^+$ . These hydrogen atoms can be described as “acidic hydrogens”. Not all hydrogen atoms in an acid must be acidic. Use acetic (ethanoic) acid ( $\text{CH}_3\text{COOH}$ ) to explain these statements. Use the concept of electronegativity of atoms in your response.

In ethanoic acid,  $\text{CH}_3\text{COOH}$ , only 1 of the 4 hydrogen atoms is described as “acidic” because only the hydrogen atom bonded to the oxygen atom can be lost as  $\text{H}^+$ . The O-H bond is polar, with the electrons pulled away from the hydrogen towards the electronegative atom, oxygen. Hence the hydrogen can leave the oxygen atom, without its electron, which stays with the oxygen atom, giving it a negative charge.

The 3 hydrogen atoms attached to carbon are involved in non-polar covalent bonding. Hydrogen and carbon have almost identical electronegativity, so the electrons are evenly shared and the hydrogen is not acidic.

12. All bases must contain a pair of electrons to accept the  $\text{H}^+$  from the acid. Draw electron dot formulae for the following bases and hence demonstrate how they accept protons from acids.





13. Because salts contain ions which can act as acids or bases, salts do not necessarily form neutral solutions:

- Strong acids form very weak bases while strong bases form very weak acids.
- As a result, chloride ions from hydrochloric acid are very weak bases and do not react with water.
- Metal ions such as  $\text{Na}^+$ ,  $\text{K}^+$ ,  $\text{Ca}^{2+}$  and  $\text{Ba}^{2+}$  do not react with water.
- A weak acid, like acetic (ethanoic), forms a strong base, the acetate (ethanoate) ion,  $\text{CH}_3\text{COO}^-$ .

Use the above rules to identify and explain a salt that forms:

- a neutral solution

$\text{NaCl}$  forms a neutral solution, because the sodium ion does not react with water and the chloride ion is a very weak base and does not react with water, coming from a very strong acid,  $\text{HCl}$ .

- an acidic solution

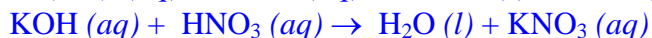
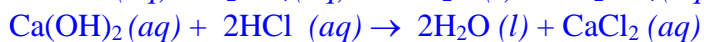
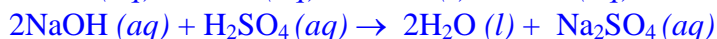
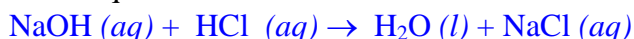
$\text{NH}_4\text{Cl}$  forms an acidic solution, as the ammonium ion acts as an acid, coming from the moderate base, ammonia. Chloride ion is a very weak base and does not react with water, coming from a very strong acid,  $\text{HCl}$ .

- a basic solution

Sodium ethanoate forms a basic solution, because the sodium ion does not react with water but the ethanoate ion acts as a base, coming from the weak acid, ethanoic acid.

14. A salt formed from a strong acid and strong base is neutral or almost neutral.  
eg.  $\text{NaCl}$ ,  $\text{Na}_2\text{SO}_4$ ,  $\text{CaCl}_2$ ,  $\text{KNO}_3$

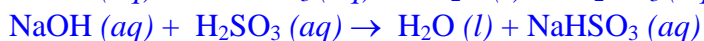
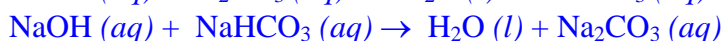
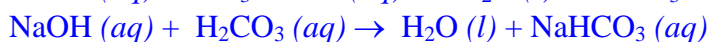
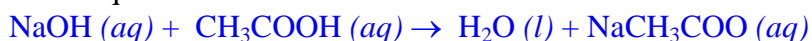
Write equations for the neutralisation reactions which form these salts.



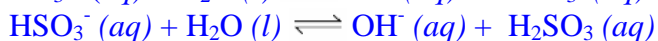
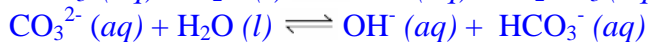
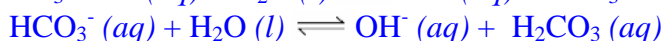
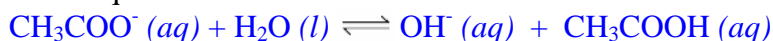
15. A salt formed from a weak acid and a strong base is alkaline (as the ions in the salt react with water to produce hydroxide ions).

eg.  $\text{NaCH}_3\text{COO}$ ,  $\text{NaHCO}_3$ ,  $\text{Na}_2\text{CO}_3$ ,  $\text{NaHSO}_3$ .

Write equations for the neutralisation reactions which form these salts.



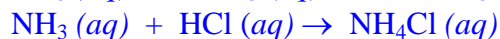
Write equations to show the reactions of the ions which cause these solutions to be alkaline.



All these reactions produce hydroxide ions; hence the solutions are alkaline.

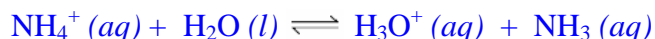
16. A salt formed from a strong acid and weak base is acidic (as the ions in the salt react with water to produce hydronium ions).  
eg.  $\text{NH}_4\text{NO}_3$ ,  $\text{NH}_4\text{Cl}$ .

Write equations for the neutralisation reactions which form these salts.



Write an equation to show the reaction of the ion which causes these solutions to be acidic.

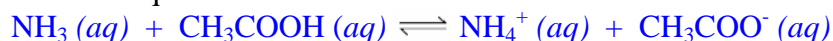
For both salts:



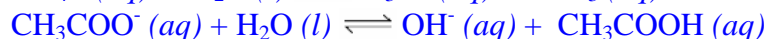
Hydronium ions are produced; hence the solution of the salt is acidic.

17. A salt formed from a weak acid and a weak base is close to neutral (as 1 of the ions reacts to produce hydroxide ions and the other reacts to produce hydronium ions).  
eg.  $\text{NH}_4\text{CH}_3\text{COO}$

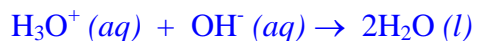
Write an equation for the neutralisation reaction which forms this salt.



Write equations to show the reactions of the ions which cause this solution to be close to neutral.



The ammonium ion produces hydronium ions and the ethanoate ion produces almost the same number of hydroxide ions. These neutralise each other and the overall solution is close to neutral.



18. Identify neutralisation as a proton transfer reaction which is exothermic.  
Write an ionic equation to represent all neutralisation reactions and show the enthalpy term.

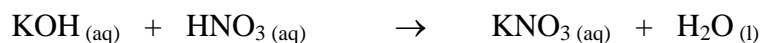
Neutralisation involves the hydronium ion losing a proton which is gained by the hydroxide ion. The reaction gives out heat energy and hence is exothermic.



## Other Questions relating to Theories of Acids and Bases

(Note: these questions are not necessarily of HSC standard but will be useful to test your knowledge and understanding of this section)

- In 1815, Davy proposed that acids
  - donate protons to other substances
  - transfer electrons to other substances
  - contain hydrogen that could be replaced by another element like an active metal
  - contain oxygen
- In 1884, Arrhenius suggested that acids
  - donate protons
  - accept protons
  - contain replaceable hydrogen
  - ionise in water to produce hydrogen ions
- In 1923, Lowry (from Britain) and Bronsted (from Denmark) independently proposed that
  - an acid is a proton donor
  - a base is a proton donor
  - water produces the protons for acids
  - an acid donates electrons to bases
- Which species in the following reaction is acting as a proton acceptor?



- $\text{K}^+$
  - $\text{NO}_3^-$
  - $\text{OH}^-$
  - $\text{H}_2\text{O}$
5.  $\text{CH}_3\text{CH}_2\text{COOH} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{X} + \text{Y}$

Propanoic acid is a weak acid and interacts with water as it ionises. The correct formula for substances X and Y would be

- |     | X                                    | Y                      |
|-----|--------------------------------------|------------------------|
| (A) | $\text{CH}_3\text{CH}_2\text{O}^-$   | $2\text{H}_2\text{O}$  |
| (B) | $\text{CH}_3\text{CH}_2\text{OO}^-$  | $\text{H}^+$           |
| (C) | $\text{CH}_3\text{COOH}$             | $\text{H}_3\text{O}^+$ |
| (D) | $\text{CH}_3\text{CH}_2\text{COO}^-$ | $\text{H}_3\text{O}^+$ |



The conjugate base of the hydrobromic acid in the above equation is

- (A)  $\text{H}_2\text{O}$
- (B)  $\text{HBr}$
- (C)  $\text{H}_3\text{O}^+$
- (D)  $\text{Br}^-$



The conjugate acid of the hydroxide ion is

- (A)  $\text{HCl}$
- (B)  $\text{KOH}$
- (C)  $\text{KCl}$
- (D)  $\text{H}_2\text{O}$

8. The conjugate base of the strong acid  $\text{HNO}_3$  as it ionises in water would be

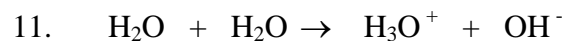
- (A)  $\text{H}_3\text{O}^+$
- (B) very weak
- (C)  $\text{H}_2\text{O}$
- (D) very strong

9. The reason that the salt  $\text{KCl}$  has a  $\text{pH} = 7$  is because

- (A) all salts have a  $\text{pH}$  of 7
- (B) the potassium and the chloride ions each has a concentration of  $1 \times 10^{-7} \text{ mol/L}$
- (C) the ions in the salt neither donate nor accept protons
- (D) the water acts as the acid or the base

10. The reaction between ammonia (a weak base) and sulfuric acid (a strong acid) should produce a salt with a  $\text{pH}$  of about

- (A) 1
- (B) 5
- (C) 7
- (D) 9



The conjugate base for water is

- (A)  $\text{H}_2\text{O}$
- (B)  $\text{H}_3\text{O}^+$
- (C)  $\text{OH}^-$
- (D)  $2\text{H}_2\text{O}$

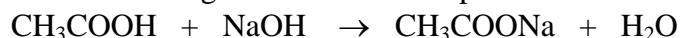
12. Predict the pH of the final solution after 300 mL of 0.1 mol/L HCl reacts with 200 mL of 0.1 mol/L NaOH.

(A) 9  
(B) 7  
(C) 6  
(D) Between 1 and 2

13. An amphiprotic substance

(A) is one that contains a positive and negative proton  
(B) can become either a positive or negative ion  
(C) has the ability to act as a proton donor or acceptor  
(D) is the same as an amphoteric substance

14. In the following reaction there is a proton transfer.



The proton transfer is from

(A)  $\text{CH}_3\text{COOH}$  to  $\text{CH}_3\text{COONa}$   
(B)  $\text{NaOH}$  to  $\text{H}_2\text{O}$   
(C)  $\text{CH}_3\text{COOH}$  to  $\text{H}_2\text{O}$   
(D)  $\text{CH}_3\text{COOH}$  to  $\text{OH}^-$

16. Clearly distinguish between the terms amphiprotic and amphoteric giving examples to illustrate your answer. Explain why aluminium oxide should be called amphoteric rather than amphiprotic.

Amphiprotic substances can ACT as acids or bases by gaining or losing a proton.

e.g.  $\text{HCO}_3^-$ , the hydrogen carbonate ion, reacts with both strong acids like  $\text{H}_3\text{O}^+$  and with strong bases,  $\text{OH}^-$ .



The term amphoteric refers to oxides or hydroxides of only a few elements because they can REACT with (neutralise) both acids and bases such as sodium hydroxide. Examples are aluminium oxide, zinc oxide, lead II oxide, tin II oxide and chromium III oxide.



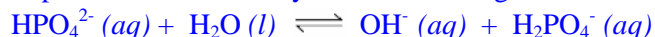
17. From your studies on chemical spills, outline how an acid spill on a highway could be cleaned up with the minimum damage to the environment.

Cover the spill with the powder of a weak base, such as sodium hydrogen carbonate. The acid is neutralised by the base. The solid powder is used to cover the acid, rather than a solution, so that the heat generated does not cause boiling of the solution to the same extent as might occur if a solution of strong base were added directly to the acid spill. The neutralised solution is diluted with water and washed off the highway.

## **Buffer Solutions**

1. Qualitatively describe the effect of buffers with reference to a specific example in a natural system.

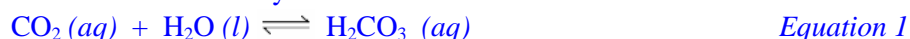
Buffers are equilibrium solutions which contain similar concentrations of moderately weak acids and their conjugate bases (which are moderately weak bases). The conjugate acid/base pair prevents large pH change when small amounts of strong acids or strong bases enter the equilibrium system. The hydrogen carbonate ion/carbonate ion combination in the blood acts as a natural buffer system. The hydrogen phosphate/dihydrogen phosphate system buffers the kidneys, where urine is produced. It is important that the kidney tissue does not get too acidic despite the excretion of an acidic compound.



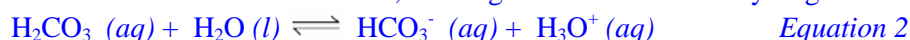
If small quantities of a strong acid are added to the buffer mixture, the equilibrium shifts to the right. If the mixture were to receive small quantities of a strong base (unlikely), then the equilibrium would shift to the left to absorb this extra  $\text{OH}^-$ .

2. Blood must remain in the pH range 7.35 – 7.45. The blood contains a carbonic acid / hydrogen carbonate buffer system. Write equations to show how the carbonic acid is formed and how the hydrogen carbonate ion is formed.

Carbonic acid forms by the reaction of carbon dioxide in water.



The carbonic acid is a weak acid, reacting with water to form hydrogen ions.



3. High concentration of carbon dioxide in the blood leads to an abnormal medical condition. Use the concept of buffers and the above equations to explain how a healthy person prevents this condition.

If the  $[\text{CO}_2(\text{aq})]$  increases, equation 1 moves to the right. This increases the  $[\text{H}_2\text{CO}_3]$ , so equation 2 also moves to the right, producing extra hydronium ions. Normally, in a healthy person, the excess hydronium ions react with the hydrogen carbonate ions which are present in the blood in ample amounts, and the reverse reaction of equation 2 reduces the build-up of hydronium ions and prevents the pH dropping too low.

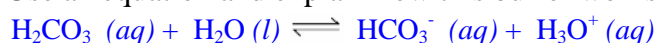
4. Explain what would happen if the blood of a healthy person were exposed to high concentration of an alkali.

Equilibrium 2 would move to the right, as the hydroxide ions from the alkali would remove some of the hydronium ions from the equilibrium. In a healthy person, the buffer would shift to the right to make more hydronium ions and hence keep the pH steady.

5. In the laboratory, an acetic acid / sodium acetate buffer can be used. When equimolar solutions are mixed the pH is 4.74.

What are equimolar solutions? Solutions containing the same number of moles of acetic acid and sodium acetate.

Use an equation and explain how this buffer works.

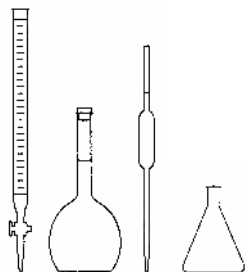


The pH stays 4.74. If a small amount of a strong acid is added, the equilibrium shifts to the left to absorb this added hydronium ion. If a small amount of a strong base is added, the hydroxide ions remove some of the hydronium ions from the equilibrium, so the equilibrium shifts to the right to make more.

## **Titrations**

1. Explain the difference between qualitative and quantitative analysis.  
Qualitative analysis involves observation only, whereas quantitative analysis involves measurement (mass, volume, etc).
2. Explain the difference between volumetric and gravimetric analysis. Are these examples of qualitative or quantitative analysis?  
Volumetric analysis involves measurement of volumes, whereas gravimetric involves measurement of mass/weight. Both are forms of quantitative analysis.
3. What is titration?  
Titration is a volumetric analytical technique used for determining the concentration of a solution, when the concentration of the other reacting solution is known and volumes of the 2 solutions are accurately measured.
4. What is the equivalence point of a titration?  
The point in the titration when the correct number of moles of 1 reactant has been added to react exactly with the number of moles of the other reactant.
5. How is an indicator chosen for a specific titration?  
In an acid-base reaction, the indicator is chosen so that it changes colour at the pH of the solution at the known equivalence point.
6. What is a standard solution?  
  
A standard solution is one that can be made up at an exact, known (and non-changing) concentration.
7. What is a primary standard solution? Give 2 examples of primary standards (1 should be a suitable acid and the other a suitable base).  
  
A primary standard solution is one that can be used as the starting point in a series of titrations to determine the concentration of other solutions. The primary standard is normally a crystalline solid, which can be weighed out accurately (will not absorb water or carbon dioxide from the atmosphere), has a relatively high molar mass (so that errors in transferring the solid are minimised in terms of moles).  
  
Oxalic acid is a suitable primary standard when an acid is needed and anhydrous sodium carbonate is a suitable base.

8. The following equipment is used in all titrations. Name each piece and state how each piece of equipment is rinsed just prior to use.



1.      2.      3.      4.

1.      Burette – washed finally with the solution that will be delivered from it.
  2.      Volumetric flask – washed with distilled water.
  3.      Pipette – washed with the solution to be delivered from it.
  4.      Conical flask – washed with distilled water.
9. Describe the techniques used in preparing a standard solution of a base and in using this solution in a titration. Ensure that you include details of all steps involved in accurate determination of the unknown concentration of an acidic solution.

To make a standard solution of a base:

Weigh out accurately the required mass of the basic primary standard (sodium carbonate). Transfer the mass exactly to a volumetric flask. Use a funnel; wash into the volumetric flask with wash bottle.

To use the base in a titration:

Pipette out of volumetric flask an exact volume (say 25.0 mL) of known concentration base into the clean, rinsed with water, conical flask.

Fill a 50.00 mL burette with the unknown acidic solution. (Burette should have been rinsed with a small quantity of this acidic solution).

Add 3 drops of suitable indicator to the conical flask. (Probably methyl orange).

Titrate known base (sodium carbonate) with unknown acid until methyl orange just changes from yellow to pink.

Calculate moles of sodium carbonate in conical flask and hence (from balanced equation) calculate moles of acidic solution transferred from burette (from a known volume).

Calculate concentration of acidic solution.



10. A primary standard needs to be

- (A) a liquid
- (B) a gas
- (C) an alkali
- (D) pure, crystalline and stable

11. 15.43 g of pure barium hydroxide was dissolved in water and made up to exactly 500 mL in a volumetric flask. Calculate the concentration of the solution.

Molar mass barium hydroxide ( $\text{Ba(OH)}_2$ ) = 171.36 g

Moles  $\text{Ba(OH)}_2 = 15.43/171.36 = 0.0900 \text{ mol}$

$c = n/V = 0.0900/0.500 \text{ mol/L} = 0.180 \text{ mol/L}$

12. What mass of pure sulphuric acid must be dissolved in 100 mL to make a 0.550 mol/L solution?

Molar mass (M)  $\text{H}_2\text{SO}_4 = 98.08 \text{ g}$

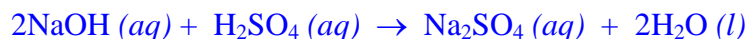
$n = m/M = c V$

Hence  $m = c VM = 0.550 \times 0.100 \times 98.08 \text{ g} = 5.39 \text{ g}$

13. How many moles of HCl are there in 45.3 mL of 0.148 mol/L hydrochloric acid solution?

$n = c V = 0.148 \times 45.3/1000 \text{ mol/L} = 0.00670 \text{ mol}$

14. 25.0 mL of a solution of sodium hydroxide was pipetted into a flask, a few drops of a suitable indicator were added, and the solution titrated with 0.123 mol/L sulphuric acid solution from a burette. 27.4 mL of sulphuric acid was required to reach the equivalence point. Calculate the concentration of the sodium hydroxide solution in mol/L and g/L.



2 moles            1 mole

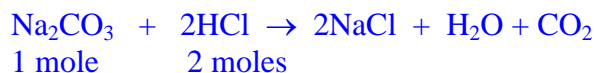
Moles  $\text{H}_2\text{SO}_4 = c V = 0.123 \times 27.4/1000 \text{ mol}$

Hence moles  $\text{NaOH} = 2 \times 0.123 \times 27.4/1000 = 0.00674 \text{ mol}$

Concentration  $\text{NaOH} = n/V = 0.00674/0.025 = 0.270 \text{ mol/L}$

Concentration in g/L =  $0.270 \text{ mol/L} \times \text{molar mass NaOH} = 0.270 \times (22.99 + 16.00 + 1.01)$   
 $= 10.8 \text{ g/L}$

15. 5.24 g of anhydrous sodium carbonate was dissolved in water in a volumetric flask and the volume made up to 250 mL. 10.0 mL of this solution was pipetted into a conical flask and titrated with HCl. 21.6 mL was required to reach the endpoint. Calculate the concentration of the HCl solution.



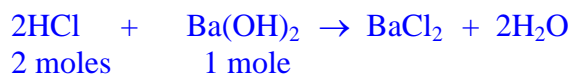
$$\begin{aligned} \text{Moles anhydrous (without water) sodium carbonate} &= 5.24 / (45.98 + 12.01 + 48.00) \\ &= 5.24 / 105.99 \\ &= 0.0494 \text{ mol} \end{aligned}$$

$$\text{Concentration sodium carbonate} = 0.0494 / 0.250 = 0.198 \text{ mol/L}$$

$$\begin{aligned} \text{Moles Na}_2\text{CO}_3 \text{ in 10.0 mL of solution} &= 10.0 / 1000 \times 0.198 \text{ mol} = 0.00198 \text{ moles} \\ \text{Hence moles of HCl} &= 2 \times 0.00198 \text{ mol} = 0.00396 \text{ mol} \end{aligned}$$

$$\text{Concentration HCl} = n/V = 0.00396 / 0.0216 = 0.183 \text{ mol/L}$$

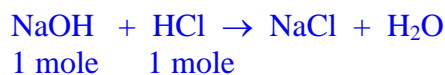
This solution was then used to determine the concentration of an unknown barium hydroxide solution. 25.0 mL of the barium hydroxide required 28.4 mL HCl for exact neutralisation. Calculate the concentration of the barium hydroxide solution.



$$\begin{aligned} \text{Moles HCl used in titration} &= cV = 0.183 \times 0.0284 = 0.00520 \text{ mol} \\ \text{Hence moles Ba(OH)}_2 &= 0.5 \times 0.00520 = 0.00260 \text{ mol} \\ \text{Concentration Ba(OH)}_2 &= 0.00260 / 0.025 = 0.104 \text{ mol/L} \end{aligned}$$

16. In order to determine the concentration of an HCl solution, a student carefully diluted 10.0 mL to 250.0 mL, then titrated 10.0 mL of the diluted solution with 0.147 mol/L sodium hydroxide solution. 21.5 mL was needed to reach the endpoint. Calculate the concentration of HCl.

Original concentrated HCl solution was diluted by a factor of 25.



$$\begin{aligned} \text{Moles NaOH used in titration} &= cV = 0.147 \times 0.0215 = 0.00316 \text{ mol} \\ \text{Hence moles HCl reacted} &= 0.00316 \text{ mol} \\ \text{Concentration of diluted HCl} &= 0.00316 / 0.0100 \text{ mol/L} = 0.316 \text{ mol/L (to 3 s.f.)} \end{aligned}$$

$$\text{Hence concentration of the original HCl solution} = 25 \times 0.316 \text{ mol/L} = 7.90 \text{ mol/L}$$

17. Vinegar is a solution of acetic acid in water. The concentration of the acetic acid in the vinegar was determined by titration. The vinegar was systematically diluted by a factor of 5.

(a) State the equipment needed for this systematic dilution.

Pipette (say 20.0 mL), volumetric flask (100.0 mL)

(b) 25.0 mL of the diluted vinegar was placed in a conical flask. What piece of glassware would be used?

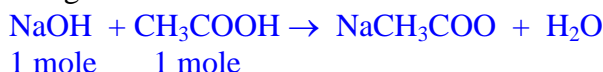
Pipette (25.0 mL)

(c) The dilute acetic acid was titrated with 0.100 mol/L NaOH. An endpoint was reached after a titre of 37.1 mL.

(i) What indicator would be used for this titration?

Phenolphthalein

(ii) Calculate the concentration of acetic acid in both the diluted and original vinegar solutions.



Moles NaOH used in titration =  $cV = 0.100 \times 0.0371 = 0.00371 \text{ mol}$

Hence moles  $\text{CH}_3\text{COOH} = 0.00371 \text{ mol}$

Concentration of diluted  $\text{CH}_3\text{COOH} = 0.00371/0.0250 = 0.148 \text{ mol/L}$

Since the original vinegar was diluted by a factor of 5, concentration of original acetic acid solution =  $5 \times 0.148 = 0.742 \text{ mol/L}$

(iii) Calculate the concentration of acetic acid in the original vinegar in grams per 100 mL.

$$\begin{aligned} 0.742 \text{ mol/L} &= 0.742 \times (\text{molar mass } \text{CH}_3\text{COOH}) \text{ g/L} = 0.742 \times 60.04 = 44.7 \text{ g/L} \\ &= 4.47 \text{ g/100 mL} \end{aligned}$$

(d) Suggest a sequence of steps that would have been done prior to this titration to standardise the sodium hydroxide solution. Include in your explanation reasons why the sodium hydroxide could not be a primary standard.

Sodium hydroxide cannot be weighed out accurately, as it absorbs water and carbon dioxide from the atmosphere and hence its purity and mass changes when exposed to air. Hence it cannot be used as a primary standard.

Oxalic acid is a primary standard and can be weighed out accurately.

The concentration of an oxalic acid solution is determined by calculation (moles/volume).

The oxalic acid can then be titrated against the sodium hydroxide.

25.0 mL of the sodium hydroxide solution is pipetted into a conical flask and the oxalic acid titrated from the burette until the phenolphthalein indicator changes from pink to colourless.

Oxalic acid is diprotic, so 2 moles of NaOH are needed for 1 mole of oxalic acid.

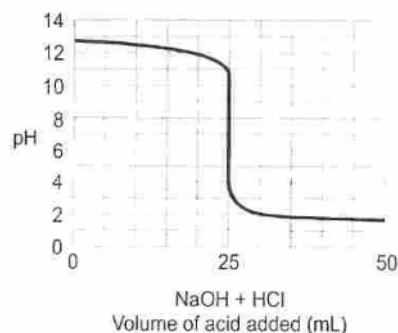
The concentration of the NaOH is determined by calculation.

## Titration pH Graphs

For each graph below assume the following:

- acid in the burette is added to base in the receiving flask
- all solutions have a concentration of 0.1 mol/L

1.

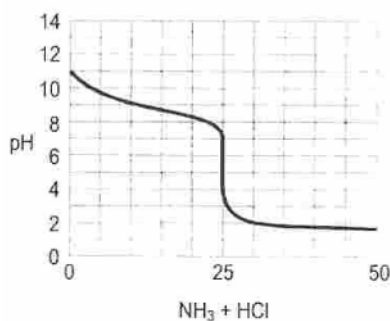


Strong acid is added to strong base. Initially pH remains high as excess of strong base. At equivalence point pH drops suddenly as all base neutralised and now excess of strong acid.

Litmus or bromothymol blue is best but phenolphthalein will give a good result as pH changes so rapidly.

Describe and explain the shape of this curve.  
Suggest suitable indicator(s) for use in this titration.

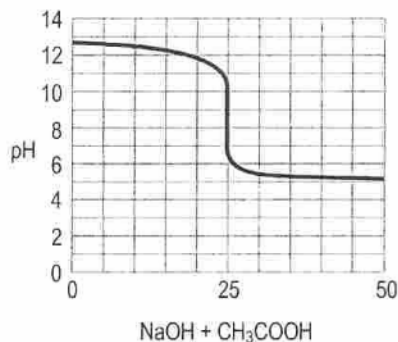
2.



Strong acid-weak base.  
pH drops quickly as weak base.  
pH drops suddenly at equivalence point.  
pH at equivalence point is <7.  
After equivalence point, pH drops to low value, as strong acid in excess.  
Use methyl orange indicator.

Describe and explain the shape of this curve.  
Suggest suitable indicator(s) for use in this titration.

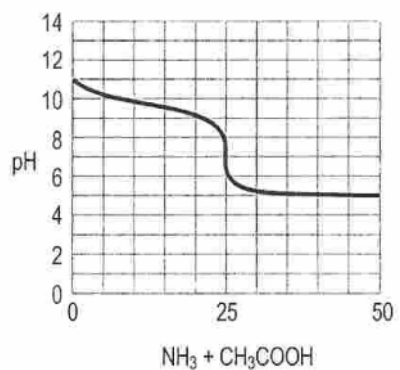
3.



Strong base-weak acid.  
pH initially high and remains high as strong base.  
Equivalence point at pH >7.  
pH does not drop below 5 as only weak acid in excess.  
Use phenolphthalein indicator.

Describe and explain the shape of this curve.  
Suggest suitable indicator(s) for use in this titration.

4.

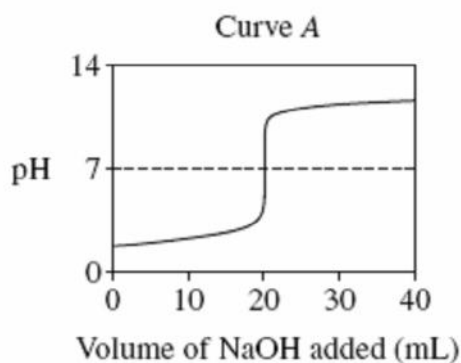


Weak acid-weak base.  
 pH never very high as weak base.  
 pH drops (but only by 5-6 pH levels) at equivalence point.  
 pH after equivalence point never drops to low value as only weak acid in excess.  
 pH at equivalence point close to 7.  
 Litmus or bromothymol blue used as indicators.

Describe and explain the shape of this curve.  
 Suggest suitable indicator(s) for use in this titration.

5. Consider the following titration curves.

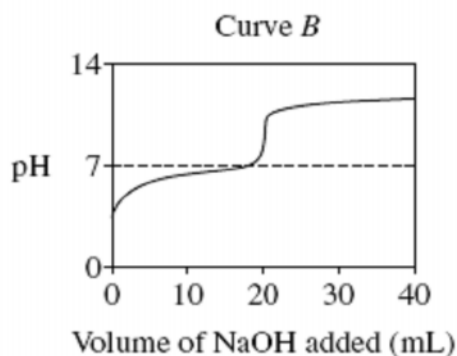
A titration was conducted by adding NaOH from a teflon-coated burette to HCl in a conical flask. The pH in the flask was recorded during the titration and Curve A was produced.



The student was instructed to use either litmus or bromothymol blue as a suitable indicator to determine the endpoint.

- (a) Explain why litmus or bromothymol blue are the most appropriate indicators for this titration.  
 The pH at the equivalence point is 7, so an indicator which changes colour at pH 7 is needed.

A second titration was conducted by adding NaOH to a different acid. The pH in the flask was recorded during the titration and Curve B was produced.



- (b) Nominate a suitable indicator for the second titration.

Phenolphthalein

- (c) Explain why the results of the first titration are not inaccurate if an indicator other than litmus or bromothymol blue is used.

In the first titration, the pH rises suddenly from about 4 to 10 after 20 mL of the NaOH had been added – the equivalence point. If phenolphthalein were used, the end point of the titration would still have been recorded at the equivalence point, as phenolphthalein changes colour from pH 8 to pH 10. The sudden rise in pH when a strong acid – strong base titration occurs ensures that there is only 1 drop of solution difference between a pH of 4 and a pH of 10.

- (d) State a general principle you must use in choosing a suitable indicator for any given titration.

The type of titration must be known (e.g. strong acid versus strong base, weak acid versus strong base).

The pH of the equivalence point must be predicted ( $7$ ,  $>7$ ,  $<7$ ).

The indicator must change colour at the correct pH (the pH at the equivalence point).

For example, for the strong base-weak acid titration, the pH at the equivalence point is between 8 and 10. Phenolphthalein changes colour in this range and so should be used. Methyl orange could not be used as it changes colour between pH values of 3 and 5.

## Section 5

### ESTERIFICATION

1. What functional group is present in all members of the alkanol homologous series?  
– OH (the hydroxyl group).
2. Draw the structural formula for the following alkanols: methanol, ethanol, 1-propanol, 2-propanol.  
 $\text{CH}_3 - \text{O} - \text{H}$                        $\text{CH}_3 - \text{CH}_2 - \text{O} - \text{H}$                        $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{O} - \text{H}$   
methanol                                  ethanol                                  1-propanol  
  
 $\text{CH}_3 - \text{CH}(\text{O} - \text{H}) - \text{CH}_3$   
2-propanol
3. What is the general formula for the alkanols?  
 $\text{C}_n\text{H}_{2n+1}\text{OH}$
4. Would you expect all alkanols to have similar chemical properties? Why/why not?  
Yes. The functional group, the –OH group, is the same in all alkanols and it is this group which determines the majority of the chemical properties (not combustion). The position of the –OH group can alter the properties.
5. Explain why the physical properties of the alkanols depend on both the length of the carbon chain and the properties of the functional group.  
  
The functional group is the hydrophilic –OH group, which is a polar group and which allows hydrogen bonding. The attraction between neighbouring molecules determines the physical properties, so hydrogen bonding and polar attractions caused by the –OH group have a great affect on the physical properties. The longer the carbon chain, the greater the number of atoms in the molecule. Hence the greater the interaction between neighbouring non-polar parts of molecules. Also, the weight of a molecule affects the physical properties. (Heavier molecules change state at higher temperatures and so are more likely to be solids or liquids than gases).
6. Draw the structural formula for the following alkanoic acids: methanoic acid, ethanoic acid, propanoic acid. From your structures derive a general formula.  
 $\text{HCOOH}$                                    $\text{CH}_3\text{COOH}$                                    $\text{CH}_3\text{CH}_2\text{COOH}$   
methanoic acid                                  ethanoic acid                                  propanoic acid  
  
General Formula =  $\text{C}_n\text{H}_{2n}\text{O}_2$
7. What do all the above molecules have in common? Describe the significance of this common feature.  
  
The functional group – COOH, i.e. the carboxylic acid group. This functional group gives the molecules their characteristic properties (both chemical and physical).

8. Consider the melting and boiling points of alkanolic acids and alkanols.  
Fill in the names of the alkanolic acids.

1-alkanol	MP (°C)	BP (°C)	Alkanolic acid	MP (°C)	BP (°C)
methanol	-98	65	methanoic acid	8	101
ethanol	-114	78	ethanoic acid	17	118
1-propanol	-126	97	propanoic acid	-21	141
1-butanol	-90	118	butanoic acid	-5	163
1-pentanol	-79	138	pentanoic acid	-34	186
1-hexanol	-52	157	hexanoic acid	-3	205
1-heptanol	-34	176	heptanoic acid	-8	223
1-octanol	-16	195	octanoic acid	16	239

State and explain the trends in melting and boiling points of both alkanols and alkanolic acids.

As the length of the carbon chain increases, for both alkanols and alkanolic acids the boiling points increase. This trend is explained by the greater number of atoms and greater weight (and hence greater dispersion forces) as the carbon chain increases.

The melting points do not follow the same pattern. Methanol and ethanol have higher melting points than propanol, even though the chain length is shorter. This is because of the strength of hydrogen bonding between the small molecules in the solid state. As the carbon chain becomes longer and hence more dominant, the effect of the hydrogen bonding is cancelled out by the impact of the non-polar carbon chain.

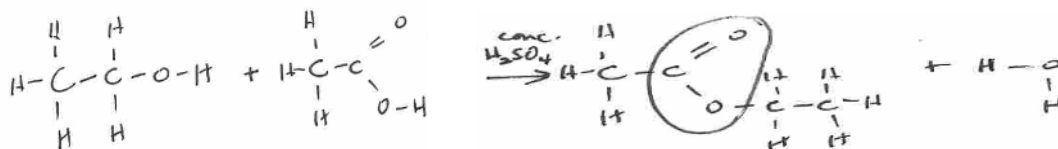
In acids, the melting points show no definite trend as the carbon chain length changes. There would be a relationship between the shape of the molecules and the way they are packed together in the solid, which would explain the melting point pattern.

Acids have higher melting and boiling points than the corresponding alcohols. This can be explained by the greater strength of hydrogen bonding in acids than in alcohols. The geometry of the carboxylic acid group allows 2 hydrogen bonds per pair of molecules whereas the geometry of the hydroxyl groups in alcohols only allows 1 hydrogen bond per pair of molecules. Of less significance is the higher molecular weight of the acid.

How would the melting and boiling points of the corresponding alkanes compare with the data shown in the table above?

Alkanes will have lower melting and boiling points than the corresponding alcohols or acids.

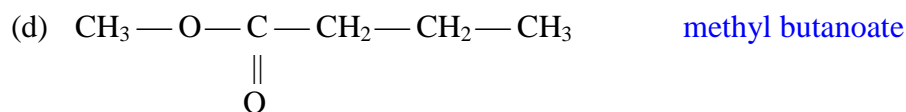
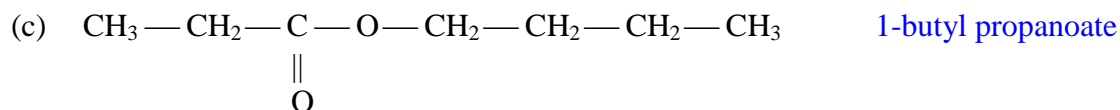
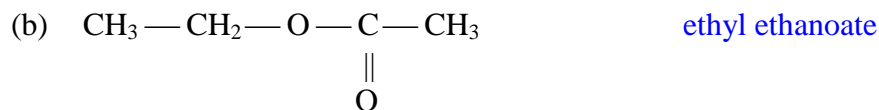
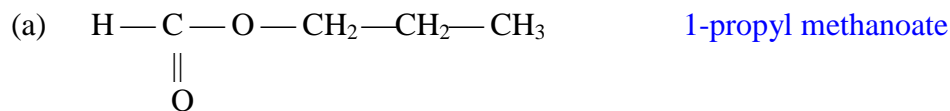
9. Write the chemical equation, using full structural formulae for the reaction between ethanol and ethanoic acid. Name the ester and draw a circle around the ester structural group (common to all esters).



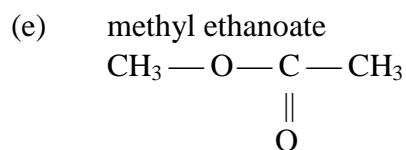
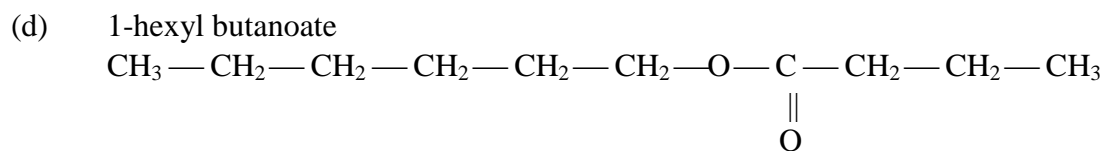
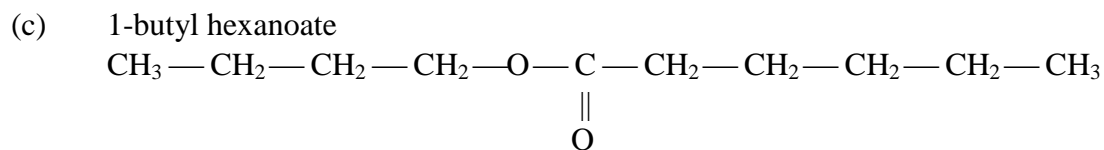
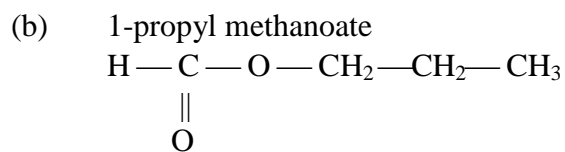
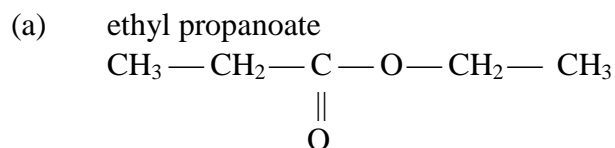
Ester is ethyl ethanoate.



10. Name the following molecules.



11. Draw the structural formula for:

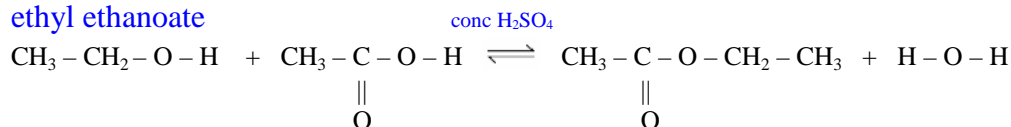


12. The table below shows some esters used in processed foods and cosmetics.

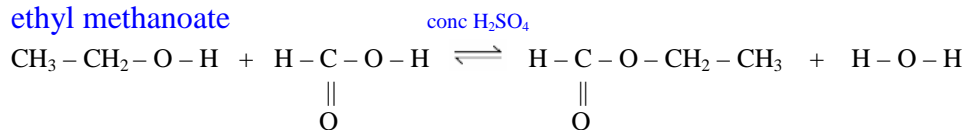
Smell	Ester
banana	pentyl ethanoate
orange	octyl ethanoate
pear	propyl ethanoate
rum	ethyl methanoate
nail lacquers	butyl ethanoate
nail polish remover	ethyl ethanoate

Write chemical equations, showing full structural formulae, for at least 3 of the reactions to form the above esters.

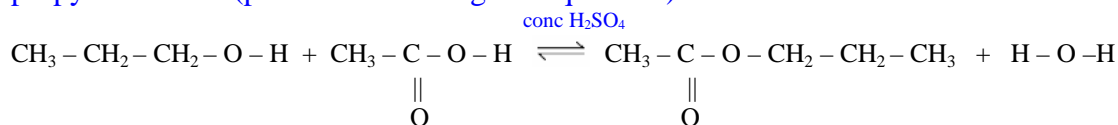
ethyl ethanoate



ethyl methanoate



propyl ethanoate (pear – Note: changed in question)



13. By considering the polarity of esters and the intermolecular forces involved, explain why esters have much lower MP and BP than alkanolic acids and alkanols of similar molecular weight.

Esters are polar molecules (the oxygen atoms are electronegative and the overall molecules are not symmetrical), so esters are held together by dipole-dipole forces. However, these are not as strong as the hydrogen bonds which attract neighbouring alcohols or acids. (Note that esters do not have hydrogen attached directly to oxygen atoms.) Stronger intermolecular forces result in higher melting points and boiling points.

14. Explain why esters are insoluble in water (or only slightly soluble in the case of the lowest molecular weight esters), whereas alcohols and alkanolic acids of low molecular weight are soluble in water.

Since esters cannot form hydrogen bonds they do not dissolve in water. (Note that esters do not have hydrogen attached directly to oxygen atoms.) Alcohols and acids can form hydrogen bonds with water molecules, so the low molecular weight acids and alcohols dissolve in water. As the carbon chain gets longer, the total attraction due to hydrogen bonding is not great enough to overcome the repulsive forces between the hydrocarbon chains and water.

15. Explain the process used in the laboratory to prepare ethyl propanoate. Refer to the use of reflux and explain why concentrated sulfuric acid is added to the reaction mixture.

Ethanol, propanoic acid and the catalyst, concentrated sulfuric acid, are placed in a flask attached to a reflux condenser.

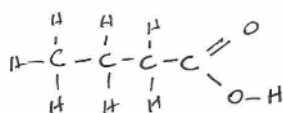
The mixture is heated for an hour, with the flame height and the water flow rate through the condenser controlled to ensure that no gases escape into the laboratory. Boiling chips in the flask ensure even boiling.

After refluxing for an hour, the mixture is cooled and the reaction stopped by addition of sodium carbonate (to remove the catalyst) and dried (to remove water) by addition of magnesium sulfate crystals (a drying agent). This removal of water prevents the backwards reaction (hydrolysis of the ester) occurring.

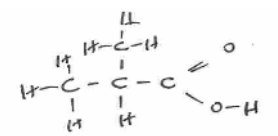
A separating funnel is used to separate the aqueous and non-aqueous layers. Esters are insoluble in water (apart from low molecular weight esters like methyl methanoate), so the ester will be in the upper layer in the separating funnel. The bottom layer is discarded.

The upper layer is distilled (or fractionally distilled) and the fraction with the boiling point of the ester is collected.

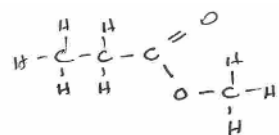
16. Draw structural formulae to represent all the compounds of molecular formula  $C_4H_8O_2$ . Name each molecule and state the homologous series to which each isomer belongs.



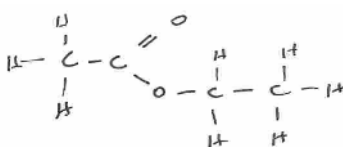
butanoic acid  
acid



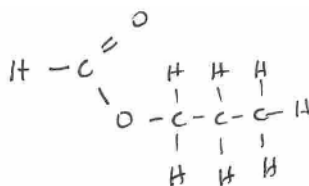
2-methylpropanoic acid  
acid



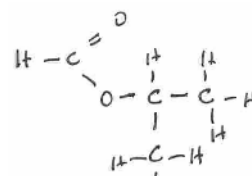
methyl propanoate  
ester



ethyl ethanoate  
ester

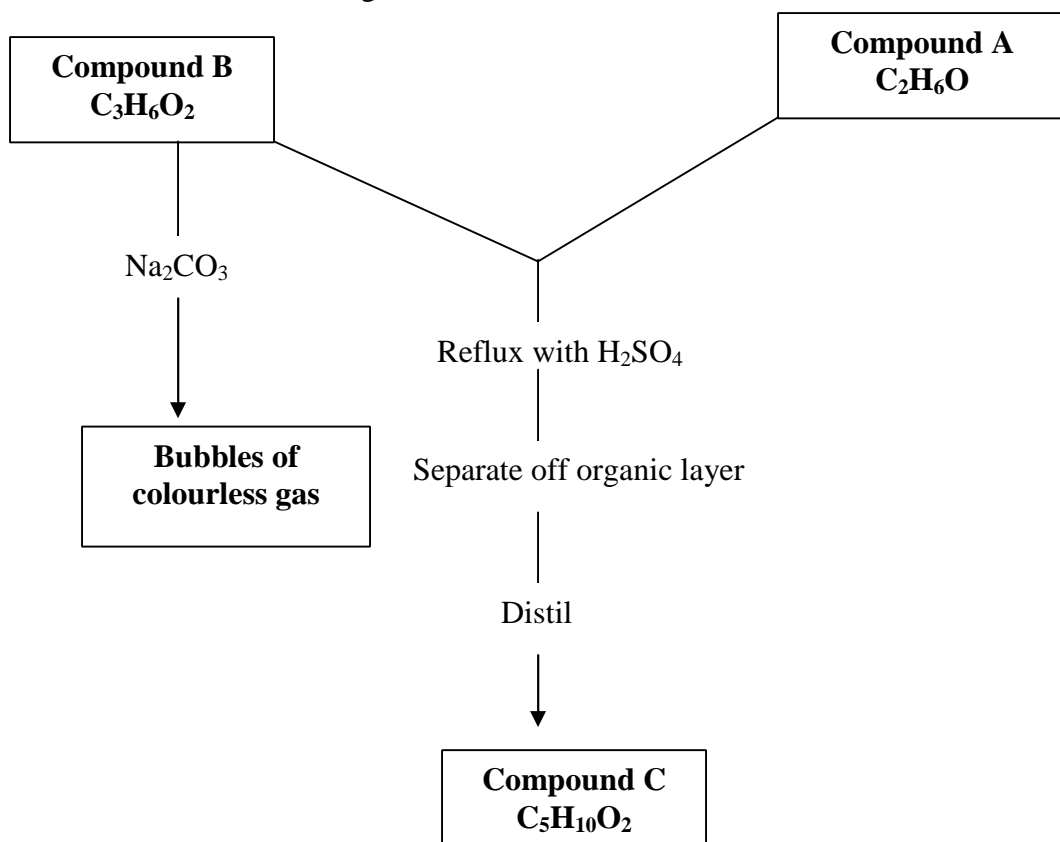


1-propyl methanoate  
ester



2-propyl methanoate (ester)

17. Consider the following reaction scheme:

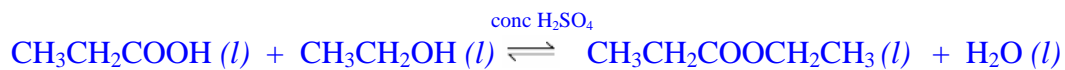


Compound A is an alkanol. Identify B and C and name them.

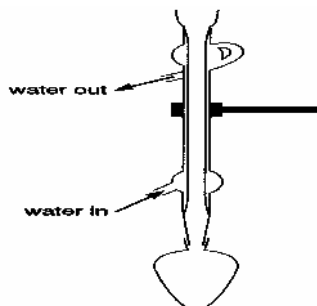
B = propanoic acid,  $CH_3CH_2COOH$

C = ethyl propanoate,  $CH_3CH_2COOCH_2CH_3$

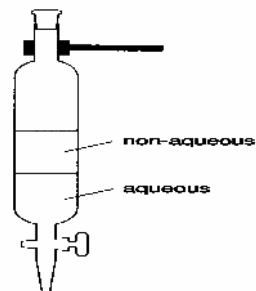
Write equations for the two reactions involved.



18. The following pieces of apparatus are used in the preparation of an ester. Name each piece of apparatus and explain how and why each piece of apparatus is used.



Reflux apparatus (flask with condenser attached)



Separating funnel

Reflux apparatus is used so that ester can be safely prepared and collected. Without a condenser to trap the volatile gases and return them to the reaction flask, the product ester (as well as the reactant alcohol and acid) could escape into the room. As a result, little or no ester would be prepared. The method ensures safe preparation (no flammable or irritating vapours are released). It also ensures that heat is constantly applied, as esterification is an endothermic reaction.

A separating funnel is used after refluxing to separate the aqueous and non-aqueous layers. Esters are normally insoluble in water (apart from low molecular weight esters like methyl methanoate), so the ester will be in the upper layer in the separating funnel. The bottom layer is discarded. The ester can then be separated from any reactants or products in the non-aqueous layer by distillation (or fractional distillation).

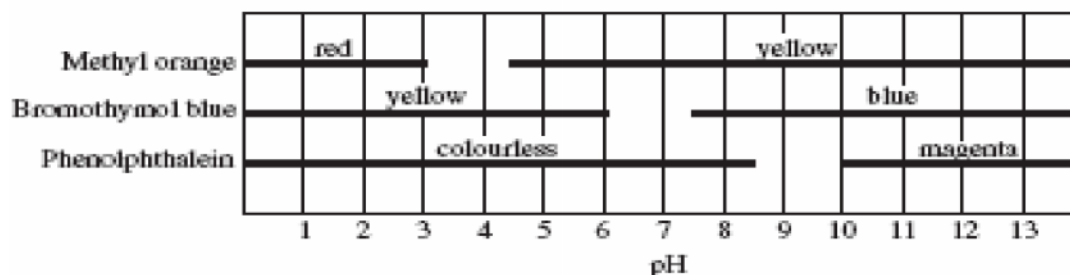
## Other Questions relating to Esterification

(Note: these questions are not necessarily of HSC standard but will be useful to test your knowledge and understanding of this topic)

- Which of the following compounds has the lowest boiling point?  
(A) methane  
(B) methanoic acid  
(C) methanol  
(D) ethane
- Which of the following compounds is not influenced by intermolecular “hydrogen bonding”?  
(A) ethanol  
(B) propanoic acid  
(C) methanoic acid  
(D) propane
- A catalyst  
(A) slows down reactions so they are safer  
(B) speeds up the reaction and is used up in the process  
(C) speeds up the reaction but is not part of the products at the end  
(D) is always a solid like iron or platinum
- The catalyst used to produce esters is concentrated sulfuric acid. An advantage of using this acid (apart from its role as a catalyst) is to  
(A) break down the C-O bonds in the reaction  
(B) increase the temperature of the reaction flask  
(C) absorb the water produced in the reaction  
(D) assist with the process of refluxing
- The products formed from the reaction of 1-propanol and butanoic acid, using concentrated sulfuric acid as a catalyst, are:  
(A) butyl propanoate and water  
(B) 1-propyl butanoate and hydrogen  
(C) 1-propyl butanoate and water  
(D) butyl propanoate and hydrogen
- The ester ethyl methanoate is formed from:  
(A) ethanol and methanoic acid  
(B) methanol and ethanoic acid  
(C) ethanol and ethanoic acid  
(D) ethylene and methanoic acid

## Questions relating to the Acidic Environment (similar to past HSC questions)

1. The graph shows the colour ranges of the acid–base indicators methyl orange, bromothymol blue and phenolphthalein.



A solution is yellow in methyl orange, yellow in bromothymol blue and colourless in phenolphthalein.

What is the pH range of the solution?

- (A) 4.5 to 6.0  
 (B) 6.0 to 7.5  
 (C) 7.5 to 8.5  
 (D) 8.5 to 10.0
2. A group of students produced a red solution by boiling red cabbage leaves in water. When dilute sodium hydroxide was added to the solution, it turned purple. When dilute hydrochloric acid was added to the red solution, no colour change occurred.

Which of these substances, when added, is most likely to cause the red solution to change colour?

- (A) Baking soda  
 (B) Concentrated hydrochloric acid  
 (C) Orange juice  
 (D) Vinegar
3. Sulfur burns in air as shown below:
- $$\text{S (s)} + \text{O}_2 \text{ (g)} \rightarrow \text{SO}_2 \text{ (g)}$$

What volume of sulfur dioxide gas will be released at 25°C and 100 kPa when 8.00 g of sulfur is burnt?

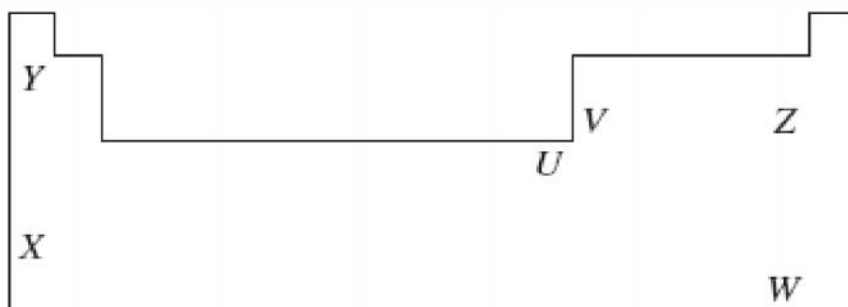
- (A) 3.10 L  
 (B) 6.20 L  
 (C) 12.40 L  
 (D) 24.79 L
4. The conjugate base of  $\text{HCO}_3^-$  is
- (A)  $\text{H}_2\text{CO}_3$   
 (B)  $\text{CO}_3^-$   
 (C)  $\text{CO}_3^{2-}$   
 (D)  $\text{HCO}_3^{2-}$

5. Le Chatelier's Principle can be used to predict the  
(A) identity of products of a chemical reaction  
(B) effect of changes in temperature on the rates of reactions  
(C) effect of catalysts on the position of equilibrium reactions  
(D) effect of changes in the concentration of chemical substances in equilibrium
6. Ozone reacts with nitric oxide according to the equation:  
 $\text{NO (g)} + \text{O}_3\text{(g)} \rightarrow \text{NO}_2\text{(g)} + \text{O}_2\text{(g)}$   
0.99 g NO (g) was mixed with 1.08 g O<sub>3</sub> (g).  
What is the maximum volume of NO<sub>2</sub> (g) produced at 0°C and 100 kPa?  
(A) 0.51 L  
(B) 0.56 L  
(C) 0.68 L  
(D) 0.75 L
7. Phosgene is prepared from the reaction of carbon monoxide and chlorine in the presence of a catalyst:  
 $\text{CO (g)} + \text{Cl}_2\text{(g)} \rightleftharpoons \text{COCl}_2\text{(g)} \quad \Delta H = -9.93 \text{ kJ mol}^{-1}$   
Which of the following sets of conditions would produce the smallest yield of phosgene?  
(A) High temperature, high pressure  
(B) Low temperature, low pressure  
(C) Low temperature, high pressure  
(D) High temperature, low pressure
8. Which of the following is a common use for ethyl pentanoate?  
(A) Flavouring  
(B) Fuel  
(C) Indicator  
(D) Solvent
9. A series of observations was made about a first-hand investigation.
- A cup of carbonated soft drink was poured from a newly-opened, chilled bottle.
  - Many bubbles were observed in the cup.
  - The cup was then left on the kitchen bench for several hours.
  - After this time, it was observed that there were no bubbles visible in the cup.
- Why were there no bubbles visible in the cup after several hours?  
(A) The carbon dioxide had dissolved into the soft drink as the temperature increased.  
(B) The carbon dioxide had been neutralised by the water in the soft drink.  
(C) Carbon dioxide is less soluble at higher temperatures and lower pressures.  
(D) Carbon dioxide is more soluble at higher temperatures and lower pressures.



10. What would be the pH of a  $0.01 \text{ mol L}^{-1}$  solution of sulfuric acid?
- (A) Less than 2.0  
 (B) Exactly 2.0  
 (C) Between 2.0 and 7.0  
 (D) Greater than 7.0
11. Which of the following pairs would form a buffer solution?
- (A)  $\text{HCl (aq)}/\text{Cl}^- \text{ (aq)}$   
 (B)  $\text{HNO}_3 \text{ (aq)}/\text{NaOH (aq)}$   
 (C)  $\text{CH}_3\text{COOH (aq)}/\text{CH}_3\text{COO}^- \text{ (aq)}$   
 (D)  $\text{H}_2\text{SO}_4 \text{ (aq)}/\text{HSO}_4^- \text{ (aq)}$

The diagram, which applies to questions 12 and 13 below, represents the Periodic Table. The positions of six different elements are shown.

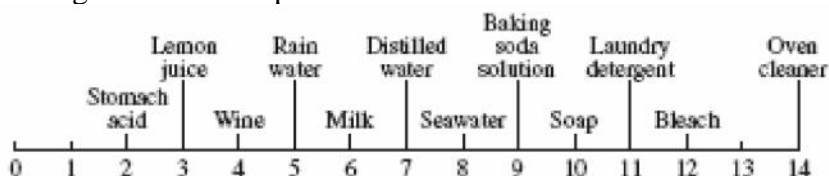


12. The elements which form the most basic and most acidic oxides, respectively, are:
- (A) X and Y  
 (B) X and Z  
 (C) Y and W  
 (D) U and V
13. The elements most likely to form amphoteric oxides are:
- (A) X and Y  
 (B) X and Z  
 (C) Y and W  
 (D) U and V
14. A sulfuric acid solution has a concentration of  $5 \times 10^{-3} \text{ mol L}^{-1}$ . Assuming complete ionisation, the pH of the solution is:
- (A) 4.0  
 (B) 2.0  
 (C) 2.3  
 (D) 2.6

15. Which statement best represents Lavoisier's definition of an acid?

- (A) Acids contain oxygen.
- (B) Acids are proton donors.
- (C) Acids contain replaceable hydrogen.
- (D) Acids ionise in solution to form hydrogen ions.

16. The figure shows the pH values of some substances.



Based on the pH values shown in the figure, which of the following statements about the concentration of hydrogen ions is correct?

- (A) It is twice as great in milk as that in lemon juice.
- (B) It is 10000 times greater in soap than in milk.
- (C) It is three times greater in wine than in bleach solution.
- (D) It is 100 times greater in distilled water than in baking soda solution.

17. A student wished to determine whether 3 solutions were acidic or basic and tested each solution with the following indicators:

3

Indicator	Colour change	pH range
Methyl orange	red to yellow	3.2–4.4
Methyl red	red to yellow	4.8–6.0
Thymol blue	yellow to blue	8.0–9.6
Alizarin	red to purple	11.0–12.4

His results were as follows:

Indicator added	Colour of solution A	Colour of solution B	Colour of solution C
Methyl orange	yellow	yellow	yellow
Methyl red	yellow	yellow	yellow
Thymol blue	blue	blue	yellow
Alizarin	purple	red	red

What conclusions can he draw about the acid/base nature of each solution?

A is highly basic (pH >12.4), B is weaker than A but still basic (pH >9.6 but <11.0), C has a pH between 6.0 and 8.0 (weakly acidic, through neutral, to weakly basic).

18. (a) Outline the procedure you would use to prepare a standard solution of sodium hydrogen carbonate from solid sodium hydrogen carbonate. 3
- Work out the molar mass of sodium hydrogen carbonate.  
Decide the volume of the solution (500 mL, 100 mL) you want to prepare.  
Calculate the mass of sodium hydrogen carbonate you need to weigh out.  
Weigh it out accurately to 4 significant figures (if the balance allows).  
Clean a volumetric flask of the correct size.  
Carefully transfer the solid to the volumetric flask, using a funnel and rod and washing any grains left in the funnel into the volumetric flask with distilled water.  
Add distilled water, swirl to dissolve the salt, add more distilled water until the bottom of the meniscus is aligned with the calibration mark.  
Stopper the volumetric flask and label it.
- (b) Calculate the mass of solid sodium hydrogen carbonate required to make 500 mL of 0.015 mol L<sup>-1</sup> solution. 1
- Molar mass NaHCO<sub>3</sub> = 84.01 g  
No. of moles needed = cV = 0.015 x 0.500 mol = 0.0075 mol  
Mass of NaHCO<sub>3</sub> needed = 0.0075 x 84.01 g = 0.63 g (2 significant figures only as 0.015 has only 2 significant figures)
19. (a) Define the term amphoteric. 1
- A substance is amphoteric if it is able both to donate a proton and to accept a proton.
- (b) Write TWO chemical equations to show that the dihydrogen phosphate ion (H<sub>2</sub>PO<sub>4</sub><sup>-</sup>) is amphoteric. 2
- Acting as an acid:  $\text{H}_2\text{PO}_4^- (\text{aq}) + \text{OH}^- (\text{aq}) \rightleftharpoons \text{HPO}_4^{2-} (\text{aq}) + \text{H}_2\text{O} (\text{l})$   
Acting as a base:  $\text{H}_2\text{PO}_4^- (\text{aq}) + \text{H}_3\text{O}^+ (\text{aq}) \rightleftharpoons \text{H}_3\text{PO}_4 (\text{aq}) + \text{H}_2\text{O} (\text{l})$
20. Justify the use of reflux for the preparation of esters in a school laboratory. 3

**Justification.**

The method of reflux is used so that ester can be safely prepared and collected.

Without a condenser to trap the volatile gases and return them to the reaction flask, the product ester (as well as the reactant alcohol and acid) could escape into the room. As a result, little or no ester would be prepared. It also ensures that heat is constantly applied, as esterification is an endothermic reaction and requires constant input of energy.

The method ensures safe preparation, as no flammable or irritating vapours are released.

21. A natural indicator was made by heating red flowers with water. Describe the steps you would take to determine the usefulness of the resulting red solution as an acid–base indicator.

3

A useful indicator must be a weak acid or base and must change colour sharply over a definite and narrow pH range.

The red solution should be tested with water, a strong acid (HCl) and a strong base (NaOH). If the solution is to be used as an indicator, it must have a different colour in acid and base and must change colour over a narrow pH range.

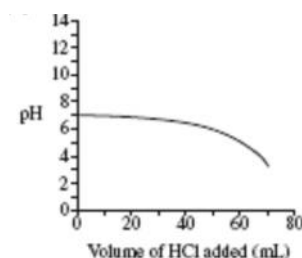
Once the liquid has proved to be useful in differentiating a strong acid from a strong base, the acid and base should be diluted systematically to produce a range of test solutions covering the pH range from (say) 2 to 12.

The coloured solution should then be added to each of these to determine the pH range when the colour change occurs. Once this is determined, more test solutions within this range can be prepared to more accurately determine the sensitivity of the indicator.

22. Dilute HCl was added to 100 mL of dilute natural buffer solution with an initial pH of 7.0 and the change in pH was plotted against the volume of HCl added.

Explain the shape of the graph.

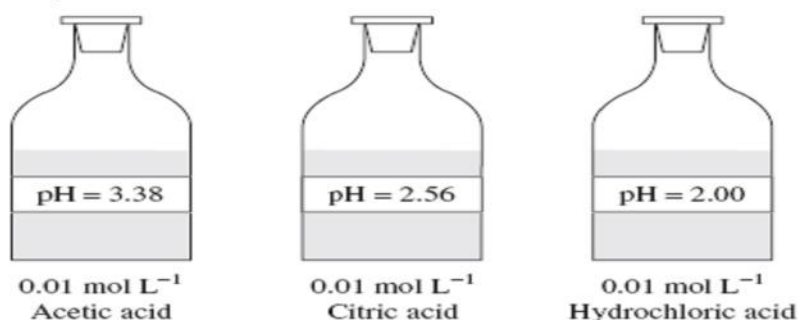
2



The pH of the buffer solution does not change significantly from 7, despite the addition of the strong acid, until 30–40 mL of the acid has been added. This is because the buffer contains equimolar (or close to equimolar) amounts of a moderate acid and its conjugate base (itself a moderate base).

On addition of the hydrochloric acid, the base in the buffer mixture will react with the added hydronium ions, to form its conjugate acid (which is only of moderate strength and hence will not ionise to produce many hydronium ions). As a result, the pH will not change significantly. However, once all the base in the buffer solution has reacted, the system cannot act as a buffer and the pH falls as more hydrochloric acid is added.

23. The diagram shows THREE bottles of different acids.



- (a) Explain why the 3 acids have different pH values. 2
- The solutions all have the same concentration. However, the strong acid, hydrochloric acid is completely ionised and hence has a  $\text{pH} = -\log_{10}[\text{H}^+] = 2$ .  
 The citric acid and acetic acid are weak acids, with the citric acid being about 8% ionised and the acetic acid only 1% ionised.  
 As a result, the citric acid produces more hydrogen ions than the acetic acid and its pH is lower, but not as low as the pH of the strong acid.

- (b) A student was asked to make exactly 100 mL of a solution of an acid of  $\text{pH} = 5.00$ , using ONE of the above acids. Which acid would she choose? Outline the steps she would take to prepare the required solution. 3

She would use hydrochloric acid, as it ionises completely.

She needs to make a solution with  $[\text{H}^+] = 10^{-5} \text{ mol/L}$ , from a solution currently at  $10^{-2} \text{ mol/L}$ . She needs to dilute by a factor of 1000. She should take 1.00 mL from the container of hydrochloric acid (0.01 mol/L) using a 1.00 mL pipette and transfer it to a 1000 mL volumetric flask and make up to the calibration mark using distilled water.

24. 25.0 mL of  $0.12 \text{ mol L}^{-1}$  standard barium hydroxide solution was titrated with nitric acid. The results were recorded in the table.

Titration	Volume of nitric acid used (mL)
1	20.4
2	18.1
3	18.2
4	18.1

Calculate the concentration of the nitric acid solution.

Use 18.13 mL as the volume of nitric acid. Discard the rough reading.

Moles  $\text{Ba}(\text{OH})_2 = cV = 0.12 \times 0.025 = 0.0030 \text{ mol}$

Moles  $\text{HNO}_3 = 2 \times 0.0030$  (as mole ratio in reaction is 2 moles  $\text{HNO}_3$  to 1 mole of  $\text{Ba}(\text{OH})_2$ )

Moles  $\text{HNO}_3 = 0.0060 \text{ mol}$

$[\text{HNO}_3] = 0.0060 / 0.01813 \text{ mol/L} = 0.331 \text{ mol/L}$  or  $0.33 \text{ mol/L}$  (to 2 s.f.) (since concentration of barium hydroxide only given to 2 s.f.)

25. To determine the mass of calcium carbonate ( $\text{CaCO}_3$ ) in an antacid tablet, the following procedure was used:

- The tablet was crushed and then placed in a beaker.
- A pipette was used to add 25.0 mL of 0.100 mol  $\text{L}^{-1}$  hydrochloric acid to the crushed tablet in the beaker.
- Once the reaction between the calcium carbonate and hydrochloric acid had stopped, phenolphthalein indicator was added to the reaction mixture.
- A burette was then used to add 0.050 mol  $\text{L}^{-1}$  sodium hydroxide to the beaker to neutralise the excess hydrochloric acid.
- The phenolphthalein changed from colourless to pink after 26.0 mL of the sodium hydroxide solution had been added.

(a) Write a balanced chemical equation for the reaction that occurred between the calcium carbonate in the tablet and the hydrochloric acid. 1



(b) How many moles of hydrochloric acid were added to the tablet? 1

$$\text{Moles HCl added to the tablet} = cV = 0.100 \times 0.0250 \text{ mol} = 0.00250 \text{ mol}$$

(c) Explain why the solution was colourless after the phenolphthalein was first added.  
The solution had an excess of hydrochloric acid and phenolphthalein is colourless in acidic solutions (in any solution below pH of 8). 1

(d) Explain why the colour of the solution changed to pink after 26.0 mL of the sodium hydroxide solution had been added. 1  
The excess hydrochloric acid had been just neutralised by the sodium hydroxide. With one extra drop, the solution was now alkaline. Hence the phenolphthalein turned pink.

(e) Calculate the mass of calcium carbonate in the original antacid tablet. 3

$$\begin{aligned} \text{Moles of NaOH required to neutralise the excess acid} &= 0.050 \times 0.0260 = 0.0013 \text{ mol} \\ \text{Hence moles of HCl used to react with tablet} &= 0.0025 - 0.0013 = 0.0012 \text{ mol} \end{aligned}$$

Since 2 moles of HCl react with 1 mole of  $\text{CaCO}_3$  (molar mass 100 g)

$$\begin{aligned} \text{Moles CaCO}_3 \text{ in original tablet} &= 0.5 \times 0.0012 = 0.00060 \text{ mol} &= 0.00060 \times 100 \text{ g} \\ & &= 0.060 \text{ g (2 s.f. only)} \end{aligned}$$

since  $[\text{NaOH}]$  was only given to 2 s.f.)

26. The ester, ethyl butanoate, is to be prepared in the school laboratory.

- (a) Draw the structural formula for ethyl butanoate.

1



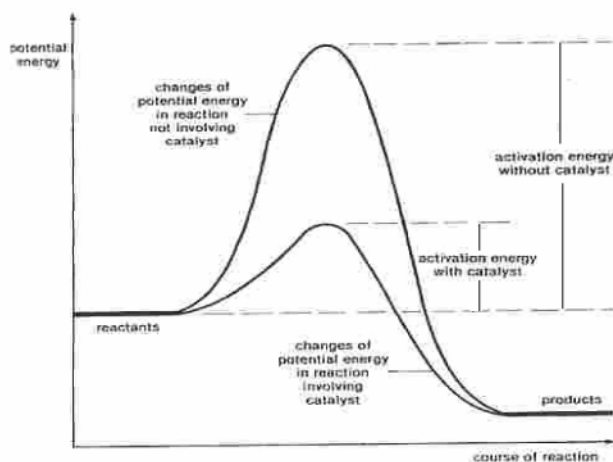
- (b) Concentrated sulfuric acid is used as a catalyst in this reaction. Explain why a catalyst is needed and use a diagram to show the effect of a catalyst on an endothermic reaction such as esterification.

3

A catalyst is used to speed up the equilibrium reactions (both forward and reverse).

As a result, the equilibrium is reached faster.

All catalysts work by providing an alternate pathway for the reaction, which has lower activation energy than the non-catalysed reaction.



- (c) Name the TWO other chemicals needed for the preparation of ethyl butanoate.

2

Ethanol and butanoic acid.

- (d) Three steps are needed to prepare and isolate the ester. The first step involves refluxing. The second step involves the use of a separating funnel. Justify the need for this second step in the laboratory process.

2

The products and unchanged reactants form 2 immiscible layers, which need to be separated from each other. The ester is normally in the upper, non-aqueous layer.

While the mixture is in the separating funnel, sodium carbonate can be added to remove any acid remaining in the mixture (sulfuric acid, the catalyst and any unreacted butanoic acid). Water can also be removed from the reaction mixture by adding magnesium sulfate or calcium chloride (these salts act as drying agents). This stops the reverse reaction happening.

The lower aqueous layer is discarded, leaving the ester and any non-aqueous impurities in the upper layer.

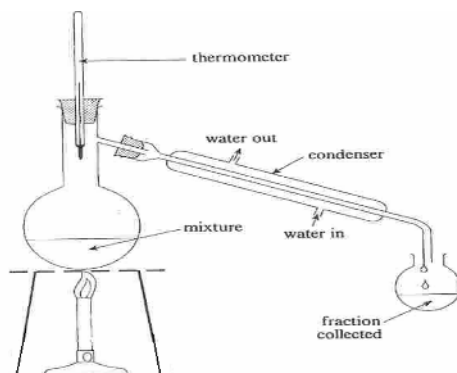
- (e) Distillation is used to separate the ester from the other substances still present in the mixture. Explain why the ester is likely to have the lowest boiling point of the substances in the mixture.

Esters are polar molecules but do not hydrogen bond. As a result, the intermolecular forces are weaker than those in the alcohol or acid which could be present in the non-aqueous layer and esters would have lower BP than the alcohols or acids.

2

(f) Draw a diagram of the distillation equipment used.

2



(g) There are several isomers of ethyl butanoate.

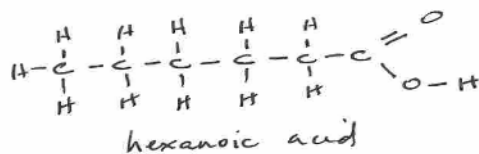
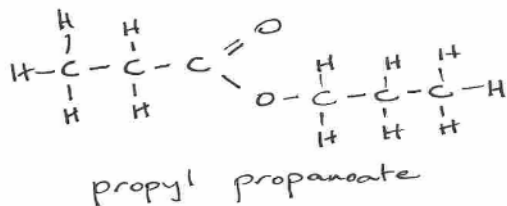
(i) Explain what is meant by this statement.

1

There are several compounds which have the same molecular formulae as ethyl butanoate but have different structural formulae.

(ii) Draw the structural formulae of 2 different isomers of ethyl butanoate. ONE of these compounds should be another ester and ONE should be a member of a different homologous series of compounds.

3





**Answers to Multiple Choice Questions**

Question No.	Sect 1 Page 10	Sect 2 Page 17	Sect 2 Page 20	Sect 3 Page 29	Sect 3 Page 34	Sect 4 Page 43	Sect 5 Page 62	Overall Page 63
1.	C	A	D	A	A	C	A	A
2.	D	C	C	A	D	D	D	A
3.		B	B	D	D	A	C	B
4.		A	C	C	D	C	C	C
5.		A	C		A	D	C	D
6.		C	D		A	D	A	A
7.		B			D	D		D
8.		B				B		A
9.		A				C		C
10.		B				B		A
11.		A				C		C
12.						D		B
13.						C		D
14.						D		B
15.								A
16.								D

## DATA SHEET

Avogadro constant, $N_A$ .....	$6.022 \times 10^{23} \text{ mol}^{-1}$
Volume of 1 mole ideal gas: at 100 kPa and	
at 0°C (273.15 K) .....	22.71 L
at 25°C (298.15 K) .....	24.79 L
Ionisation constant for water at 25°C (298.15 K), $K_w$ .....	$1.0 \times 10^{-14}$
Specific heat capacity of water .....	$4.18 \times 10^3 \text{ J kg}^{-1} \text{ K}^{-1}$

### Some useful formulae

$$\text{pH} = -\log_{10}[\text{H}^+] \qquad \Delta H = -m C \Delta T$$

### Some standard potentials

$\text{K}^+ + \text{e}^-$	$\rightleftharpoons$	$\text{K(s)}$	-2.94 V
$\text{Ba}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Ba(s)}$	-2.91 V
$\text{Ca}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Ca(s)}$	-2.87 V
$\text{Na}^+ + \text{e}^-$	$\rightleftharpoons$	$\text{Na(s)}$	-2.71 V
$\text{Mg}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Mg(s)}$	-2.36 V
$\text{Al}^{3+} + 3\text{e}^-$	$\rightleftharpoons$	$\text{Al(s)}$	-1.68 V
$\text{Mn}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Mn(s)}$	-1.18 V
$\text{H}_2\text{O} + \text{e}^-$	$\rightleftharpoons$	$\frac{1}{2}\text{H}_2(\text{g}) + \text{OH}^-$	-0.83 V
$\text{Zn}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Zn(s)}$	-0.76 V
$\text{Fe}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Fe(s)}$	-0.44 V
$\text{Ni}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Ni(s)}$	-0.24 V
$\text{Sn}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Sn(s)}$	-0.14 V
$\text{Pb}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Pb(s)}$	-0.13 V
$\text{H}^+ + \text{e}^-$	$\rightleftharpoons$	$\frac{1}{2}\text{H}_2(\text{g})$	0.00 V
$\text{SO}_4^{2-} + 4\text{H}^+ + 2\text{e}^-$	$\rightleftharpoons$	$\text{SO}_2(\text{aq}) + 2\text{H}_2\text{O}$	0.16 V
$\text{Cu}^{2+} + 2\text{e}^-$	$\rightleftharpoons$	$\text{Cu(s)}$	0.34 V
$\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O} + 2\text{e}^-$	$\rightleftharpoons$	$2\text{OH}^-$	0.40 V
$\text{Cu}^+ + \text{e}^-$	$\rightleftharpoons$	$\text{Cu(s)}$	0.52 V
$\frac{1}{2}\text{I}_2(\text{s}) + \text{e}^-$	$\rightleftharpoons$	$\text{I}^-$	0.54 V
$\frac{1}{2}\text{I}_2(\text{aq}) + \text{e}^-$	$\rightleftharpoons$	$\text{I}^-$	0.62 V
$\text{Fe}^{3+} + \text{e}^-$	$\rightleftharpoons$	$\text{Fe}^{2+}$	0.77 V
$\text{Ag}^+ + \text{e}^-$	$\rightleftharpoons$	$\text{Ag(s)}$	0.80 V
$\frac{1}{2}\text{Br}_2(\text{l}) + \text{e}^-$	$\rightleftharpoons$	$\text{Br}^-$	1.08 V
$\frac{1}{2}\text{Br}_2(\text{aq}) + \text{e}^-$	$\rightleftharpoons$	$\text{Br}^-$	1.10 V
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+ + 2\text{e}^-$	$\rightleftharpoons$	$\text{H}_2\text{O}$	1.23 V
$\frac{1}{2}\text{Cl}_2(\text{g}) + \text{e}^-$	$\rightleftharpoons$	$\text{Cl}^-$	1.36 V
$\frac{1}{2}\text{Cr}_2\text{O}_7^{2-} + 7\text{H}^+ + 3\text{e}^-$	$\rightleftharpoons$	$\text{Cr}^{3+} + \frac{7}{2}\text{H}_2\text{O}$	1.36 V
$\frac{1}{2}\text{Cl}_2(\text{aq}) + \text{e}^-$	$\rightleftharpoons$	$\text{Cl}^-$	1.40 V
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^-$	$\rightleftharpoons$	$\text{Mn}^{2+} + 4\text{H}_2\text{O}$	1.51 V
$\frac{1}{2}\text{F}_2(\text{g}) + \text{e}^-$	$\rightleftharpoons$	$\text{F}^-$	2.89 V

Aylward and Findlay, *SI Chemical Data* (5th Edition) is the principal source of data for this examination paper. Some data may have been modified for examination purposes.

PERIODIC TABLE OF THE ELEMENTS

PERIODIC TABLE OF THE ELEMENTS																																			
1 H 1.008 Hydrogen		KEY																																	
		Atomic Number		Symbol of element		Atomic Weight		Name of element																											
3	Li 6.941 Lithium	4	Be 9.012 Beryllium	79	Au 197.0 Gold	79	Au 197.0 Gold	5	B 10.81 Boron	6	C 12.01 Carbon	7	N 14.01 Nitrogen	8	O 16.00 Oxygen	9	F 19.00 Fluorine	10	Ne 20.18 Neon																
11	Na 22.99 Sodium	12	Mg 24.31 Magnesium	25	Mn 54.94 Manganese	26	Fe 55.85 Iron	27	Co 58.93 Cobalt	28	Ni 58.69 Nickel	29	Cu 63.55 Copper	30	Zn 65.39 Zinc	31	Ga 69.72 Gallium	32	Ge 72.61 Germanium	33	As 74.92 Arsenic	34	Se 78.96 Selenium	35	Br 79.90 Bromine	36	Kr 83.80 Krypton								
37	Rb 85.47 Rubidium	38	Sr 87.62 Strontium	39	Y 88.91 Yttrium	40	Zr 91.22 Zirconium	41	Nb 92.91 Niobium	42	Mo 95.94 Molybdenum	43	Tc [98.91] Technetium	44	Ru 101.1 Ruthenium	45	Rh 102.9 Rhodium	46	Pd 106.4 Palladium	47	Ag 107.9 Silver	48	Cd 112.4 Cadmium	49	In 114.8 Indium	50	Sn 118.7 Tin	51	Sb 121.8 Antimony	52	Te 127.6 Tellurium	53	I 126.9 Iodine	54	Xe 131.3 Xenon
55	Cs 132.9 Cesium	56	Ba 137.3 Barium	57-71	Lanthanides La Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu	72	Hf 178.5 Hafnium	73	Ta 180.9 Tantalum	74	W 183.8 Tungsten	75	Re 186.2 Rhenium	76	Os 190.2 Osmium	77	Ir 192.2 Iridium	78	Pt 195.1 Platinum	79	Au 197.0 Gold	80	Hg 200.6 Mercury	81	Tl 204.4 Thallium	82	Pb 207.2 Lead	83	Bi 209.0 Bismuth	84	Po [210.0] Polonium	85	At [210.0] Astatine	86	Rn [222.0] Radon
87	Fr [223.0] Francium	88	Ra [226.0] Radium	89-103	Actinides Ac Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr	104	Rf [261.1] Rutherfordium	105	Db [262.1] Dubnium	106	Sg [263.1] Seaborgium	107	Bh [264.1] Bohrium	108	Hs [265.1] Hassium	109	Mt [268] Moscovium	110	Uun — Ununilium	111	Uuu — Ununilium	112	Uub — Unbibium	113	Uuq — Ununquadium	114	Uuh — Ununhexium	115	Uus — Ununseptium	116	Uuo — Ununoctium	117	Uus — Ununseptium	118	Uuo — Ununoctium

Lanthanides

57	La	138.9	Lanthanum	58	Ce	140.1	Cerium	59	Pr	140.9	Praseodymium	60	Nd	144.2	Neodymium	61	Pm	[146.9]	Promethium	62	Sm	150.4	Samarium	63	Eu	152.0	Europium	64	Gd	157.3	Gadolinium	65	Tb	158.9	Terbium	66	Dy	162.5	Dysprosium	67	Ho	164.9	Holmium	68	Er	167.3	Erbium	69	Tm	168.9	Thulium	70	Yb	173.0	Ytterbium	71	Lu	175.0	Lutetium
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Actinides

89	Ac	[227.0]	Actinium	90	Th	232.0	Thorium	91	Pa	231.0	Protactinium	92	U	238.0	Uranium	93	Np	[237.0]	Neptunium	94	Pu	[239.1]	Plutonium	95	Am	[241.1]	Americium	96	Cm	[244.1]	Curium	97	Bk	[249.1]	Berkelium	98	Cf	[252.1]	Californium	99	Es	[252.1]	Einsteinium	100	Fm	[257.1]	Fermium	101	Md	[258.1]	Mendelevium	102	No	[259.1]	Nobelium	103	Lr	[262.1]	Lawrencium
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Where the atomic weight is not known, the relative atomic mass of the most common radioactive isotope is shown in brackets.  
The atomic weights of Np and Tc are given for the isotopes <sup>237</sup>Np and <sup>99</sup>Tc.

# HSC CHEMISTRY

## STUDENT WORKBOOKS

### About the author:

Bronwen Hegarty is an experienced teacher of HSC Chemistry and has been the convenor of Examination Committees and Chemistry writing teams. She continues to teach Intensive Chemistry Revision courses and is working to assist Chemistry teachers and their students in a number of schools. She has recognised the need for these workbooks in schools, not only to provide revision for students throughout their Year 11 and HSC courses, but also to reduce the workload on teachers in preparing course materials and sequencing lesson plans. The workbooks are designed to be a low cost alternative to creating and photocopying teacher-designed student worksheets, assignments and revision questions. The format will facilitate student-centred teaching and learning and provide a framework for students to monitor their progress through the HSC syllabus. Teacher Editions of the workbooks include a CD and are available only to teachers, to provide them with fully worked solutions for each workbook.

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To obtain copies of these workbooks, contact Bronwen Hegarty:

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