

THE ACIDIC ENVIRONMENT

What is this topic about?

To keep it as simple as possible, (K.I.S.S.) this topic involves the study of:

1. ACIDS, BASES & INDICATORS
2. ACIDS IN OUR ENVIRONMENT
3. ACIDS & THE pH SCALE
4. ACID-BASE THEORY & TITRATION
5. ESTERIFICATION

...all in the context of Chemistry in our environment and society.

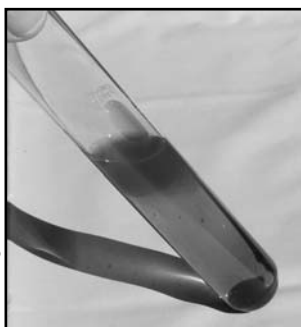
but first, an introduction...

Way back in years 8, 9 and 10 you would have studied some basic Chemistry, and it probably involved studying acids and bases and indicators... so these are familiar terms, even if you've forgotten the details.

This topic begins by reminding you of this simple way to classify all chemical substances:

**Acid,
Base,
or Neutral**

and how **"Indicators"** can be used to identify them.



Then, we look at the **Chemistry of Oxide Compounds**, and link that to **Acidity**.



Photo by Ken Kiser

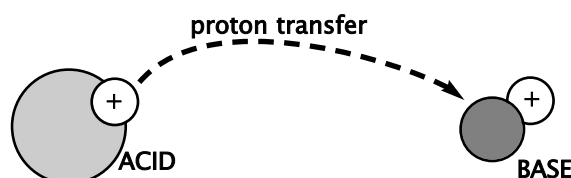
You will study the sources and problems that can result from

Acids in the Environment

and, revise & practice

Mole Calculations

The details of **Acid-Base Reactions** will be studied

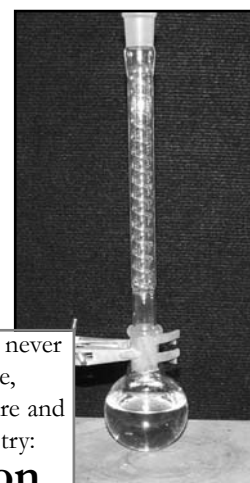


...and you will learn the real meaning of the pH scale, used for measuring acid and base strength.

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



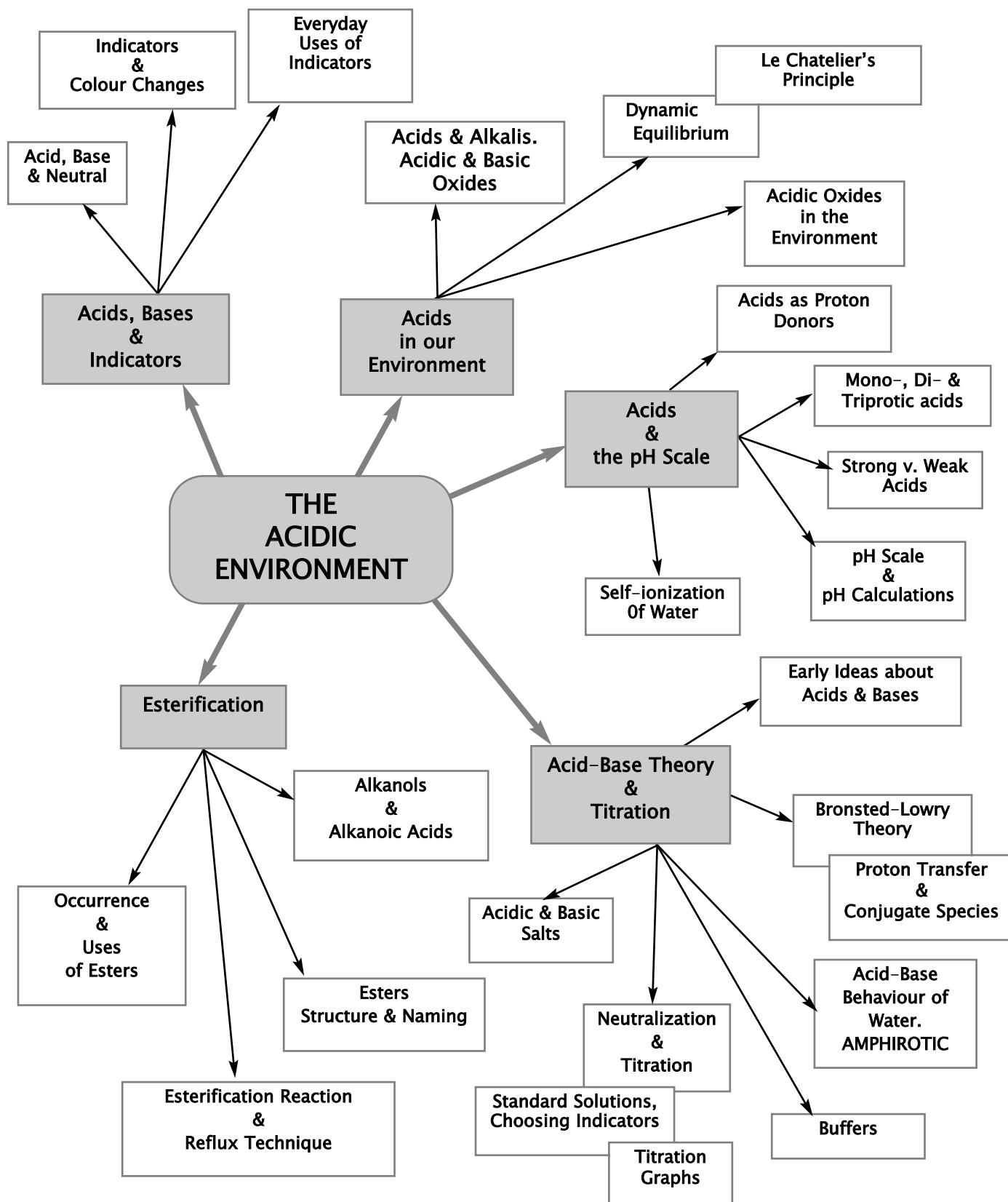
Along the way, you will learn some important techniques in **Chemical Analysis**



and a process you might never have heard of before, but it is common in nature and in the Chemical Industry:
Esterification

CONCEPT DIAGRAM ("Mind Map") OF TOPIC

Some students find that memorizing the OUTLINE of a topic helps them learn and remember the concepts and important facts. As you proceed through the topic, come back to this page regularly to see how each bit fits the whole. At the end of the notes you will find a blank version of this "Mind Map" to practise on.

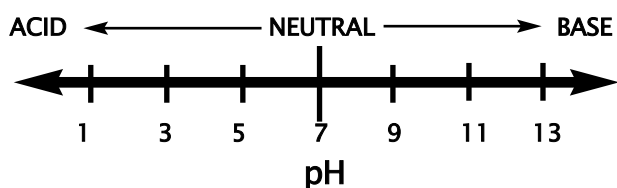


1. ACIDS, BASES & INDICATORS

Acid, Base or Neutral?

The chemical definitions of acid and base will come later. For now, you are reminded of what you may have learnt in earlier Science classes.

Acids and bases are chemical opposites; if you add one to the other they “destroy” (neutralize) each other, and the end product is neutral. Acids and bases are the opposite ends of a chemical property called “acidity”, which is measured by a numerical scale called “pH”.



The word “acid” comes from the Latin for “sour”, and refers to the fact that natural, acidic chemicals (e.g. vinegar) are sour-tasting.

Indicators

Indicators are chemicals which change colour according to the acidity of the solution they are in.

The original indicators were natural extracts from plants or other living things. Some, such as litmus, are still in use today, as well as newer, synthetically made chemicals.

Practical Work: A Natural Indicator

You may have done practical work to prepare and test a natural indicator.



A good example is the common garden plant Hydrangea.

If you collect a flower head and put it through a blender with a little water and ethanol, the filtered liquid extract will work as a simple indicator.

In acid, the Hydrangea flower extract is a bluish colour.

In a base, it turns pink-ish.

Modern Laboratory Indicators

The syllabus requires that you are familiar with the common laboratory indicators listed below.

You will have done experimental work, adding 2 drops of indicator to test tubes of acid, base and pure water (which is neutral) and recorded the colours produced.

Your results should have agreed with the following:

Indicator	Colour in		
	Acid	Water	Base
Litmus	pink	purple	blue
Phenolphthalein	clear	clear	red/pink
Methyl Orange	red	yellow	yellow
Bromothymol blue	yellow	green/blue	blue

Choosing an Indicator

Why are so many indicators needed?

Litmus is useful for general identification of acidic or basic substances. However, its colour change is rather indistinct, and can occur over quite a range of pH values... it is not a “sharp” change.

In contrast, Phenolphthalein cannot tell you the difference between a glass of water and sulfuric acid. However, the colour change is very distinctive, and occurs suddenly at a very specific pH value... it is very sharp. This is not much use in general identification of substances, but in certain methods of chemical analysis it is very important.

So, each indicator has a purpose and preferred use.

Everyday Uses of Indicators

- Soil Testing. Some plants grow best in acidic soils; others need slightly alkaline (basic) conditions. Farmers and keen gardeners use simple test kits containing an indicator and colour chart, to test the soil. They can then adjust the soil pH to get the best results.
- Water Testing. Swimming pools need regular testing for acidity to better maintain their water quality and hygiene. Aquariums must be maintained at very specific pH levels for the health of their inhabitants.
- Effluent Testing. Acidity is a useful way to assess the levels of certain types of pollution from industries. Industry technicians and Government authorities use indicators to monitor the pH of waste water and natural waterways.

Practical Work: Classifying Household Substances

You may have done laboratory work using various liquid and paper indicators to classify a range of household substances and foodstuffs as either acid, base or neutral.

By using appropriate indicators (such as “Universal”) you may have even been able to differentiate between substances that are “mildly acidic” and “strongly acidic”.



Photo by Jan Friml



A Few Typical Results

Household Substances Found to be...

ACIDIC

Vinegar
Orange juice
Tomato juice

NEUTRAL

Salt
Sugar
Shampoo
Milk
Liquid detergent

BASIC

Soap
Floor cleaner
Drain cleaner

Worksheet 1

Part A Fill in the Blanks

Acids and bases are chemical a)..... If you add one to the other, they b)..... each other. Acidity is measured by the c)..... scale. On this scale, a neutral substance has a value of d)..... Values above this indicate e)..... substances, while values below indicate f)..... substances.

Indicators are chemicals which g)..... according to the h)..... of the solution they are in. The original indicators were extracts from i).....

The colour changes for the following indicators need to be learnt.

<u>Indicator</u>	<u>Acid</u>	<u>Neutral</u>	<u>Base</u>
Litmus	j).....	purple	k).....
Phenolphthalein	l).....
Methyl orange	m).....
Bromothymol blue	n).....	green/blue	o).....

In everyday situations, indicators are used for purposes such as p)..... for farming and gardening, q)..... for pools and aquariums, and for monitoring r)..... from industries.

Part B Practice Problems

Each solution listed below has been tested with one or more indicators, and the colour is given. For each, state if it is acidic, neutral or basic.

<u>Solution A:</u>	Phenolphthalein is clear. Methyl orange is red.
<u>Solution B:</u>	Phenolphthalein is pink. Methyl orange is yellow.
<u>Solution C:</u>	Phenolphthalein is clear. Methyl orange is yellow.
<u>Solution D:</u>	Bromothymol blue is blue. Methyl orange is yellow.
<u>Solution E:</u>	Phenolphthalein is clear. Litmus is pink.

2. ACIDS IN OUR ENVIRONMENT

Laboratory Acids and “Alkalis”

You are reminded that the common laboratory acids are:

Name	Formula	Solution of
Hydrochloric acid	HCl	$\text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$
Sulfuric acid	H_2SO_4	$2\text{H}^+_{(\text{aq})} + \text{SO}_4^{2-}_{(\text{aq})}$
Nitric acid	HNO_3	$\text{H}^+_{(\text{aq})} + \text{NO}_3^-_{(\text{aq})}$

The most familiar laboratory bases are soluble hydroxide compounds which are sometimes called “alkalis”.

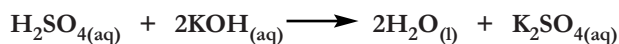
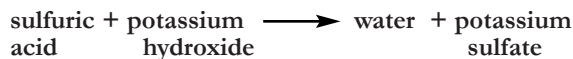
Name	Formula	Solution of
Sodium hydroxide	NaOH	$\text{Na}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
Potassium hydroxide	KOH	$\text{K}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

The Acid-Alkali Reaction: Neutralization

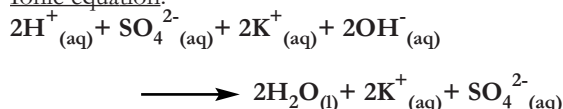
You should be familiar with this reaction from earlier Science studies:



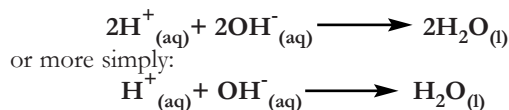
Example:



Ionic equation:



If you study this equation you will see that the potassium and sulfate ions are spectators. You can leave them out to form the net ionic equation:



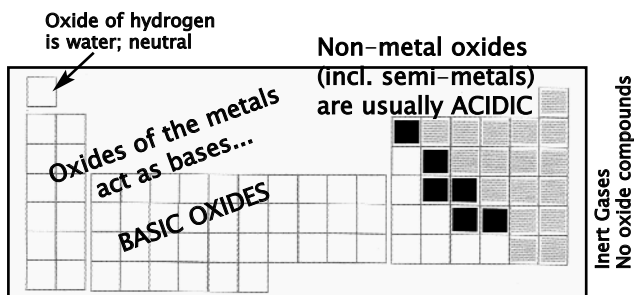
All the familiar laboratory acids are solutions containing hydrogen ions (H^+). The laboratory “alkali” bases are solutions containing hydroxide ions (OH^-). These will always react to form water, so acid and base have neutralized each other.

The other product is always a soluble, ionic compound. These are known collectively as “salts”.

Try the WORKSHEET at the end of section

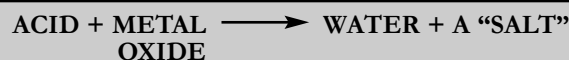
Acid-Base Properties of the Oxide Compounds

The common laboratory acids and alkali-bases are not the whole story. Most of the oxide compounds of the elements show some acid-base behaviour too, and the periodic table reveals another pattern.

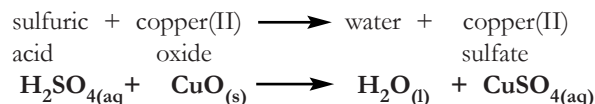
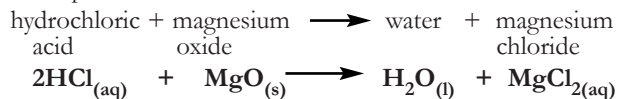


Basic Oxides of the Metals

Most of the oxides of the metallic elements are considered basic because they can neutralize acids, forming water and a “salt”.



Examples:

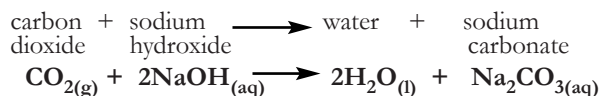
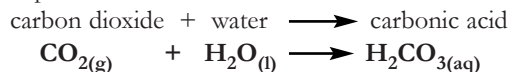


Acidic Oxides of the Non-Metals

Many of the oxide compounds of the non-metal elements are acidic because they will:

- react with water to form an acid
- and/or
- react with a base by neutralizing it, and form water and a “salt”.

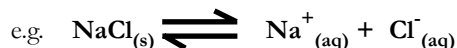
Examples:



Try the WORKSHEET at the end of section

Chemical Equilibrium

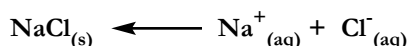
The concept of “Dynamic Equilibrium” was introduced in a previous Preliminary topic. The example used then was the equilibrium between a solid ionic lattice and dissolved ions in a saturated solution.



At equilibrium, it seems (macroscopically) that nothing is happening. However, at the atomic level, solid salt is dissolving into the solution;



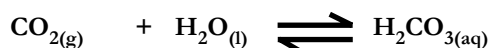
...but at the same time, some ions are precipitation out of solution to form a solid lattice;



...and these 2 processes are occurring at the same rate.

Equilibrium of Carbon Dioxide and Water

When CO_2 dissolves in water, it doesn't merely dissolve, but reacts to form the weak acid “carbonic acid”.



This is a dynamic equilibrium situation. Even when the process seems finished, there are actually 2 reactions (one forwards \longrightarrow , one backwards \longleftarrow) occurring at the same rate, so that nothing appears to be happening.

Shifting an Equilibrium

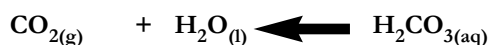
A chemical equilibrium is “dynamic”, meaning that things are moving back and forth. This also means that it is possible to “upset” an equilibrium and cause it to shift to a new balance between reactants and products.



For example, when you open and pour a fizzy soft drink it forms bubbles in the glass.

The sealed drink contains dissolved CO_2 , but as soon as it is opened, the gas begins coming out of solution, because the equilibrium has shifted to the left.

Eventually a new equilibrium is established with a lot less gas dissolved... we say it has “gone flat”.



Equilibrium shifting to the left results in less carbonic acid and the formation of more CO_2 gas. The shift in the equilibrium was caused by the pressure change which occurred when the bottle/can was opened.

Practical Work: Mass & Volume of CO_2 in a Soft Drink

You will have done a simple laboratory exercise to “de-carbonate” a fizzy soft drink and measure/calculate the mass and volume of CO_2 gas released.

If you weigh the soft drink and container accurately, then release the pressure so that it goes flat (this may take a day or more) you will have measured a small loss of mass, due to CO_2 gas escaping.



Typical Results & Analysis

For a 300mL bottle of lemonade:

Start mass = 536.9g
Final mass = 535.8g
 \therefore mass loss = 1.1g

moles of CO_2 :

$$n = m/\text{MM} \\ = 1.1/44.01$$

$$n(\text{CO}_2) = 0.025 \text{ mol.}$$

Volume of CO_2 (assuming standard conditions)

$$V = n \times V_M \\ = 0.025 \times 24.8 = 0.62 \text{ L}$$

\therefore **0.62 L (620mL) of CO_2 gas was released**

Factors That Can Cause an Equilibrium Shift

Pressure

If a reaction involves a gas, (as in the example at left) any change in the pressure of that gas will shift the equilibrium. With soft drinks, increasing the pressure of CO_2 drives the equilibrium to the right; decreasing CO_2 pressure, shifts it to the left.

Temperature

If you kept the gas pressure constant, but raised the temperature, you would find this equilibrium would shift left. (i.e. less carbonic acid) Lowering the temperature would shift it to the right (more carbonic acid).

Which way an equilibrium shifts due to temperature, depends on whether the reaction is exo- or endo-thermic. This is explained later.

Concentration

The equilibrium could also be shifted by altering the concentration of (say) the carbonic acid. If you added extra $\text{H}_2\text{CO}_{3(aq)}$ somehow, the equilibrium would shift left. If you reduced the concentration, (e.g. by adding an alkali which would react and destroy it) the equilibrium would shift towards the right.

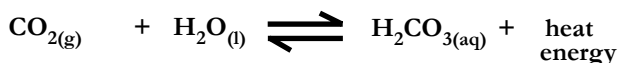
Le Chatelier's Principle

In 1885, the French Chemist Henri Le Chatelier discovered the underlying pattern in these equilibrium shifts. The equilibrium always shifts in the direction which counteracts the change that upset it in the first place.

Le Chatelier's Principle:
If a system in equilibrium is disturbed, the system will adjust itself in the direction which counteracts the disturbance

Temperature Effect on Equilibrium

The dissolving of CO_2 to form carbonic acid is an exothermic reaction, so heat energy may be considered as one of the products:



So, if the temperature is raised, the equilibrium shifts left



... because shifting left would use up heat and lower the temperature again.

If the temperature is lowered, the equilibrium shifts right



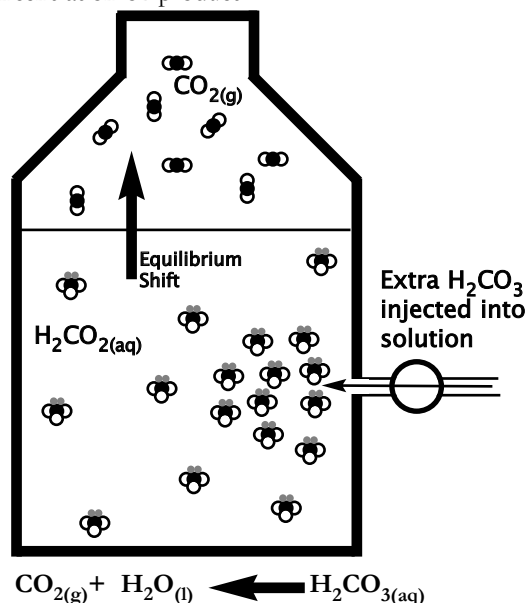
... in an attempt to release heat energy and raise the temperature again.

The Response of an Endothermic Reaction is exactly the opposite

Effect of Concentration

Imagine a sealed container with an equilibrium mixture within.

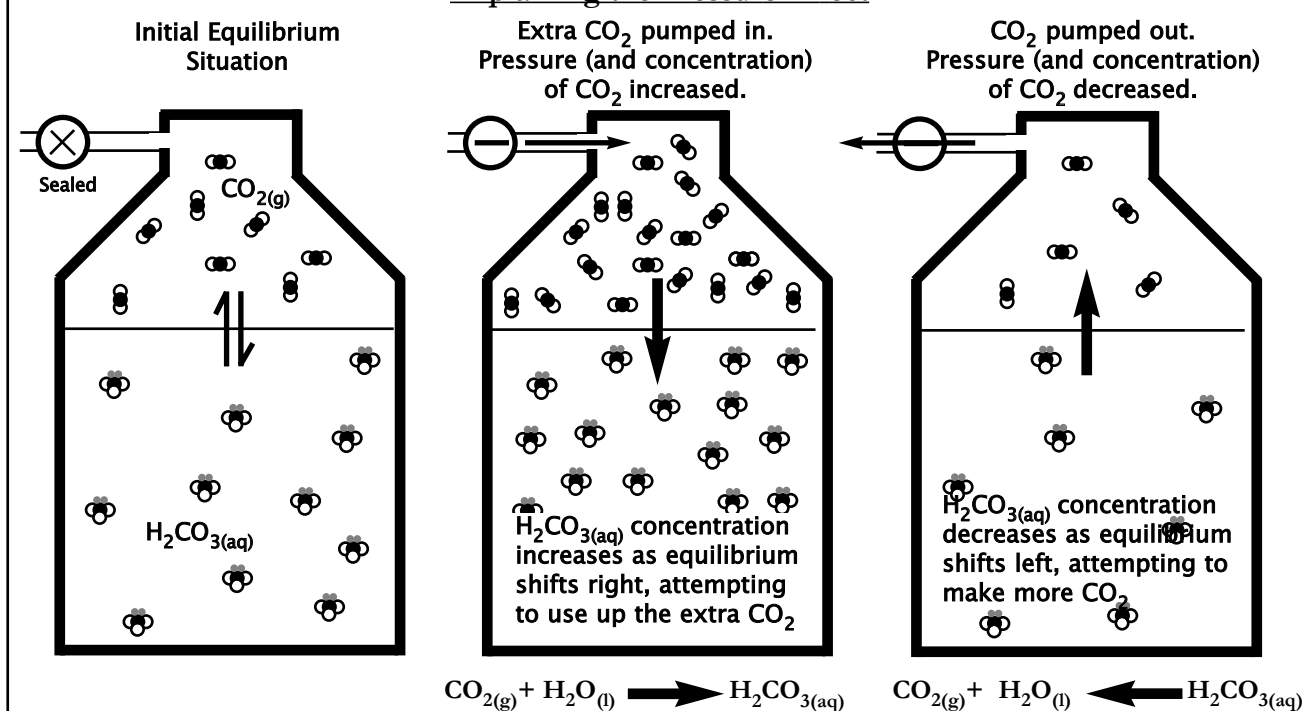
If you injected extra product so its concentration increased, the equilibrium shifts left, attempting to reduce the concentration of product



If you somehow removed product so its concentration was reduced, the equilibrium would shift right in an attempt to make more.

In every case, the equilibrium shift tries to counteract the change

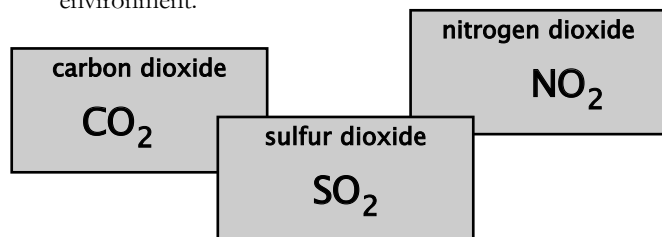
Explaining the Pressure Effect



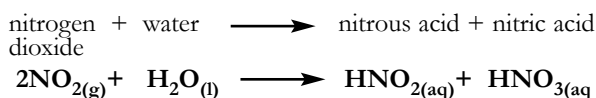
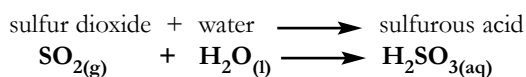
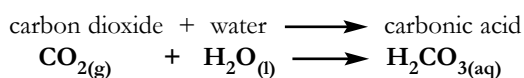
Acidic Oxides in the Environment

Learning about Equilibrium was a necessary diversion... now back to acids.

There are 3 main acidic oxides that are of concern in the environment:

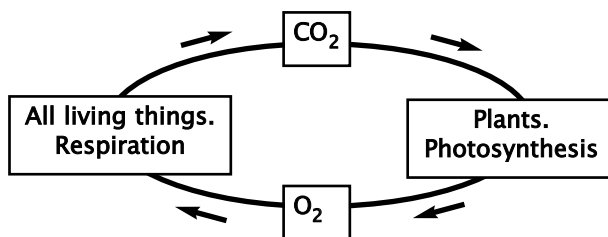


All 3 are gases produced by both natural processes and by human activities. All 3 react with water in the environment to form acids:



Carbon Dioxide

is a part of the great "Carbon-Oxygen" cycle in nature.



CO_2 is also released into the air by natural bush fires, and by volcanic eruptions, but generally there's a balance.

For millions of years, huge quantities of carbon have been "locked-away" in fossil materials such as coal and petroleum.

For the last 100 years or so, human activity has been releasing this "fossil carbon" as CO_2 by burning the fossil fuels. CO_2 levels have risen 30% or more, upsetting the balance, world-wide.

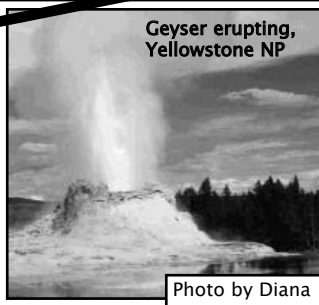
This is believed to be causing environmental problems ("Global warming") due to the "Greenhouse Effect".

Sources of Acidic Oxides in the Environment

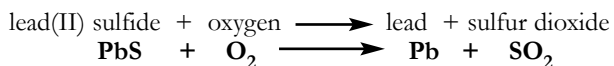
Sulfur Dioxide

is released into the air from volcanic eruptions and hot-springs, but the natural levels are extremely low.

Human activities can pollute the air of a region by releasing SO_2 from:

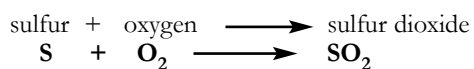


Smelting of Sulfide Ores (Revise Prel. topic "Metals")
The ores of some metals (esp. lead, zinc, copper) contain sulfide compounds. To extract the metal, the ore is roasted with air:



Burning of Fossil Fuels

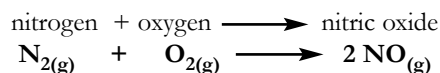
Some fuels, especially coal, contain small amounts of sulfur-containing compounds. When the coal is burnt, the sulfur burns too:



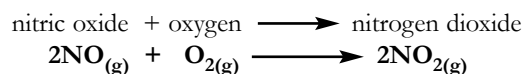
The main environmental concern with SO_2 release is "Acid Rain"

Nitrogen Dioxide

also occurs naturally in extremely small quantities due to the reaction between O_2 and N_2 when lightning provides the necessary energy:



Nitric oxide is NOT acidic, but it reacts with O_2 :



The combination of $\text{NO}_{(g)}$ and $\text{NO}_{2(g)}$ is referred to as the " NO_x " gases.

High temperature combustions (especially coal-burning power stations and vehicle engines) produce large amounts of NO_x gases.

NO_x pollution can cause toxic "smog" around large cities.



Environmental Impacts

Each of these acidic oxides can cause a different major environmental problem:

carbon dioxide



CO_2 is the weakest acid of these 3, and its acidity is not the problem. You should already be aware of the “Greenhouse Effect” and “Global Warming”

Acidity is not the main concern with the NO_x gases either. In the next topic you will study the Chemistry of “smog” and ozone pollution. (But NO_2 does contribute to “Acid Rain”)

nitrogen dioxide



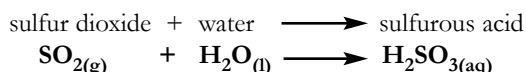
sulfur dioxide



and
“Acid Rain”

The main problem with SO_2 pollution is its acidity.

SO_2 reacts with water in the environment:



Sulfurous acid is a strong acid and where SO_2 pollution is serious, the acidity of rainfall stings the eyes, corrodes metals, erodes stone buildings and monuments, and can have serious environmental impacts:

- Lakes and wetlands can become acidic enough to kill plants and animals and disrupt the food-chains and the normal ecological balance.
- Forests can be killed by acidity of the rain, and the leaching of the soils by acids.

Deforestation

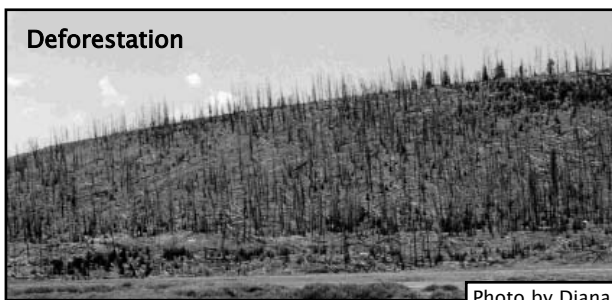


Photo by Diana

In the 1960's (when Acid Rain became known) many forests were seriously damaged, including the famous “Black Forest” of Germany. In Canada, 15,000 lakes were known to be “dead” in an ecological sense.

Since then, emissions of SO_2 have been limited and damage reduced, but the threat of Acid Rain is still a serious one, especially in rapidly industrializing countries such as China.

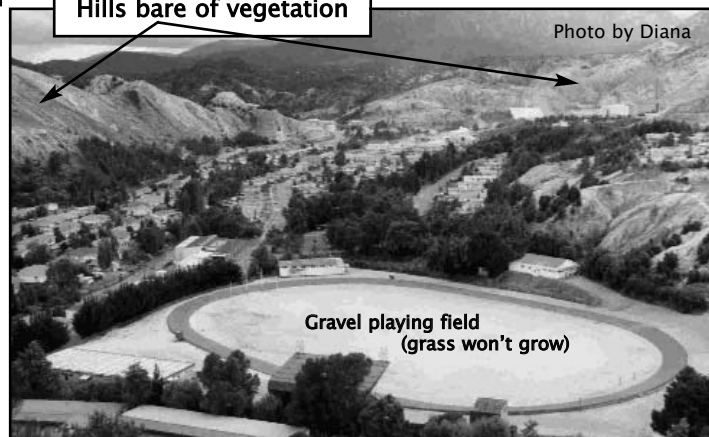
The Evidence for Acidic Oxide Pollution

We know that CO_2 levels in the atmosphere have increased because measurements have been collected over many years. The measurements of SO_2 and NO_x gases have not been collected for as long, and these gases are rapidly “washed” out of the air by rain, so the evidence for their presence is not so certain.

What we can be certain about are the localized effects of SO_2 pollution in places where it is, or was, prevalent. Some examples from the 1960's were mentioned before. Closer to home and to the present, is the evidence of devastation around Queenstown, Tasmania.

Hills bare of vegetation

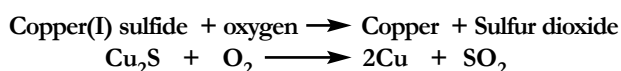
Photo by Diana



The environmental damage around Queenstown was caused by the release of SO_2 from the Mt. Lyell copper mine and smelter. Although it shut down over 20 years ago, the environment has still not recovered.

How Much Gas is Produced?

The Queenstown Copper Industry was smelting a copper ore containing mainly copper(I) sulfide. After concentrating the ore by froth flotation, the material was smelted by roasting in a furnace with a blast of air to provide oxygen:



How much $\text{SO}_{2(g)}$ is produced from each tonne of Cu_2S ?

Solution: 1 Tonne = 1,000 kg = 1.00×10^6 gram

$$\begin{aligned} \text{moles of Cu}_2\text{S: } n &= m/\text{MM} & (\text{MM} = 159.2\text{g}) \\ &= 1.00 \times 10^6 / 159.2 \\ &= 6.28 \times 10^3 \text{ mol.} \end{aligned}$$

mole ratio in equation is 1:1

$$\therefore \text{moles of SO}_2 \text{ produced} = 6.28 \times 10^3 \text{ mol.}$$

$$\begin{aligned} \text{mass(SO}_2\text{): } m &= n \times \text{MM} & (\text{MM} = 64.07\text{g}) \\ &= 6.28 \times 10^3 \times 64.07 \\ &= 4.02 \times 10^5 \text{ g (402 kg)} \end{aligned}$$

$$\begin{aligned} \text{volume(SO}_2\text{): } V &= n \times V_{\text{M}_3} & (\text{at SLC}) \\ &= 6.28 \times 10^3 \times 24.8 \\ &= 1.56 \times 10^5 \text{ L (156,000 litres!!)} \end{aligned}$$

Worksheet 2

Part A Fill in the blank spaces

Acids and bases react with each other to form a)..... and a "b)....." and c)..... each other.

The oxides of most metals act as d)....., in that they will react with an acid to form e)..... and a Oxides of non-metals mostly act as f)..... because they will either

- dissolve in/react with water to form an g)..... and/or
- react with bases and h)..... them.

Chemical i)..... occurs when a reaction, and its opposite are occurring at the same j)....., so that the concentrations of reactant(s) and k)..... do not change. The equilibrium is said to be l).....

An equilibrium can be "upset" by a change in m)....., or When a change occurs, it will "shift" to a new equilibrium position according to n)..... Principle. This states that when an equilibrium is disturbed, the system will shift in the direction which o).....

There are 3 acidic oxides which are of concern environmentally:

- Carbon dioxide is a natural part of the p).....-..... cycle in nature. Human activities have increased the CO₂ levels mainly from q)..... The major problem arising is "r)..... warming" due to the "s)....."
- Sulfur dioxide occurs naturally in t)..... eruptions and u)..... Human activities which release SO₂ include burning of v)..... and the w)..... of some metal ores. The main environmental problem is "x)....." which can cause serious ecological damage to lakes and forests.
- Nitrogen dioxide and nitric oxide (collectively known as y)..... gases) are produced naturally in very small amounts by z)..... Human activities which produce them are high temperature combustions in aa)..... and Environmentally, NO_x gases are the main cause of toxic "ab)....." in large cities, but also can contribute to Acid ac).....

**WHEN COMPLETED, WORKSHEETS
BECOME SECTION SUMMARIES**

Part B Practice Problems

1. Simple Acid-Alkali Reactions

- a) Name the salt formed in a reaction between:
- hydrochloric acid & calcium hydroxide
 - sulfuric acid & magnesium hydroxide
 - nitric acid & barium hydroxide

- b) Write a balanced symbol equation for the reaction of:
- hydrochloric acid and lithium hydroxide
 - sulfuric acid and sodium hydroxide
 - nitric acid and magnesium hydroxide

2. Reactions of Basic Oxides

Write a balanced symbol equation for the reaction of:

- sulfuric acid & iron(II) oxide
- hydrochloric acid & magnesium oxide
- nitric acid & copper(II) oxide

3. Reactions of Acidic Oxides

- carbon dioxide reacts with calcium hydroxide to form water and calcium carbonate. (This is the "limewater" reaction) Write a balanced equation for the reaction.
- P₂O₅ is an acidic oxide. It reacts with water to form phosphoric acid, H₃PO₄. Write the balanced equation.
- Sulfur trioxide reacts with water to form a strong acid. Write a balanced equation, and name the acid.

4. Le Chatelier's Principle

- a) NO_{2(g)} reacts with itself: $2\text{NO}_{2(g)} \rightleftharpoons \text{N}_2\text{O}_{4(g)}$

The reaction to the right is exothermic, so heat can be considered as a "product".



Note that as the reaction proceeds to the right, 2 moles of gas form 1 mole of gas, so in a fixed volume container the pressure would drop as the reaction proceeds to the right.

Imagine a sealed container in which a mixture of these gases has reached equilibrium.

State which way the equilibrium would shift if each of the following disturbances were made to the mixture. Explain each answer.

- Increase in temperature.
- Compress the mixture, thereby increasing pressure.
- Injecting extra N₂O₄, without changing pressure.
- Decrease the temperature.
- Spray in a little water. (NO₂ dissolves, N₂O₄ does not.)
- Decreasing the total gas pressure.

continued...

4. Le Chatelier's Principle (continued)

b) If hydrogen iodide (covalent molecular) is dissolved in water, some of the molecules ionize, and an equilibrium is reached:



What is the effect on this equilibrium of: (Explain each)

- adding $\text{NaI}_{(\text{aq})}$ solution, which increases the concentration of iodide ions.
- Adding NaOH , which reacts with H^+ ions, and reduces their concentration.
- Dissolving extra HI in the solution.

iv) It is found that raising the temperature of an equilibrium mixture has the effect of increasing the concentration of ions. Deduce whether the reaction as written is exo- or endothermic.

c) Ammonia is manufactured from its elements by the reaction

$$\text{N}_{2(\text{g})} + 3\text{H}_{2(\text{g})} \rightleftharpoons \text{NH}_{3(\text{g})} + \text{heat}$$

- To maximize the yield of ammonia, the reaction is carried out under very high pressure. Explain how this helps.
- The temperature of the reaction is kept fairly high to speed up the rate of the reaction. What effect does higher temperature have on the equilibrium?
- During the reaction, ammonia is constantly removed from the reaction vessel, and more reactant gases constantly pumped in. Explain the effect this has.

Volume of Gases under Different Conditions

You already know that the conditions known as "SLC" (Standard Laboratory Conditions) means

- 25°C
- and • 1 atmosphere of pressure (= 100 kPa)

At SLC, 1 mole of any gas occupies **24.8 Litres**

What about at other temperatures and pressures?

Another set of standard conditions commonly used in Chemistry is known as "Standard Temperature & Pressure" (STP):

- 0°C
- and • 1 atmosphere of pressure

At STP, 1 mole of any gas occupies **22.7 Litres**

Notes:

- You do NOT need to remember these values. They are always given in the Data Sheet in tests/exams.
- There is a formula for calculating the volume at any temperature and pressure, but its use is not required in this course.

5. Molar Gas Volumes

a) What is the volume of:

- 2.59 mol of O_2 at SLC?
- 0.0453 mol of H_2 at STP?
- 120 mol of CO_2 at STP?
- 4.67×10^{-2} mol of N_2 at SLC?

b) How many moles of gas is:

- 12.4L of He at SLC?
- 250mL (=0.250L) of O_2 at STP?
- 10,000L of N_2 at SLC?
- 1.00mL of Ar at STP?

6. Mass - Volume of Gases

a) What is the mass of:

- 5.00L of CO_2 at SLC?
- 5.00L of H_2 at SLC?
- 100L of Ne at STP?
- 25.0mL of O_2 at STP?

b) What is the volume (at SLC) of:

- 100g of CO_2 ?
- 100g of He?
- 1.50g of N_2 ?
- 1.00kg of Ar?

7. Problems Involving Reactions

a) Carbon dioxide gas reacts with aqueous calcium hydroxide (limewater) to form water and insoluble calcium carbonate.

- Write a balanced equation for the reaction.
- What mass of calcium carbonate would be formed by the reaction, if 1.00L of CO_2 , measured at SLC?
- What volume of CO_2 (at STP) must be absorbed by limewater in order to precipitate 1.75g of CaCO_3 ?

b) In the smelting of zinc, the crushed, concentrated ore is zinc sulfide. This is roasted in a blast of air, forming zinc oxide and sulfur dioxide gas.

- Write a balanced equation for the reaction.
- Calculate the volume of pure oxygen gas (at SLC) required for complete reaction with 1.00 Tonne of zinc sulfide.
- Assuming the air used is 21% oxygen, what volume of air must be supplied to the smelter for each tonne of ore?
- From your answer to (ii), use Avogadro's Hypothesis to find the volume of SO_2 (at SLC) released from one tonne of ZnS .

**FOR MAXIMUM MARKS SHOW
FORMULAS & WORKING,
APPROPRIATE PRECISION & UNITS
IN ALL CHEMICAL PROBLEMS**

3. ACIDS & THE pH SCALE

Acids as Proton Donors

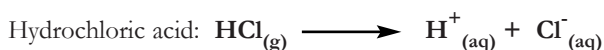
So far, you have learnt a number of things about acids, but not the most basic thing... the chemical definition of just what an acid is!

What defines all acids is their ability to donate a hydrogen ion (H^+) to another species. Since a hydrogen ion is really just a “naked” proton from the nucleus of the hydrogen atom, the formal definition of an acid is:

An acid is a chemical species which DONATES PROTONS

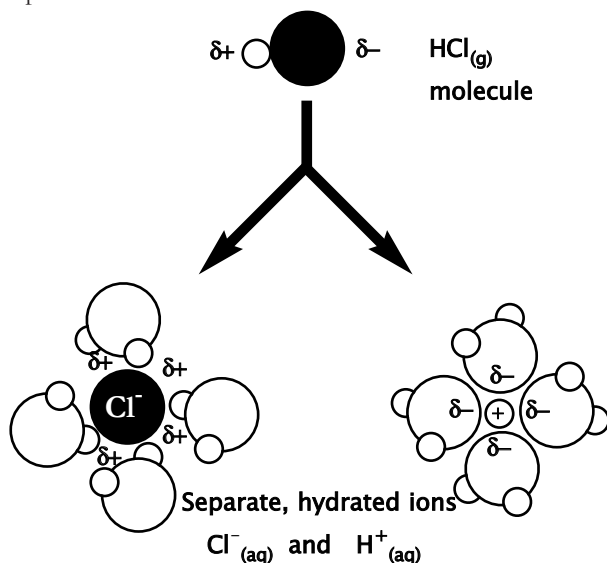
Acids in Aqueous Solution

When acids dissolve in water, we often just imagine an ionization such as this example:

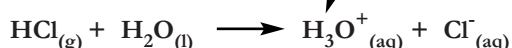


However, this is NOT the full story!

In a previous topic (Preliminary topic “Water”) you learnt about how ions are “hydrated” by polar water molecules when in solution. The dissolving of HCl in water could be represented as:



Although there may be many water molecules clinging to a H^+ ion, for simplicity we imagine there is just one, and it forms a special ion, called the “hydronium ion”, H_3O^+ .



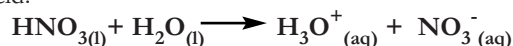
HCl is an acid because it has “donated” a proton to a water molecule. The hydronium ion in the solution is an acid because it can, in turn, donate a proton to other species.

In water solution, all acids produce hydronium ions.

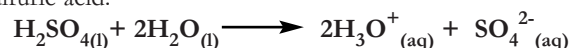
Formation of Hydronium Ions in Water

More examples:

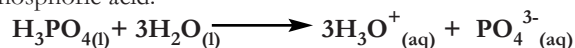
Nitric acid:



Sulfuric acid:



Phosphoric acid:



Some useful words:

MONOPROTIC = donating 1 proton (e.g. HCl)

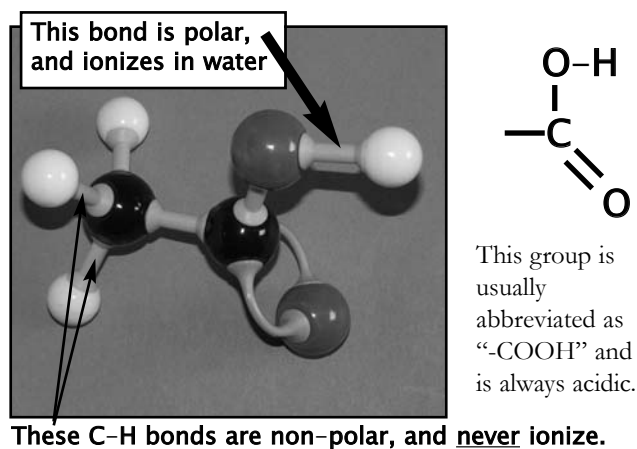
DIPROTIC = donating 2 protons (e.g. H_2SO_4)

TRIPROTIC = donating 3 protons (e.g. H_3PO_4)

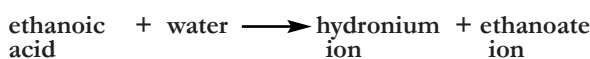
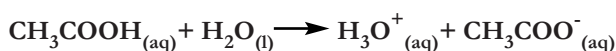
The “Organic” (Carbon-Based) Acids

Many naturally-occurring, biological molecules are acids too. These are carbon-compounds, made by living things and most contain a special chemical group you should become familiar with: the “-COOH” group.

Perhaps the most important is ethanoic acid, CH_3COOH :



When this molecule is dissolved in water the O-H bond ionizes, and donates a proton to a water molecule:



The syllabus requires you to know about ethanoic acid, and “citric acid” (next page)

Try the WORKSHEET at the end of section

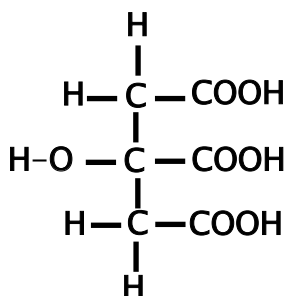
Citric Acid

is familiar as the acid in citrus fruits... oranges & lemons.

Its correct systematic name is

2-hydroxypropane-1,2,3-tricarboxylic acid

Molecular formula is $C_6H_8O_7$, but the structural formula is more meaningful:



Remember that only the $-\text{COOH}$ groups are acidic, and you will see that citric acid is **triprotic**

Strong & Weak Acids

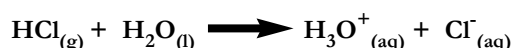
In everyday usage, a “strong” solution might mean the same as “concentrated” (i.e. having a lot of solute) and “weak” can mean the same as “dilute”.

FROM HERE ON, YOU MUST NOT USE THESE TERMS THAT WAY. “Strong” & “weak” have particular meanings with regard to acids.

STRONG Acid = Total Ionization in Solution

WEAK Acid = Partial Ionization in Solution

For example, HCl is a STRONG acid; when added to water the reaction...

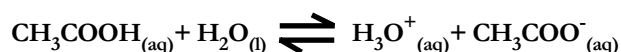


...goes fully to completion.

100% of the HCl molecules are ionized.

If the concentration of HCl was (say) 1 molL^{-1} , then the concentration of hydronium ions will also be 1 molL^{-1} .

In contrast, ethanoic acid is a WEAK acid; when added to water, the reaction...

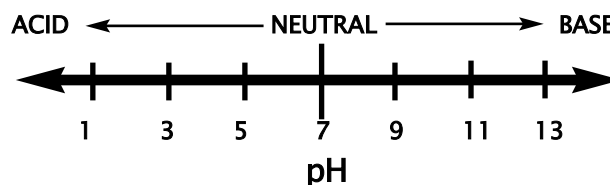


...reaches an equilibrium with only about 1% of the molecules ionized. (i.e. the equilibrium lies well to the left, favouring the reactant molecules.) If the concentration of the solution was 1 molL^{-1} , then the hydronium ion concentration would be only about 0.01 molL^{-1} .

It is quite possible to have a “concentrated solution of a weak acid” or a “dilute solution of a strong acid”... just be careful with these precise meanings!

The pH Scale

has already been introduced, and you should be familiar with it in a descriptive way.



Where do these numbers come from?

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$$

pH is the negative logarithm of the molar concentration of hydronium ions

Notes:

1. You must know that [square brackets] around any chemical species means “molar concentration of”.

2. “ \log_{10} ” means the “logarithm to base 10”. This is a mathematical function, best understood by example:

If $100 = 10^2$, then $\log_{10}(100) = 2$

$1,000 = 10^3$, then $\log_{10}(1,000) = 3$

$500 = 10^{2.699}$, then $\log_{10}(500) = 2.699$

pH values are therefore, powers of 10.

3. Different calculators may handle log functions differently. You must find out or figure out how to do log functions on your calculator.

Example Calculations: **pH and $[\text{H}_3\text{O}^+]$**

1. If the concentration of hydronium ions is $[\text{H}_3\text{O}^+] = 0.500 \text{ molL}^{-1}$, what is the pH?

Solution: **$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$**

$$= -\log_{10}[0.500]$$

$$= -(-0.301)$$

$$\therefore \text{pH} = 0.301$$

2. If the concentration of hydronium ions is $[\text{H}_3\text{O}^+] = 0.00252 \text{ molL}^{-1}$, what is the pH?

Solution: **$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$**

$$= -\log_{10}[0.00252]$$

$$= -(-0.260)$$

$$\therefore \text{pH} = 2.60$$

3. If the pH = 3.75, what is $[\text{H}_3\text{O}^+]$?

Solution: **$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$**

$$\text{so } [\text{H}_3\text{O}^+] = \text{Inverse}(\log(-3.75))$$

$$\therefore [\text{H}_3\text{O}^+] = 0.000178 \text{ molL}^{-1}$$

Try the WORKSHEET at the end of section

More about pH

Since the numbers on the pH scale are powers of 10, it follows that if an acid solution has a pH one unit lower than another, it is actually 10 times more acidic. Two units on the pH scale represents 100 times (10^2) difference in $[\text{H}_3\text{O}^+]$.

Example:

Acid "P": $[\text{H}_3\text{O}^+] = 10^{-5} \text{ molL}^{-1}$ pH = 5

Acid "Q": $[\text{H}_3\text{O}^+] = 10^{-3} \text{ molL}^{-1}$ pH = 3

Acid "R": $[\text{H}_3\text{O}^+] = 10^{-2} \text{ molL}^{-1}$ pH = 2

In the examples above, acid "Q" has a $[\text{H}_3\text{O}^+]$ value 100 times larger than "P", and its pH is 2 units lower.

Acid "R" has a $[\text{H}_3\text{O}^+]$ value 10 times higher than "Q", and its pH is 1 unit lower.

The pH scale is said to be "logarithmic", because the values are logarithms... powers of 10.

Measuring pH

You may have carried out some simple laboratory experiments, using a pH meter (or data-logger probe) to measure the pH of various solutions.



You may have recorded the pH value, then used your knowledge of the pH scale to decide if each solution tested was

Acidic (pH < 7)

or

Neutral (pH = 7)

or

Basic (pH > 7)

pH of Strong & Weak Acids

Remember the special meanings of "strong" and "weak" with regard to acids.

You may have used a pH meter to measure the pH of several different acids of exactly the same concentration.

1.00 molL⁻¹ Hydrochloric Acid

When $\text{HCl}_{(g)}$ dissolves in water:



Since HCl is a STRONG ACID, the ionization is 100%.

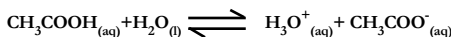
If $[\text{HCl}] = 1.00 \text{ molL}^{-1}$,
then $[\text{H}_3\text{O}^+] = 1.00 \text{ molL}^{-1}$

and $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$
 $= -\log_{10}[1.00]$
 $\therefore \text{pH} = 0$

Sure enough, by experiment you will have found that a 1 molL⁻¹ HCl solution has a pH = 0.

1.00 molL⁻¹ Ethanoic Acid

When $\text{CH}_3\text{COOH}_{(g)}$ dissolves in water:



Since this is a WEAK ACID, the ionization is incomplete.

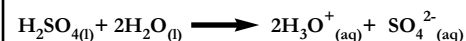
If $[\text{CH}_3\text{COOH}] = 1.00 \text{ molL}^{-1}$,
then $[\text{H}_3\text{O}^+] < 1.00 \text{ molL}^{-1}$

and the pH can be expected to be well above the values found for HCl or H_2SO_4 .

Sure enough, by experiment you will have found that a 1 molL⁻¹ CH_3COOH solution has a pH ≈ 2 .

1.00 molL⁻¹ Sulfuric Acid

When $\text{H}_2\text{SO}_{4(l)}$ dissolves in water:



Since H_2SO_4 is a STRONG ACID, the ionization is 100%. It is DIPROTIC, so each mole of H_2SO_4 produces 2 moles of H_3O^+ .

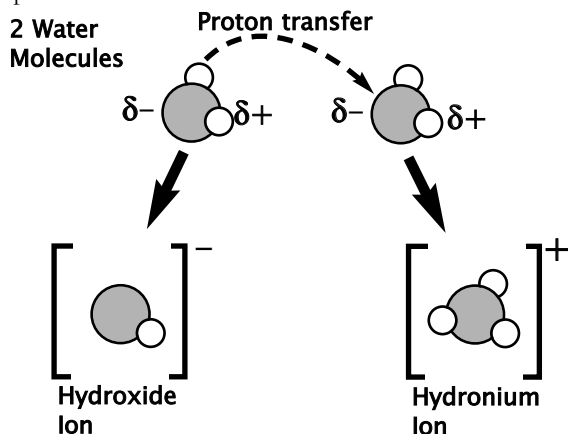
If $[\text{H}_2\text{SO}_4] = 1.00 \text{ molL}^{-1}$,
then $[\text{H}_3\text{O}^+] = 2.00 \text{ molL}^{-1}$

and $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$
 $= -\log_{10}[2.00]$
 $= -(0.301)$
 $\therefore \text{pH} = -0.30$

Sure enough, by experiment you will have found that a 1 molL⁻¹ H_2SO_4 solution has a pH = -0.30. (below zero)

The Self-Ionization of Water

Water itself can be considered as an extremely weak acid, since in pure water a few water molecules ionize and donate a proton to another molecule.



In pure water, at 25°C, it turns out that the concentration of hydronium ions, $[\text{H}_3\text{O}^+] = 1.00 \times 10^{-7} \text{ molL}^{-1}$.

So the pH of pure water:

$$\begin{aligned} \text{pH} &= -\log_{10}[\text{H}_3\text{O}^+] \\ &= -\log_{10}[1.00 \times 10^{-7}] \\ &= -(-7.00) \end{aligned}$$

$$\therefore \text{pH} = 7.00$$

- That's why $\text{pH} = 7$ is **neutral** on the pH scale!
- In an acid solution, $[\text{H}_3\text{O}^+] > 1.00 \times 10^{-7} \text{ molL}^{-1}$, so $\text{pH} < 7$
- In a basic solution, $[\text{H}_3\text{O}^+] < 1.00 \times 10^{-7} \text{ molL}^{-1}$, so $\text{pH} > 7$

Acids as Food Additives

Acids are common food additives. If you read the "Ingredients" listing on many processed foods, from tomato sauce to ice-cream topping to dried fruits or tinned sausages, you may find "food acid (260)" (or something similar).

"Food acid (260)" is in fact ethanoic acid, the main chemical in vinegar, and a common ingredient in many recipes.

The reasons that acids are added to many foods are:

- **Preservatives.** Adding acids to a food lowers the pH and makes it more difficult for some bacteria or fungi to grow within the food. This is the main reason why some foods, such as tomato sauce, do not go bad even when not refrigerated. Acids most commonly used are SO_2 & ethanoic.
- **Flavour.** Many foods, such as jams, sauces and fruit-flavoured drinks will taste better if they have that "sharp" or sour effect that acids give. Acids commonly used are citric and ethanoic acids.
- **Nutrition.** Some foods have vitamins added to increase their nutritional value. Most commonly used is ascorbic acid (vitamin C).



Photo by Norbert Machmek

Some Naturally Occurring Acids & Bases

Acids

Ever been bitten by a bull-ant or stung by a bee?

You were injected with "formic acid" (methanoic acid), HCOOH .

Ethanoic acid is the main chemical in vinegar (literally "wine-sour") known and used for thousands of years.

SUGAR (grapes)
 ↓ fermentation
ETHANOL (wine)
 ↓ oxidation
ETHANOIC ACID (vinegar)

Bases

Lime is a major ingredient of cement, and has been known and used for thousands of years. Chemically it is calcium oxide, CaO .

It is prepared by roasting limestone (or even oyster shells) which results in the decomposition



"Lye" is the name given to the alkali liquid obtained by soaking wood ashes in water. The basic ingredient is a mixture of hydroxide compounds, mainly potassium and sodium hydroxides.

Lye has been used for centuries in soap-making.

Part A Fill in the Blank Spaces

4. ACID-BASE THEORY & TITRATION

Early Ideas About Acids & Bases

Naturally occurring acids, like ethanoic acid in vinegar, have been known for thousands of years, and described by their simple, observable properties such as their sour taste.

Chemistry became a modern Science just over 200 years ago, and one of the earliest scientific theories about acids was made by

Antoine Lavoisier (French, 1780's)

Lavoisier used simple plant-extract indicators (e.g. litmus) to identify acids and bases. He found by experiment that all the non-metal oxide compounds he tested produced acid solutions. He concluded that acids must contain oxygen.

Humphry Davy (English, 1820's)

Forty years later and armed with more chemical knowledge, Davy realized that Lavoisier was wrong about acids. While some acids do NOT contain oxygen (e.g. HCl), he found that all known acids contain hydrogen.

Davy also discovered that metal oxides are basic and described the patterns in the way acids react with metals to form hydrogen gas.

These early ideas were all empirical descriptions of properties; they described the properties of acids discovered by experiment, but did not include a general theory to explain or predict chemical behaviour.

That had to wait for

Svante Arrhenius (Swedish, 1880's)

The Arrhenius Theory of acids-base behaviour is still generally used in years 7-10 Science.

**Acids produce H^+ ions in water solution.
Bases produce OH^- ions (or oxide ions).**

The Arrhenius Theory was very successful in accounting for simple acid-base behaviour in water solution. It explained the neutralization reaction, and could explain strong and weak acids as being due to complete or partial ionization.

However, the Arrhenius Theory had some deficiencies:

- It could not account for acid-base behaviour that was not in water solution.
- It could not explain why many ions (which did NOT contain any hydrogen, hydroxide or oxide) showed acid or base behaviour in water solution.

You will soon learn that the main reason for all these deficiencies was that Arrhenius failed to consider the role of the solvent itself, and in water solutions this is critical.

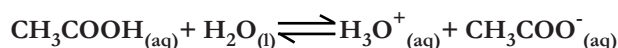
The Bronsted-Lowry Theory

was developed independently by Johannes Bronsted (Danish) and Thomas Lowry (English) in 1923. The basic concepts of this theory have already been used earlier in this topic:

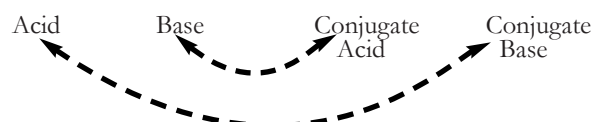
All Acid-Base Reactions involve transfer of Protons

**An Acid is a Proton Donor
A Base is a Proton Acceptor**

Central to the Bronsted-Lowry (B-L) Theory is the concept of “conjugate” species. For example, when Ethanoic acid dissolves in water:



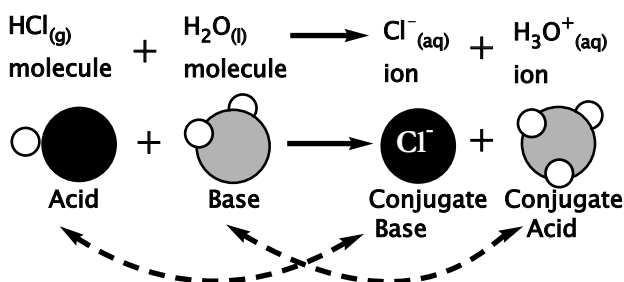
ethanoic acid + water \rightleftharpoons hydronium ion + ethanoate ion



The CH_3COO^- ion is the conjugate base of Ethanoic acid, because if this reaction was to run in reverse, the ion could act as a base and accept a proton to form ethanoic acid again.

The water molecule is acting as a base, because it accepts a proton to become the H_3O^+ ion. The H_3O^+ ion is the conjugate acid because, if the equilibrium shifts left, it can donate a proton and become a water molecule again.

Another Example: Dissolving of HCl in water:



In this case, the reaction is very unlikely to ever run in reverse, but Cl^- is still considered the conjugate base.

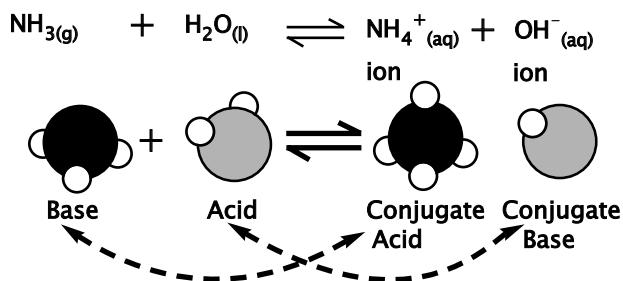
As before, water has acted as a base, and its conjugate acid is the hydronium ion.

Try the WORKSHEET at the end of section

Water Can be an Acid, Too

In both the previous examples, the water molecule acted as a base by accepting a proton to form a H_3O^+ ion. Water can also act as an acid and donate a proton.

For example, when ammonia, NH_3 dissolves:



In this reaction the water molecule acts as an acid by donating a proton. Its conjugate base is the hydroxide ion, OH^- .

This equation explains why ammonia is a base, and why a solution of ammonia ($\text{NH}_3(\text{aq})$) can be considered as a solution of ammonium hydroxide, $\text{NH}_4\text{OH}(\text{aq})$.

Chemical Species (like water) which can both donate and accept protons are called "AMPHIPROTIC"

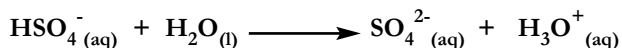
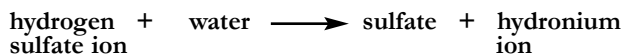
You will meet more Amphiprotic species soon

Acidic and Basic Salts

One of the weaknesses of the Arrhenius Theory was that it could not explain the results of simple experiments you may have done:

Acidic Salts

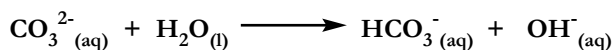
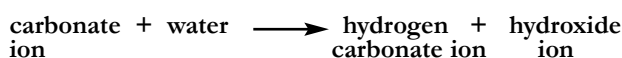
As some ions dissolve in water they react as acids, forcing the water molecule to be a base. An example is the "hydrogen sulfate ion", HSO_4^- .



If the salt used was sodium hydrogen sulfate there would also be some sodium ions in the solution. In terms of acid-base behaviour, they would be merely spectators.

Basic Salts

dissolve in water and react as bases, forcing the water molecule to be the acid. An example is the carbonate ion:



Once again, if sodium carbonate was used, there would be sodium ion "spectators" in the solution as well.

A few worth knowing...

Acidic Salts	Neutral Salts	Basic Salts
NaHSO_4	NaCl	Na_2CO_3
NH_4Cl	KNO_3	CH_3COONa
NH_4NO_3	Na_2SO_4	(sodium ethanoate)

(There is a pattern to help you remember these... later)

Try the WORKSHEET at the end of section

Practical Work: Testing the pH of Various Salt Solutions

You may have done simple experiments using a pH meter, or Universal Indicator, to test the pH of solutions of various ionic "salts", none of which seem to be obviously acids or bases, from their formulas.

You may have tested many, but they probably included...

sodium chloride, NaCl

sodium carbonate, Na_2CO_3

sodium hydrogen sulfate, NaHSO_4

...all at the same concentration, in pure water.



Typical Results

Salt	pH	Conclusion
NaCl	7	Neutral
Na_2CO_3	11	Basic
NaHSO_4	2	Acidic

The reasons WHY some salts show acid-base behaviour is explained above right.

Note: Due to CO_2 from the air dissolving in the solution, the pH of the NaCl solution may be slightly acidic.

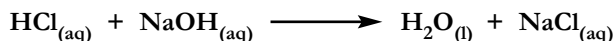
Neutralization

Now that you know about B-L Theory, we can go back to look again at the simple acid-alkali neutralization.



Example

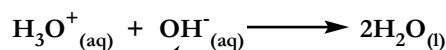
hydrochloric acid + sodium hydroxide \longrightarrow water + sodium chloride



However, now we know that what makes HCl an acid is really the formation of H_3O^+ ions in water, and the Cl^- ions are spectators, as are the Na^+ ions from the NaOH.

Leaving out the spectators, the net ionic equation is...

hydronium ion + hydroxide ion \longrightarrow water

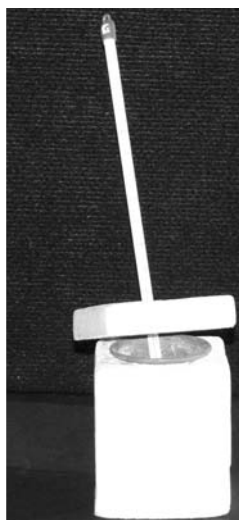
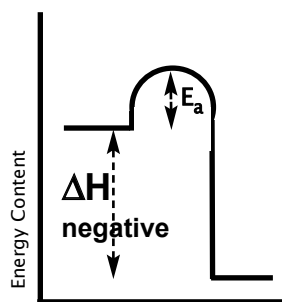


The reaction involves a proton transfer

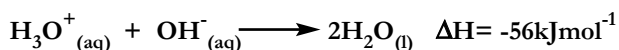
... and this is the net reaction for ALL the simple acid-alkali reactions, regardless of which acid or which alkali are used.

Heat of Neutralization

If you carried out the reaction above in a calorimeter, you will quickly find that the temperature rises... the reaction is **exothermic**.



If measured and calculated, it is found that the value for ΔH is the same regardless of which acid and alkali is used.



The information in the Right-Hand column is **not** specified by the Syllabus. It is presented here in the interests of better understanding.

Salts

from Different Acid-Base Neutralizations

Although we might view all neutralizations as essentially the same reaction and ignore the "spectator ions" which form the "salt", this prevents us noticing a useful pattern which was mentioned on the previous page.

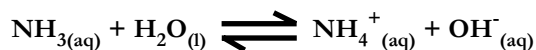
The nature of the "salt" formed depends on whether the acid and base were "strong" or "weak", as follows:

	Strong Base e.g. KOH, NaOH	Weak Base e.g. NH_3
Strong Acid e.g. HCl, HNO_3 , H_2SO_4	forms NEUTRAL salts e.g. KCl, NaNO_3 , K_2SO_4	forms ACIDIC salts e.g. NH_4Cl , NH_4NO_3 , $(\text{NH}_4)_2\text{SO}_4$
Weak Acid e.g. CH_3COOH , H_2CO_3	forms BASIC salts e.g. CH_3COONa , Na_2CO_3	forms NEUTRAL salts e.g. $\text{CH}_3\text{COONH}_4$, $(\text{NH}_4)_2\text{CO}_3$

Why is Ammonia (NH_3) a "Weak Base"?

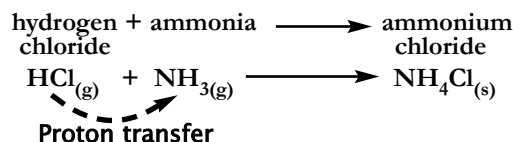
Simple!

Because when it reacts with water, the reaction does NOT ionize fully, but reaches an equilibrium with a significant concentration of un-ionized molecules present.



Acid-Base Reactions without H_3O^+ , OH^- or H_2O

Although most of the simple examples of neutralization involve the reaction at left, and usually we study reactions taking place in water, be aware that this is not always the case. For example, if you add together the dry gases $\text{HCl}_{(\text{g})}$ and $\text{NH}_{3(\text{g})}$ they will react



Although there is no water present, and no H_3O^+ or OH^- ions are formed at any time, this is clearly an acid-base reaction, according to the B-L definition.

Titration

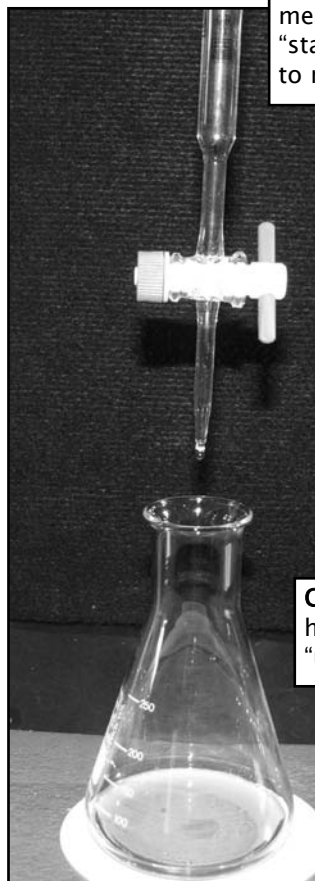
Titration has been one of the most important techniques in Chemical Analysis for over a century. Its use has diminished as electronic “probes” have become more widespread, but it remains a “must-know” part of Chemistry.

Titration can be used for any reaction for which you can determine the “end-point”. Acid-base neutralizations are ideal because indicators (or pH probes) can identify the end-point.

The purpose of titration is Chemical Analysis. It allows the concentration of an “unknown” solution to be determined by calculation, after measuring the volume of a known-concentration “standard solution” which reacts with a sample to reach the “equivalence-point” (end-point). This is the point where the reactants have been consumed in exactly the molar ratio specified by the balanced equation.

Titration Technique

- you must have a “standard solution” (see next page)
- the burette should be rinsed with small quantities of the standard solution, then filled with it. The level is then adjusted to the zero mark, ensuring that there are no air bubbles (especially in the tap and delivery spout).
- an exact, measured volume of “unknown” solution is placed in the reaction flask by pipette.
- 1-2 drops of suitable indicator are added to the flask. (“Choosing the Indicator”, next page)
- the reaction proceeds by careful addition, with mixing, of solution from the burette. Add drop-by-drop approaching the end-point, until the indicator just changes colour.
- record the volume of solution delivered by burette.
- repeat 3 times, and average the “titres” (= burette volumes). It is common to discard any values which are not in close agreement. **The rest is calculation.**



Burette
measures the volume of “standard solution” needed to reach the end-point.

Burette tap
controls the addition of “standard solution” to the “unknown solution”

Conical flask
holds a measured volume of “unknown solution”, plus indicator

Filter Paper
under flask allows better seeing of indicator colour

Example Titration Calculation

Titration of an “unknown” solution of KOH.

“Aliquot” (volume of samples) = **25.00 mL** (by pipette)

Standard solution; H₂SO₄ solution, **C = 0.04252 molL⁻¹**.

Indicator used: **Bromothymol Blue** (see next page)

“Titres” (volumes from burette)

measured (in mL): ~~34.25~~, 33.90, 33.85, 33.95.

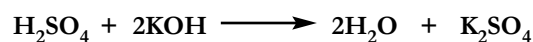
The first titre is discarded (not in close agreement with the others). **Average titre = 33.90 mL**

Titration Formula: $\frac{C_a \times V_a}{a} = \frac{C_b \times V_b}{b}$

C_a, C_b = concentrations (molL⁻¹) of acid and base solutions
V_a, V_b = volumes (mL) of each solution used.

a, b = molar co-efficients (balancing numbers) from balanced equation for reaction.

Step 1: Balanced equation (ALWAYS!!)



Step 2: Titration Formula

$$\frac{C_a \times V_a}{a} = \frac{C_b \times V_b}{b}$$

Step 3: Re-arrange and substitute.

We are trying to find the concentration of the “unknown” base, so make “C_b” the subject:

$$\begin{aligned} C_b &= \frac{b \times C_a \times V_a}{a \times V_b} \\ &= \frac{2 \times 0.04252 \times 33.90}{1 \times 25.00} \\ &= 0.1153 \end{aligned}$$

∴ Concentration of KOH solution = 0.1153 molL⁻¹.

Preparing a Standard Solution

The key to titration is to have available a suitable “standard solution”. The technique for making a solution to an exact known concentration was covered in the Preliminary topic “Water” and is revised below.

The problem is to get a substance which is of very high purity and stability to use as a “Primary Standard” to make the solution from.

The common acids like H_2SO_4 , and bases like NaOH , cannot be obtained in the pure state due to the way they rapidly absorb water and/or CO_2 from the atmosphere.

Suitable primary standard substances are:

- Base: anhydrous sodium carbonate, Na_2CO_3
- Acid: oxalic acid, COOHCOOH (diprotic)

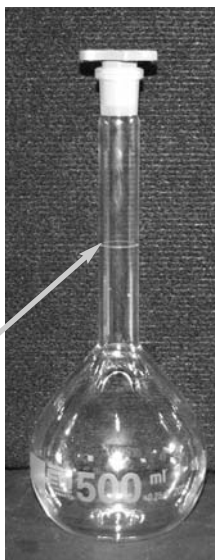
Calculate the mass of pure, dry Primary Standard solute required for solution, and weigh out accurately

Dissolve Solute in a small amount of (pure) water in a clean beaker

Carefully transfer solution into a Volumetric Flask. Rinse beaker with small amounts of water & add washings to flask

Add water to flask to fill it to the mark. (Use a dropper to avoid overshooting) Insert stopper & mix well.

Volumetric Flask



In the example (previous page) the standard solution was H_2SO_4 . This would not have been a primary standard solution.

It may have been “standardized” by titration against a primary standard Na_2CO_3 solution, so that its exact concentration is known. Then it can be used, in turn, as the titration standard solution.

Selecting the Indicator

Choosing the best indicator to use is not just a matter of which colour change you’d prefer looking at.

Each indicator changes colour over a range of pH values, and the trick is to choose an indicator which will change colour as close to the “end-point” as possible.

Indicator	Colour in		Changes at pH
	Acid	Base	
Litmus	red	blue	6 - 8
Bromothymol	blue	yellow	6.2 - 7.6
Methyl orange	red	yellow	3.1 - 4.4
Phenolphthalein	clear	pink	8.3 - 10.0

But, isn’t the end-point always at neutral, pH = 7 ??
(I hear you ask)

No, its not! Remember (see p.19) that the “salts” formed by different combinations of strong or weak acids and bases may have acidic or basic properties. This means that, at the exact end point, the pH might not be neutrality.

Examples:

Strong Acid - Strong Base

If titrating (say) H_2SO_4 against KOH , you can expect the salt K_2SO_4 to be neutral.

Therefore, choose **Bromothymol blue**, which changes colour near pH = 7.

(Actually, it doesn’t really matter, because the the last drop of chemical at the end point causes a huge pH change.)
See graph next page.

Strong Acid - Weak Base

If titrating (say) H_2SO_4 against NH_3 , you can expect the salt $(\text{NH}_4)_2\text{SO}_4$ to be acidic.

Therefore, choose **Methyl orange**, which changes colour around pH = 3-4.

Weak Acid - Strong Base

If titrating (say) CH_3COOH against KOH , you can expect the salt (potassium ethanoate, CH_3COOK) to be basic.

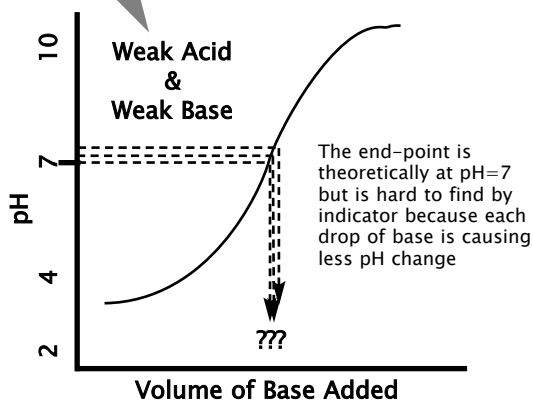
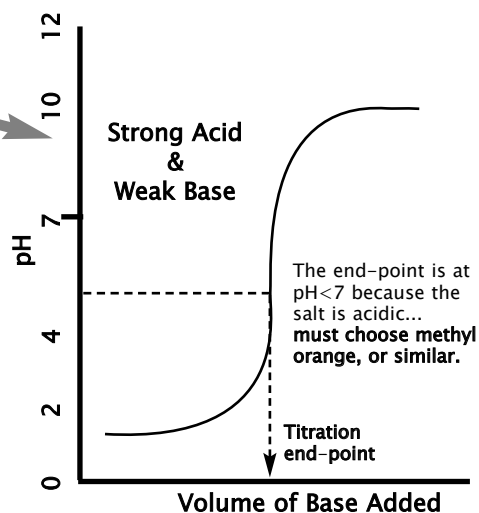
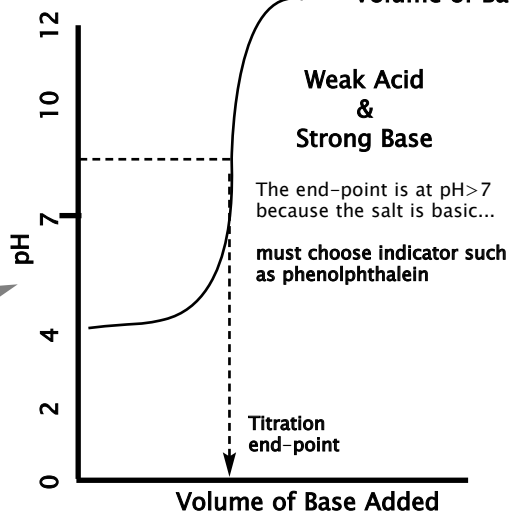
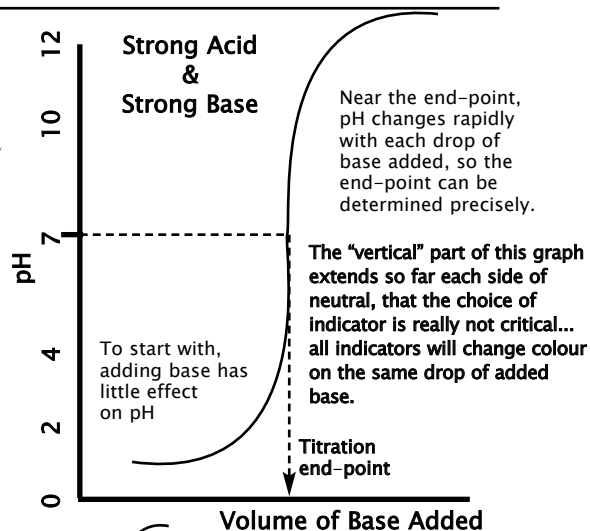
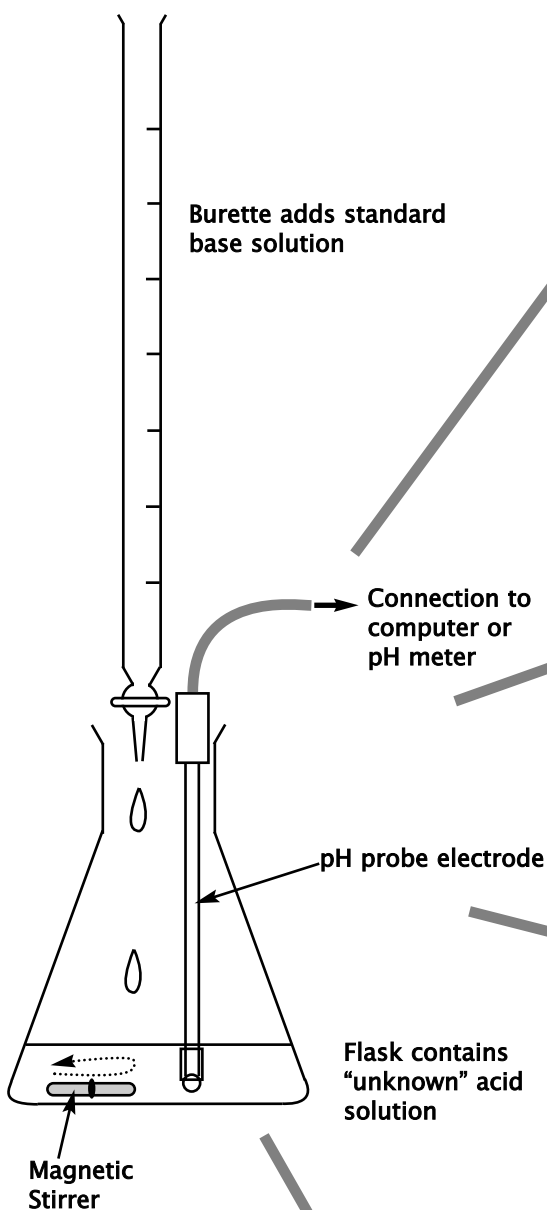
Therefore, choose **Phenolphthalein**, which changes colour around pH = 8-10.

Weak Acid - Weak Base

Titration with this combination are to be avoided, because the end-point is not very “sharp”.

Measuring pH During Titration

If a titration is carried out as shown below, with a pH meter or data-logger probe attached to a computer, the pH changes occurring in the solution can be recorded and graphed.



In every case, the end-point is the centre of the vertical (or near-vertical) part of the graph.

The graph at left shows why weak acid-weak base titrations are best avoided.

Buffers

A “Buffer”, or “buffered solution”, is a solution which can absorb significant amounts of acid or base with minimal change in pH.

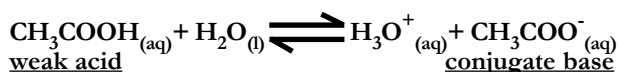
Biology students will be aware that in all living things it is vital that the conditions within the body/cells are kept very constant. Living things cannot function properly if their “internal environment” undergoes large changes in

- temperature,
- water content,
- salt concentration, (and many other chemicals) AND
- **pH level.**

The fluids inside all living things are buffered, so that pH remains remarkably constant, despite changes in the external environment, eating acidic food, breathing, excreting, etc, all of which could alter the body's pH.

How do Buffers Work?

All buffers are solutions containing a weak acid and its conjugate base.



This equilibrium constitutes a buffer solution, so long as there are significant amounts of the ethanoate ion (CH_3COO^-) present. (e.g. by adding CH_3COONa)

If acid is added (i.e. concentration $[\text{H}_3\text{O}^+]$ increases) (pH lower) then (by Le Chatelier's Principle) the equilibrium shifts left, absorbing H_3O^+ and lowering $[\text{H}_3\text{O}^+]$ again.

If base is added (i.e. concentration $[\text{H}_3\text{O}^+]$ decreases) (pH higher) then (by Le Chatelier's Principle) the equilibrium shifts right, making H_3O^+ and raising $[\text{H}_3\text{O}^+]$ again.

This way, the pH remains quite constant, despite acid or base being added.

Using Neutralization on Chemical Spills

Whether its a small vinegar spill at home, or a sulfuric acid tanker leaking after a road accident, a knowledge of neutralization can help damage control and safety.



A Natural Buffer System

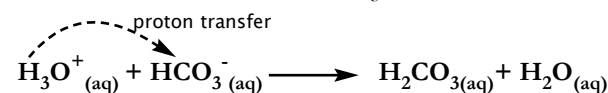
The classic example of a natural buffer system is the way our blood is maintained at a pH = 7.4, despite the fact that we keep exchanging CO_2 (acidic), excreting wastes, absorbing foods, etc.

The main chemical buffer in our blood is a solution containing both the bicarbonate ion (HCO_3^-) and its conjugate base, the carbonate ion (CO_3^{2-}).

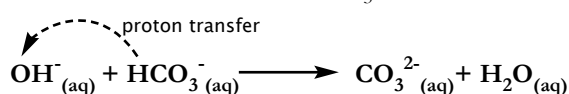
HCO_3^- is Amphiprotic

The bicarbonate ion is amphiprotic, which adds a further dimension to its buffering ability.

If the environment is acidic, HCO_3^- acts as a base:



If the environment is basic, HCO_3^- acts as an acid:



So, the total buffering system can be summarized as:



This system is highly effective at maintaining the pH of our blood. Be aware that it is not the only buffer operating.

As you'll learn below,
Amphiprotic substances
are useful in emergencies, too.

So, if someone gets splashed with a strong base, neutralize it by throwing a bucket of acid over them, right?

Wrong! Unless you can guarantee to apply the exact molar quantity for titration-like precision, then using an acid on an alkali spill (or vice-versa) can do more harm than good, and would be extremely dangerous as a first-aid method.

The best thing to do is use an **Amphiprotic substance** such as bicarbonate ion (e.g. sodium bicarbonate, NaHCO_3).

As shown by the reactions above, it can react and neutralize either an acid or a base, and once the neutralization is achieved, it stops working and poses no further threat in its own right.

It does its job, and then stops automatically... perfect!

Worksheet 4

Part A Fill in the blank spaces

Lavoisier concluded that acids must contain a).....
 Later, Davy described acids as all containing b).....,
 and being able to react with c).....

The first attempt at a complete acid-base theory was made by d)..... in 1885. According to this theory, all acids produce e)..... in solution. Bases were defined as compounds containing f)..... or ions. This theory was successful at explaining some acid-base behaviour, but failed to take into account the important role of g).....

The h)..... (B-L) Theory is much more useful for explaining things. It defines an acid as a i)....., and a base as a j)....., and all acid-base reactions involve the k)..... of one or more l).....

When an acid loses a proton it forms the “m)..... base”, and when a base gains a proton it forms the n).....

Water can act as either o)..... or depending on the chemical environment. The word for this is “p).....”

Many salts dissolve in water to form acidic or basic solutions, because of their interaction with q).....

An example of an acidic salt is r).....

An example of a basic salt is s).....

An example of a neutral salt is t).....

The “rules” governing this are:

- Neutral salts form from the neutralization of u)..... acids by v)..... bases.
- Acidic salts form from the neutralization of w)..... acids by x)..... bases.
- Basic salts form from the neutralization of y)..... acids by z)..... bases.

All neutralization reactions (in aqueous solution) involve the reaction of aa)..... and ions to form ab)..... The reaction is always ac).....-thermic.

ad)..... is a technique used for chemical analysis. A measured ae)..... of “unknown” solution is reacted with a “af)..... solution” until the “ag)..... point” is reached. The main piece of equipment involved is a ah)..... It is important to choose the ai)..... which will change aj)..... at the pH of the end-point, depending on the nature of the salt formed.

If the ak)..... is measured during titration and graphed, the graphs generally show an “S” shaped curve. The end-point is located in the middle of the al)..... part of the curve.

Buffers are solutions containing an equilibrium mixture of a am)..... and its If either acid or base is added, the equilibrium an)..... (according to ao)..... Principle) so that the change in ap)..... is minimized. Buffers are important in all aq)..... things, by helping to maintain ar)..... A specific example is as)....., where the pH is maintained by a mixture of at)..... and ions.

Part B Practice Problems

1. Reactions of Acids & Bases with Water

For each substance below, write an equation describing its monoprotic reaction with water. (In brackets is a description of how the substance behaves in each case)

For each equation, identify the conjugate acid/base of the named species.

- methanoic acid, HCOOH (acid)
- ammonia, NH_3 (base)
- hydrogen phosphate ion, HPO_4^{2-} , (acid)
- hydrogen phosphate ion, HPO_4^{2-} , (base) **Amphiprotic**
- sulfide ion, S^{2-} (base)
- cyanide ion, CN^- (base)
- hydrogen sulfide, H_2S (acid)
- nitrite ion, NO_2^- (base)
- ammonium ion, NH_4^+ (acid)
- hydrogen sulfite ion, HSO_3^- (acid)

2. Amphiprotic Substances

Each of the substances below is amphiprotic.

For each, write TWO equations to show its reaction

- with H_3O^+
- with OH^-

- dihydrogen phosphate ion, H_2PO_4^-
- hydrogen carbonate ion, HCO_3^-
- hydrogen sulfide ion HS^-

(continued...)

**WHEN COMPLETED, WORKSHEETS
BECOME SECTION SUMMARIES**

Worksheet 4 Part B (continued)

3. Acidic and Basic Salts

Each of the following salts dissolves in water, and reacts (as shown in brackets) to form an acidic or basic solution.

Write a net ionic equation (leave out spectators) to show the reaction with water.

- potassium ethanoate, CH_3COOK (basic)
- ammonium nitrate, NH_4NO_3 (acidic)
- sodium nitrite, NaNO_2 (basic)
- potassium hydrogen oxalate, KHC_2O_4 (acidic)
- lithium cyanide, LiCN (basic)

4. Titrations

a) 25.00mL of an “unknown” NaOH solution was titrated against standardized HCl, concentration = $0.09255 \text{ molL}^{-1}$. The titration was carried out 4 times, with the end-point titres being 22.50mL, 22.45mL, 23.10mL and 22.50mL.

Average these results appropriately, then write an equation, and calculate the concentration of the NaOH solution.

b) Using a $0.05025 \text{ molL}^{-1}$ standardized solution of NH_4OH (ammonia solution), 25.00mL samples of an unknown H_2SO_4 solution were titrated. The average titre was 28.32mL

- Write a balanced equation for the neutralization.
- Find the concentration of the acid.

iii) For this titration there were 3 indicators available:

Indicator	pH of Colour Change
J	$\cong 8.7$
K	$\cong 6.8$
L	$\cong 4.2$

Which indicator is most appropriate for this titration? Explain your answer.

c) A student wishes to prepare 500mL of a $0.02500 \text{ molL}^{-1}$ solution of oxalic acid (COOHCOOH), from the solid chemical.

- What are the characteristics that qualify this chemical as a “primary standard”?
- What mass needs to be accurately weighed out?
- Summarize the main steps involved in preparing the solution.

iv) The solution, when prepared, was used to determine the concentration of a KOH solution. 25.00mL aliquots of KOH required an average titre of 31.45mL. Find the concentration of the KOH. (care: oxalic acid is diprotic)

d) To find the concentration of an ammonia solution (NH_4OH), a student has 2 choices of standardized solutions she could use:

- $0.7438 \text{ molL}^{-1} \text{ HNO}_3$ solution
- $0.8863 \text{ molL}^{-1} \text{ CH}_3\text{COOH}$ solution

- Which solution should she use in the titration? Explain.
- Using 10.00mL samples of the ammonia unknown, and titrating with the appropriate acid, the average titre was found to be 12.76mL. Calculate the concentration of the ammonia solution.

**FOR MAXIMUM MARKS SHOW
FORMULAS & WORKING,
APPROPRIATE PRECISION & UNITS
IN ALL CHEMICAL PROBLEMS**

Alkanols and Alkanoic Acids

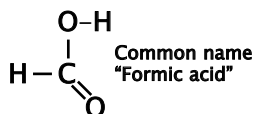
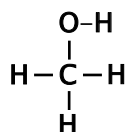
Differences Between these

Functional Groups

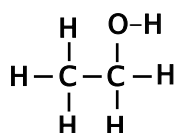
Alkanols

Alkanoic Acids

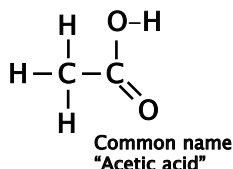
Methanoic acid
 HCOOH



Ethanoic acid
 CH_3COOH



The most important member of both Homologous Series is the 2-carbon "Eth-" compound



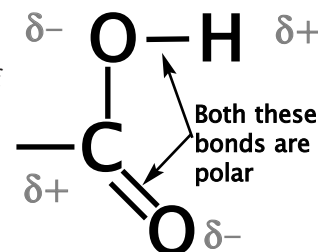
This bond is polar

δ^- O — H δ^+

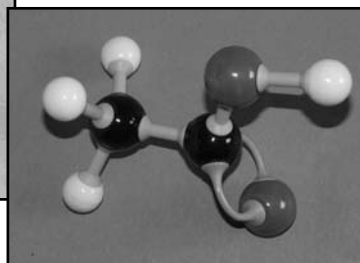
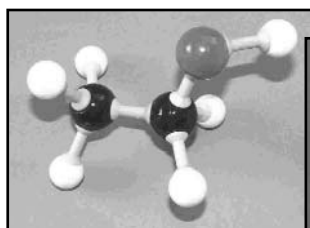
— CH₂

Now, compare the -COOH Functional Group of the Alkanoic Acids:

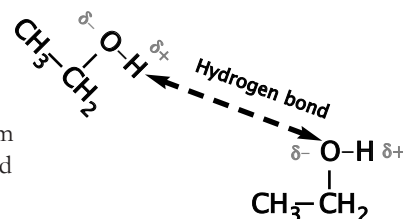
Because of the presence of another electronegative oxygen atom, there are two sets of dipoles on the molecule.



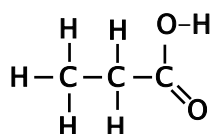
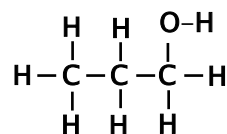
In an alkanol, the dipoles result in hydrogen bonding between molecules. For example, in **ethanol**:



Each pair of molecules can form one hydrogen bond

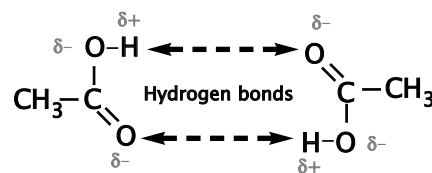


Propanoic acid
 $\text{CH}_3\text{CH}_2\text{COOH}$



In an alkanonic acid, such as **ethanoic acid**:

Each pair of molecules can form two hydrogen bonds.



You already know that the alkanols have much higher m.p.'s & b.p.'s than the corresponding alkanes, because of the hydrogen bonding between molecules.

The m.p's & b.p's of the alkanolic acids are higher still, due to the presence of twice as many hydrogen bonds.

For comparison:

	Ethane	Ethanol	Ethanoic acid
Boiling Pt ($^{\circ}\text{C}$)	-89	+78	+118

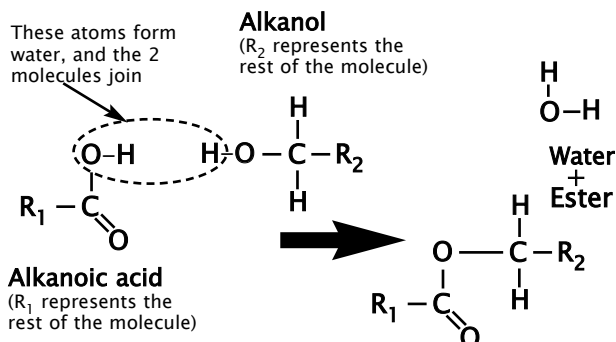
The syllabus requires you to know these as far as 8-carbon molecules; octanol and octanoic acid.

The Esters

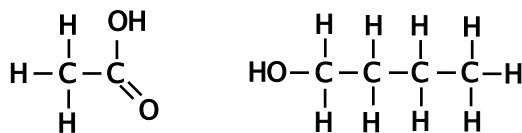
“Esters” are a group of carbon compounds formed by the reaction between an alkanol and an alkanic acid.

The reaction could be described as a “condensation” because it produces a water molecule, but it is so widespread in nature, and so important, that it rates its own name; “Esterification”.

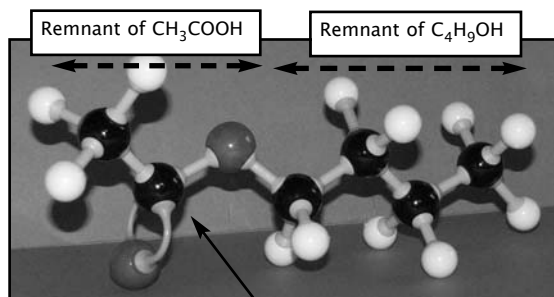
The reaction can be visualized as follows:



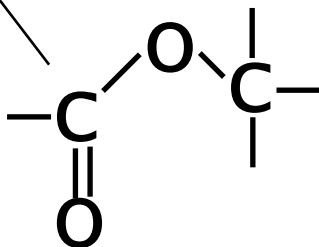
Example: If ethanoic acid reacted with butanol...



...the ester formed would look like this:



All esters contain this chemical “functional group” inside the molecule, with hydrocarbon chains on either side.



This group of atoms is polar, but both ends of the molecule are non-polar. Esters generally have low solubility in water (some exceptions), and are volatile with a strong odour... often sweet and fruity.

As you’ll learn, this is their “job”; to give the smells and flavours to many foods and perfumes.

Naming Esters

Since the esters are made by joining together 2 other molecules, it should be no surprise that they have a 2-part name.

- Alkanol name first.

Drop-off “-ANOL”, and add “-YL”.

In the example at lower left, “**butanol**” becomes “**butyl**”.

- Alkanoic acid name second.

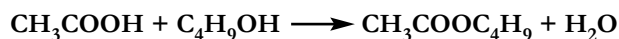
Drop-off “-IC ACID”, and add “-ATE”

In the example lower left, “**ethanoic acid**” becomes “**ethanoate**” (the same as the ion from this acid).

The name of the ester in the photo is

“butyl ethanoate”

ethanoic + butanol → butyl ethanoate + water acid

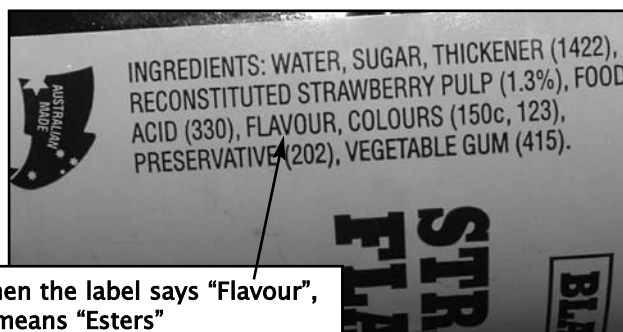


You need to be able to name any ester with up to 8 carbons on either end (nothing bigger than “octyl octanoate”)

Try the **WORKSHEET** at the end of section



The delicious smell and taste of ripe fruit is largely due to natural esters

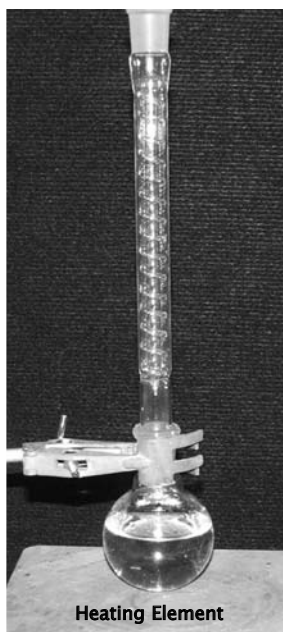


When the label says “Flavour”, it means “Esters”

Practical Work: Making an Ester

You may have made an ester in the laboratory using the process of Reflux.

A simple laboratory reflux set-up is shown.



Reflux condenser

is open to the atmosphere at the top.

This is vital to avoid any dangerous build-up of pressure which would occur in a sealed flask.

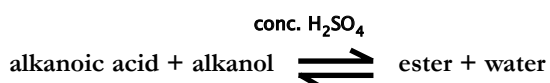
Volatile chemical vapours rise, but are condensed and drip back into the flask... "reflux".

Reaction Flask

is heated to speed up the otherwise very slow reaction.

Heating Element

As well as an alkanoic acid and an alkanol, it is usual to add a catalyst, concentrated sulfuric acid, H_2SO_4 .



As well as speeding the reaction up, the H_2SO_4 catalyst also absorbs the water product. This has the effect of shifting the equilibrium to the right (Le Chatelier's Principle) and increasing the yield of the ester.

Occurrence of Esters

Esters occur widely in nature, especially in fruits and flowers. Esters are largely responsible for the smell and taste of many foods.

There may be a complex mixture of esters and other compounds which give the complete smell of (say) a ripe strawberry, but there's always an ester giving the main smell and taste sensation.

Some examples:

Strawberry ethyl butanoate, $\text{C}_3\text{H}_7\text{COOC}_2\text{H}_5$

Orange octyl ethanoate, $\text{CH}_3\text{COOC}_8\text{H}_{17}$

Fats & Oils are Esters, Too

Esters made from very long chain "fatty acids", and the triple-alcohol molecule "glycerol" are used by all living things as high-energy foods and energy storage chemicals. We call them fats or oils, depending on their melting point.

Production and Uses of Esters

As well as their wide occurrence in living things, artificially-manufactured esters are important industrial chemicals.

They are produced by exactly the same process shown at the left... reflux of the appropriate alcohol, acid and catalyst... but on an industrial scale, of course.

Uses include:

- artificial flavours for drinks and various processed foods.
- solvents. Ethyl ethanoate is widely used as an industrial solvent (e.g. plastics industry) and is well known as the solvent for nail varnish.
- ingredients in many products, including shampoo and cosmetic products, and as "plasticizers" (softening agents) in some plastics, such as "vinyl".

Worksheet 5

Part A Fill in the Blanks

The Alkanols, also called a)....., all contain the functional group b)..... and have the general formula c)..... The -OH group contains a chemical bond which is d)....., and allows e)..... bonding between molecules. This is why the alkanols have m.p's & b.p's much higher than the corresponding f).....

The Alkanoic Acids contain the functional group g)..... This group contains 2 polar bonds, so 2 h)..... bonds can form between molecules. This is why i)..... than the alkanols.

Esters are formed by the reaction of j)..... with The other product is k)..... Esters are named by the l)..... first (with its ending changed to m).....), followed by the n)..... name (with its ending changed to o).....)

Esters are made using the technique of p)..... The reaction flask is open to the atmosphere to avoid any dangerous build-up of q)..... Volatile chemicals vapourize, but are r)..... by the s)..... and drip back into the flask. t)..... is used as a catalyst, and also improves the yield by shifting the u)..... because it absorbs water.

Esters occur widely in nature, being responsible for many of the v)..... and of foods, especially w)..... Long-chain esters of glycerol are the x)..... and

Artificially manufactured esters are used as y)..... in processed foods, as z)..... in industry and as ingredients in many products such as aa)..... and

Part B Practice problems

1. Names of Esters

Name the ester formed from

- ethanol & propanoic acid
- propanol & ethanoic acid
- pentanol & methanoic acid
- methanol and pentanoic acid
- hexanoic acid and butanol
- ethanoic acid and octanol

2. Condensed Structural Formulas

For each of the compounds above, give the condensed structural formula for the

- alkanol
 - alkanoic acid
- and
- ester

The first has been done for you as an example.

Answer

- i) ethanol = $\text{CH}_3\text{CH}_2\text{OH}$
- ii) propanoic acid = $\text{CH}_3\text{CH}_2\text{COOH}$
- iii) ester = $\text{CH}_3\text{CH}_2\text{COOCH}_2\text{CH}_3$

Note: although the alkanol comes first in naming, it may be more convenient to place the acid remnant first in the structural formula. This system is used here throughout.

3. Names from Structures

For each of the following esters:

- give the name of the ester
- name the alkanol and acid use to make it

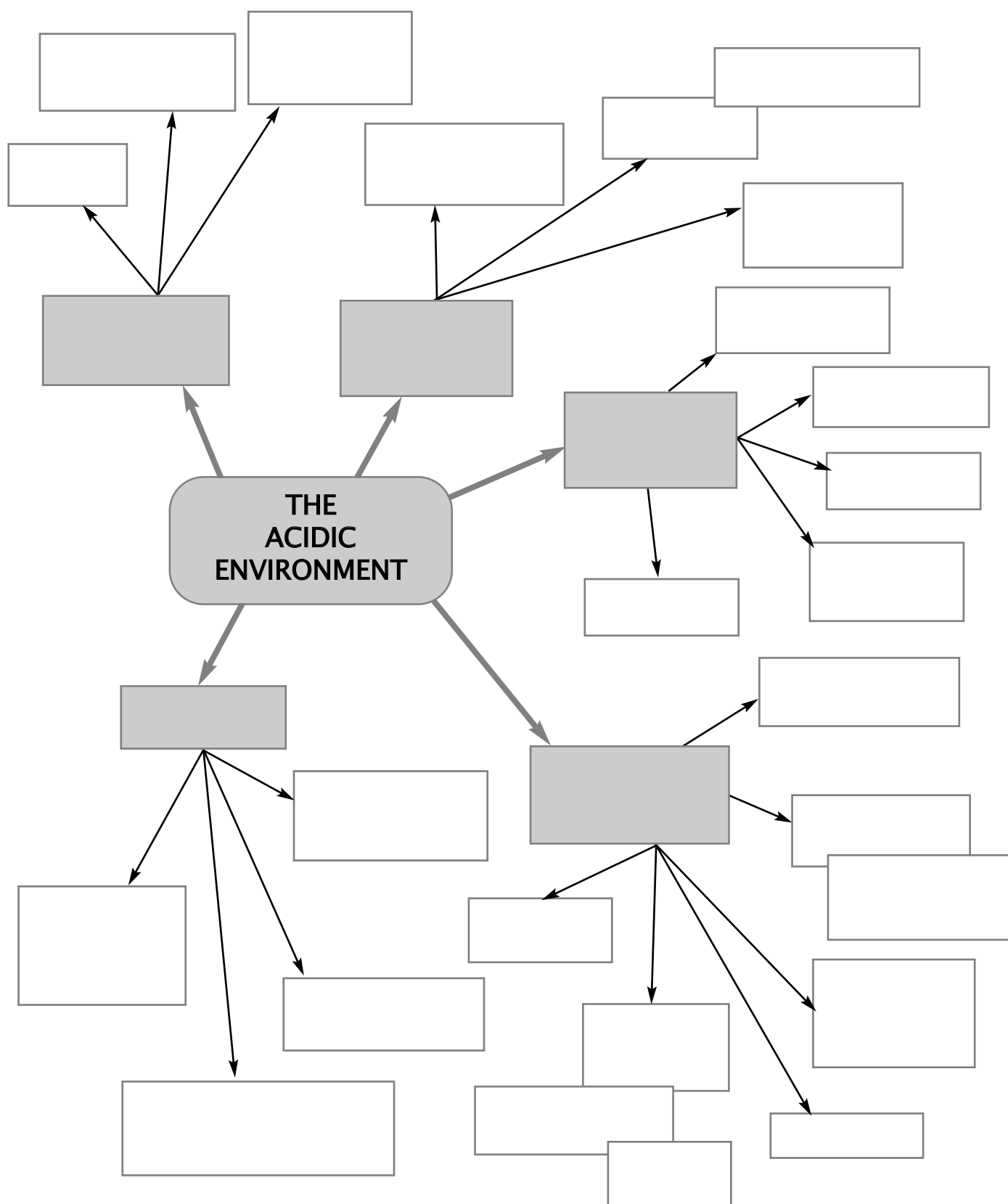
- $\text{HCOO}(\text{CH}_2)_3\text{CH}_3$
- $\text{CH}_3\text{CH}_2\text{COO}(\text{CH}_2)_3\text{CH}_3$
- $\text{CH}_3(\text{CH}_2)_3\text{COO}(\text{CH}_2)_4\text{CH}_3$
- $\text{C}_4\text{H}_9\text{COOCH}_3$
- $\text{C}_5\text{H}_{11}\text{COOC}_7\text{H}_{15}$

**WHEN COMPLETED, WORKSHEETS
BECOME SECTION SUMMARIES**

CONCEPT DIAGRAM (“Mind Map”) OF TOPIC

Some students find that memorizing the **OUTLINE** of a topic helps them learn and remember the concepts and important facts.

Practise on this blank version.



Practice Questions

These are not intended to be "HSC style" questions, but to challenge your basic knowledge and understanding of the topic, and remind you of what you NEED to know at the K.I.S.S. Principle level.

When you have confidently mastered this level, it is strongly recommended you work on questions from past exam papers.

Part A Multiple Choice

1.

The colour changes for some common indicators are shown:

Indicator	Colour in		Changes at pH
	Acid	Base	
Bromothymol	blue	yellow	6.2 - 7.6
Methyl orange	red	yellow	3.1 - 4.4
Phenolphthalein	clear	pink	8.3 - 10.0

In a solution with pH= 8.0 the colours of phenolphthalein, bromothymol blue and methyl orange (in order) would be:

- A. pink, blue, yellow
- B. clear, blue, red
- C. clear, yellow, yellow
- D. clear, blue, yellow

2.

A household substance most likely to be very basic is

- A. vinegar.
- B. soap.
- C. sugar.
- D. milk.

3.

The salt formed when nitric acid reacts with calcium hydroxide would be:

- A. calcium nitrate
- B. nitric hydroxide
- C. water
- D. calcium nitric

4.

An oxide compound, MO_2 (M is not the correct symbol) is found to react as follows:



It would be true to say that MO_2 is

- A. an acidic oxide
- B. a conjugate base
- C. amphoteric
- D. a basic oxide

5.

According to Le Chatelier's Principle, a chemical system in equilibrium, which is then disturbed, will adjust itself

- A. so that the disturbance is amplified.
- B. in the direction that releases energy.
- C. so that the disturbance is counteracted.
- D. in the direction that releases the pressure.

6.

In the reaction



the change that would shift the equilibrium to the right would be to:

- A. increase the pressure.
- B. increase the concentration of hydrogen.
- C. increase the temperature.
- D. remove NH_3 from the equilibrium mixture.

7.

A significant natural source of sulfur dioxide in the environment is:

- A. lightning storms.
- B. volcanic eruptions.
- C. burning of fossil fuels.
- D. forest fires.

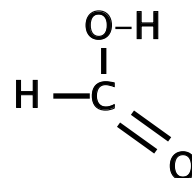
8.

The major environmental impact of NO_x gases is:

- A. Global Warming.
- B. Acid Rain.
- C. Ozone depletion.
- D. Smog.

9.

The diagram shows a molecule of the weak acid, methanoic acid.



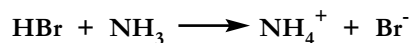
When dissolved in water, one of the bonds in the molecule may ionize.

The bond most likely to ionize is the

- A. H-C bond
- B. C=O bond
- C. O-H bond
- D. C-O bond

10.

In the reaction



it is true to say that:

- A. NH_3 is the base, because it has accepted a proton.
- B. HBr is the acid, because it has accepted a proton.
- C. NH_4^+ is the conjugate acid of HBr .
- D. Br^- is the conjugate acid of HBr .

11.

If an undissociated molecule of an acid is represented by "H-A", and the ionized acid by separate " H^+ " and " A^- " symbols, which diagram could show a dilute solution of a strong acid?

A.	B.	C.	D.
$\text{H}-\text{A}$ A^- H^+ $\text{H}-\text{A}$ $\text{H}-\text{A}$	H^+ A^- A^- H^+ H^+ H^+ A^- A^- H^+ H^+ A^- H^+	$\text{H}-\text{A}$ $\text{H}-\text{A}$ $\text{H}-\text{A}$ $\text{H}-\text{A}$	H^+ A^- H^+ A^-

12.

In a solution of pH = 10, the concentration of hydronium ions is:

- A. 10 molL^{-1} . B. $10^{10} \text{ molL}^{-1}$.
C. $10^{-10} \text{ molL}^{-1}$. D. 1 molL^{-1} .

13.

If you had 4 solutions of different acids

Hydrochloric	Nitric
Ethanoic	Sulfuric

all with exactly the same molar concentration of acid, which two only would you expect to have the same pH?

- A. sulfuric and ethanoic
B. hydrochloric and nitric
C. nitric and sulfuric
D. ethanoic and hydrochloric

14.

The hydrogen phosphate ion, HPO_4^{2-} is an amphiprotic species. If it were to act as a base, then its conjugate acid would be:

- A. H_2PO_4^- B. H_3PO_4 C. PO_4^{3-} D. HPO_3^{3-}

15.

An ionic "salt" is found to be acidic in water solution. It is likely that this salt is the product of the reaction between:

- A. a strong acid and a strong base.
B. a weak acid and a weak base.
C. a weak acid and a strong base.
D. a strong acid and a weak base.

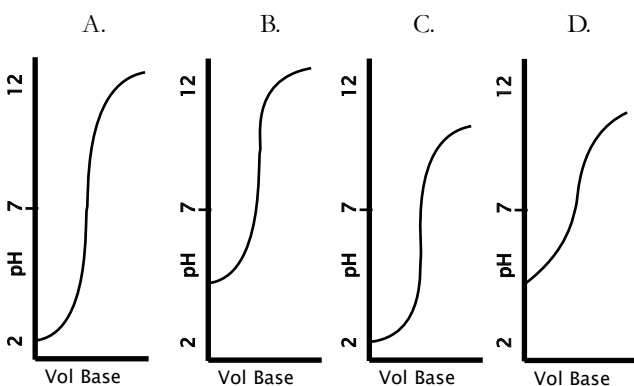
16.

The salt mentioned in Q15 was formed during a titration. The most appropriate indicator for the titration would be:
(hint: refer to the list given for Q1)

- A. bromothymol blue.
B. methyl orange.
C. phenolphthalein.
D. universal indicator.

17.

Which of the following graphs might be the "titration curve" for the titration described in Q15-16?



18.

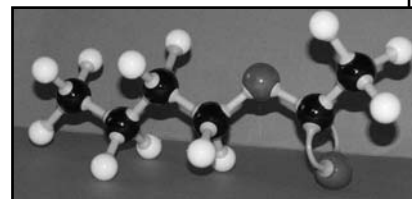
An appropriate material to use in case of an acid spill is:

- A. an amphiprotic substance, like sodium bicarbonate.
B. a strong base, like sodium hydroxide.
C. a buffer solution, like methanoic/methanoate mixture.
D. a weak base, like ammonia.

19.

The ester shown is:

- A. butyl propanoate
B. ethyl butanoate
C. butyl ethanoate
D. propyl butanoate



20.

Of these compounds, which would you expect to have the highest m.p. & b.p.?

- A. ethane B. ethene C. ethanol D. ethanoic acid

Part B Longer Response Questions

Mark values shown are suggestions only, and are to give you an idea of how detailed an answer is appropriate.

21. (4 marks)

- a) What is meant by an "acid-base indicator"?
b) Identify an everyday use of an indicator.
c) Describe how you could prepare a crude indicator from a named natural substance.

22. (7 marks)

- a) Write a balanced, symbol equation for the reaction between hydrochloric acid and magnesium oxide.
b) Explain how this equation supports the classification of magnesium oxide as a "basic oxide".
c) Write a balanced, symbol equation which shows that carbon dioxide may be considered as an "acidic oxide".

23. (8 marks)

- a) Explain why a sealed bottle of fizzy lemonade shows no signs of bubbles within the liquid, yet bubbles form immediately the lid is removed.
Include a relevant chemical equation in your answer.

- b) The reaction of carbon dioxide with water is exothermic. Given that information, predict the effect of higher temperature on the rate of bubble formation in a just-opened bottle of soft drink. Explain your prediction, naming the scientific principle involved.

24. (8 marks)

- a) Name a significant natural source of sulfur dioxide gas.
b) Use a balanced, symbol equation to explain how the smelting of some metal ores can produce sulfur dioxide.
c) Name a serious environmental effect that can be caused by sulfur dioxide, and describe a possible environmental impact it causes.
d) Calculate the volume of sulfur dioxide gas (measured at SLG) that can be formed from 1.00 tonne of sulfur.

25. (7 marks)

A certain diprotic acid can be represented by the formula " H_2A ". ("A" is not the correct symbol)

a) Write a balanced equation for the complete ionization of H_2A when added to water.

Label each species in the equation as acid or base, including conjugate species.

b) If H_2A is a strong acid, calculate the pH of a solution with concentration $[H_2A] = 0.0250 \text{ mol L}^{-1}$.

c) In fact, when this exact concentration solution was tested, it was found to have a $\text{pH} = 2.50$.

What do you conclude from this?

26. (4 marks)

Explain 2 different reasons for adding acids to processed foods. For each reason given, name a (different) acid used for that purpose.

27. (7 marks)

a) Give the definition of an acid according to the Arrhenius Theory.

b) Write an equation which Arrhenius might have used to explain why hydrogen chloride gas is an acid when dissolved in water.

c) Give the definitions for acid and base according to the Bronsted-Lowry Theory.

d) Write an equation to show why hydrogen chloride gas in water is an acid according to the B-L Theory.

28. (7 marks)

a) A solution containing carbonate ions (CO_3^{2-}) is found to be quite strongly basic.

Write an equation to explain why, and state the role of the water molecule in this reaction.

b) Write TWO different equations to show the amphiprotic nature of the hydrogen carbonate ion, HCO_3^- .

NOTICE ANY ERRORS?

Our material is carefully proof-read
but we're only human

If you notice any errors, please let us know

29. (8 marks)

A diluted vinegar sample (CH_3COOH solution) was analysed by titration with a standardized solution of KOH with concentration of $0.008263 \text{ mol L}^{-1}$.

25.00mL samples of the vinegar solution were used. The titration was done 4 times, giving burette titres of 27.35, 26.70, 26.75 and 26.65mL of KOH.

a) Choose the most appropriate indicator for this titration from the list shown in Q1, and justify your choice.

b) Write a balanced symbol equation for the reaction.

c) Explain how the titration measurements should be used.

d) Calculate the concentration of the diluted vinegar solution.

30. (4 marks)

a) What are the characteristics of a "buffer solution"?

b) In general terms, what are the usual ingredients of a buffer solution? Give a specific example.

c) Give an example of a natural buffer system.

31. (6 marks)

a) Draw structural formulas for

i) methanol

ii) propanoic acid

iii) the ester formed by reaction of the above,
and give the name of this ester.

b) Explain the need for using a reflux system to carry out this ester preparation.

c) Name the chemical you would add to the reaction flask as a catalyst.

d) Predict, in general terms, the odour of the ester.

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Answer Section

Worksheet 1

Part A

- | | |
|---------------------------|------------------------|
| a) opposites | b) neutralize |
| c) pH | d) 7 |
| e) basic | f) acidic |
| g) change colour | h) pH (acidity) |
| i) plants / lichens | |
| j) pink | k) blue |
| l) clear, clear, pink/red | m) red, yellow, yellow |
| n) yellow | o) blue |
| p) soil testing | q) water testing |
| r) effluents | |

Part B

A = acid. B = base. C = neutral. D = base. E = acid.

Worksheet 2

Part A

- | | |
|---|--------------------------------|
| a) water | b) salt |
| c) neutralize | d) bases |
| e) water and a salt | f) acids |
| g) acid solution | h) neutralize |
| i) equilibrium | j) rate |
| k) products | l) dynamic |
| m) temperature, concentration or gas pressure | |
| n) Le Chatelier's | o) counteracts the disturbance |
| p) carbon-oxygen | q) burning fossil fuels |
| r) Global | s) Greenhouse Effect |
| t) volcanic | u) hot springs/geysers |
| v) fossil fuels | w) smelting |
| x) Acid Rain | y) NOx |
| z) lightning | aa) power stations & engines |
| ab) smog | ac) Rain |

Part B

1. Simple Acid-Alkali Reactions

- a) i) calcium chloride
ii) magnesium sulfate
iii) barium nitrate
- b) i) $\text{HCl} + \text{LiOH} \longrightarrow \text{H}_2\text{O} + \text{LiCl}$
- ii) $\text{H}_2\text{SO}_4 + 2\text{NaOH} \longrightarrow 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4$
- iii) $2\text{HNO}_3 + \text{Mg}(\text{OH})_2 \longrightarrow 2\text{H}_2\text{O} + \text{Mg}(\text{NO}_3)_2$

2. Basic Oxides

- a) $\text{H}_2\text{SO}_4 + \text{FeO} \longrightarrow \text{H}_2\text{O} + \text{FeSO}_4$
- b) $2\text{HCl} + \text{MgO} \longrightarrow \text{H}_2\text{O} + \text{MgCl}_2$
- c) $2\text{HNO}_3 + \text{CuO} \longrightarrow 2\text{H}_2\text{O} + \text{Cu}(\text{NO}_3)_2$

3. Acidic Oxides

- a) $\text{CO}_2 + \text{Ca}(\text{OH})_2 \longrightarrow \text{H}_2\text{O} + \text{CaCO}_3$
- b) $\text{P}_2\text{O}_5 + 3\text{H}_2\text{O} \longrightarrow 2\text{H}_3\text{PO}_4$
- c) $\text{SO}_3 + \text{H}_2\text{O} \longrightarrow \text{H}_2\text{SO}_4$ (sulfuric acid)

4. Le Chatelier's Principle

- a) i) Left, to try to use heat & reduce temp.
ii) Right, to reduce the pressure again.
iii) Left, to decrease concentration of product.
iv) Right, to release heat & increase temp.
v) Left, to make more reactant & increase pressure.
vi) Left, to increase pressure again.
- b) i) Equilib. shifts left, to decrease conc. of iodide ion.
ii) Equilib. shifts right, to increase conc. of hydrogen ion.
iii) Equilib. shifts right, to decrease conc. of HI.
iv) Endothermic (heat is a reactant), since higher temp causes shift to right to consume heat.
- c) i) High pressure causes equilibrium to shift right to reduce the total moles of gas and pressure.
ii) Shifts equilibrium to left.
iii) Removing product & adding reactant keeps shifting equilib. to right, so yield is maximized.

5. Molar Gas volumes

- a) In each case, multiply moles by 24.8 (SLC) or 22.7 (STP)
i) 64.2L ii) 1.03L iii) $2.72 \times 10^3 \text{ L}$ iv) 1.16L
- b) In each case, divide volume by 24.8 or 22.7.
i) 0.500 mol ii) $1.10 \times 10^{-2} \text{ mol}$ iii) 403 mol iv) $4.41 \times 10^{-5} \text{ mol}$

6. Mass-Vol of Gases

- a) i) $n(\text{CO}_2) = 5.00/24.8 = 0.2016 \text{ mol}$
 $m(\text{CO}_2) = n \times \text{MM} = 0.2016 \times 44.01 = \mathbf{8.87g}$
ii) $n(\text{H}_2) = 5.00/24.8 = 0.2016 \text{ mol}$
 $m(\text{H}_2) = n \times \text{MM} = 0.2016 \times 2.016 = \mathbf{0.406g}$
iii) $n(\text{Ne}) = 100/22.7 = 4.405 \text{ mol}$
 $m(\text{Ne}) = n \times \text{MM} = 4.405 \times 20.18 = \mathbf{88.9g}$
iv) $n(\text{O}_2) = 0.0250/22.7 = 0.001101 \text{ mol}$
 $m(\text{O}_2) = n \times \text{MM} = 0.001101 \times 32 = \mathbf{3.52 \times 10^{-2} g}$
- b) i) $n = m/\text{MM} = 100/44.01 = 2.272 \text{ mol}$
 $v = 2.272 \times 24.8 = \mathbf{56.4L}$
ii) $n = m/\text{MM} = 100/4.003 = 24.98 \text{ mol}$
 $v = 24.98 \times 24.8 = \mathbf{624L}$
iii) $n = m/\text{MM} = 1.50/28.02 = 0.05353 \text{ mol}$
 $v = 0.05353 \times 24.8 = \mathbf{1.33L}$

7. Problems

- a) i) $\text{CO}_{2(g)} + \text{Ca}(\text{OH})_{2(aq)} \longrightarrow \text{H}_2\text{O}_{(l)} + \text{CaCO}_{3(s)}$
- ii) $n(\text{CO}_2) = 1.00/24.8 = 0.04032 \text{ mol}$
 $\therefore n(\text{CaCO}_3) = 0.04032 \text{ mol}$ (ratio 1:1)
 $\therefore m(\text{CaCO}_3) = n \times \text{MM} = 0.04032 \times 100.09 = \mathbf{4.04g}$
iii) $n(\text{CaCO}_3) = m/\text{MM} = 1.75/100.09 = 0.01748 \text{ mol}$
 $\therefore n(\text{CO}_2) = 0.01748 \text{ mol}$ (ratio 1:1)
 $\therefore \text{vol}(\text{CO}_2) = 0.01748 \times 22.7 = \mathbf{0.397 L}$ (397mL)
- b) i) $2\text{ZnS}_{(s)} + 3\text{O}_{2(g)} \longrightarrow 2\text{ZnO}_{(s)} + 2\text{SO}_{2(g)}$
- ii) $n(\text{ZnS}) = m/\text{MM} = 1.00 \times 10^6 / 97.46 = 10,260 \text{ mol}$
 $\therefore n(\text{O}_2) = 10,260 \times 3/2 = 15,391 \text{ mol}$ (ratio = 2:3)
 $\therefore \text{vol}(\text{O}_2) = 15,391 \times 24.8 = \mathbf{3.82 \times 10^5 L}$
- iii) $\text{vol}(\text{air}) = 3.82 \times 10^5 \times 100/21 = \mathbf{1.82 \times 10^6 L}$
- iv) Avogadro's Hypothesis: equal volumes of gases contain equal numbers of particles... the volume of SO_2 will be in the same ratio as the mole ratio of $\text{O}_2 : \text{SO}_2$

$$\therefore \text{vol}(\text{SO}_2) = 3.82 \times 10^5 \times 2/3 = \mathbf{2.55 \times 10^5 L}$$

Worksheet 3

Part A

- | | |
|--|---------------------------------------|
| a) donate protons | b) hydrogen |
| c) hydronium | d) monoprotic |
| e) diprotic | f) triprotic |
| g) CH ₃ COOH | |
| h) 2-hydroxypropane-1,2,3-tricarboxylic acid | |
| i) ionizes completely in water solution | |
| j) only partially ionizes | k) $-\log_{10}[\text{H}_3\text{O}^+]$ |
| l) powers (index numbers) | m) 10 |
| n) preserve | o) bacteria & fungi |
| p) SO ₂ and ethanoic | q) flavour |
| r) sour | s) ethanoic & citric |
| t) nutritional | u) ascorbic |
| v) formic (methanoic) | w) calcium oxide |

Part B

1. Acid Ionization in Water

- a) $\text{HCl}_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$
- b) $\text{HBr}_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{Br}^-_{(\text{aq})}$
- c) $\text{HCOOH}_{(\text{l})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{HCOO}^-_{(\text{aq})}$
- d) $\text{HCN}_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{CN}^-_{(\text{aq})}$
- e) $\text{H}_2\text{SO}_{4(\text{l})} + 2\text{H}_2\text{O}_{(\text{l})} \longrightarrow 2\text{H}_3\text{O}^+_{(\text{aq})} + \text{SO}_4^{2-}_{(\text{aq})}$
- f) $\text{H}_2\text{CO}_{3(\text{l})} + 2\text{H}_2\text{O}_{(\text{l})} \longrightarrow 2\text{H}_3\text{O}^+_{(\text{aq})} + \text{CO}_3^{2-}_{(\text{aq})}$
- g) $\text{H}_3\text{PO}_{4(\text{l})} + 3\text{H}_2\text{O}_{(\text{l})} \longrightarrow 3\text{H}_3\text{O}^+_{(\text{aq})} + \text{PO}_4^{3-}_{(\text{aq})}$
- h) $\text{C}_6\text{H}_8\text{O}_{7(\text{s})} + 3\text{H}_2\text{O}_{(\text{l})} \longrightarrow 3\text{H}_3\text{O}^+_{(\text{aq})} + \text{C}_6\text{H}_5\text{O}_7_{(\text{aq})}$

2. pH from [H₃O⁺]

use $\text{pH} = -\log[\text{H}_3\text{O}^+]$ in each case

- | | |
|--------------------|-----------------------------|
| a) 1.17 (acidic) | b) 3.08 (acidic) |
| c) -0.398 (acidic) | d) 11.6 (basic) |
| e) 2.46 (acidic) | f) 7.82 (just barely basic) |

3. pH from Acid Concentration

- a) $[\text{H}_3\text{O}^+] = 0.250$, so $\text{pH} = 0.602$
- b) $[\text{H}_3\text{O}^+] = 0.0750 \times 2$, so $\text{pH} = 0.824$
- c) $[\text{H}_3\text{O}^+] = 7.50 \times 10^{-4} \times 2$, so $\text{pH} = 2.82$
- d) $[\text{H}_3\text{O}^+] = 4.5 \times 10^{-3} \times 3$, so $\text{pH} = 1.87$
- e) $[\text{H}_3\text{O}^+] = 6.00 \times 2$, so $\text{pH} = -1.08$

4. [H₃O⁺] from pH

use inverse (or 2nd function) $\log(-\text{pH})$ on calculator

- | | |
|--|---|
| a) $5.01 \times 10^{-6} \text{ molL}^{-1}$ | b) $3.16 \times 10^{-12} \text{ molL}^{-1}$ |
| c) 2.51 molL^{-1} | d) $3.00 \times 10^{-9} \text{ molL}^{-1}$ |
| e) 1.00 molL^{-1} | |

Worksheet 4

Part A

- | | |
|------------------------|-----------------------|
| a) oxygen | b) hydrogen |
| c) metals | d) Arrhenius |
| e) hydrogen ions | f) hydroxide or oxide |
| g) the solvent (water) | h) Bronsted-Lowry |
| i) proton donor | j) proton acceptor |
| k) transfer | l) protons |
| m) conjugate | n) conjugate acid |
| o) acid or base | p) amphiprotic |

Worksheet 4 (continued)

- | | |
|--------------------------------------|-----------------------------|
| q) water | r) any ammonium salt |
| s) sodium carbonate | t) sodium chloride/sulfate |
| u) strong | v) strong |
| w) strong | x) weak |
| y) weak | z) strong |
| aa) hydronium & hydroxide | ab) water |
| ac) exo- | ad) Titration |
| ae) volume | af) standard |
| ag) equivalence (end) | ah) burette |
| ai) indicator | aj) colour |
| ak) pH | al) near-vertical |
| am) weak acid and its conjugate base | |
| an) shifts | ao) Le Chatelier's |
| ap) pH | aq) living |
| ar) constant chemical conditions | |
| as) in the blood | at) bicarbonate & carbonate |

Part B

1. Reactions of Acids & Bases with Water

In each case the conjugate is in bold print.

- a) $\text{CH}_3\text{COOH}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{CH}_3\text{COO}^-_{(\text{aq})}$
- b) $\text{NH}_{3(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{NH}_4^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
- c) $\text{HPO}_4^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{PO}_4^{3-}_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$
- d) $\text{HPO}_4^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{H}_2\text{PO}_4^-_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
- e) $\text{S}^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HS}^-_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
- f) $\text{CN}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HCN}_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
- g) $\text{H}_2\text{S}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HS}^-_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$
- h) $\text{NO}_2^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HNO}_{2(\text{aq})} + \text{OH}^-_{(\text{aq})}$
- i) $\text{NH}_4^+_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{NH}_{3(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$
- j) $\text{HSO}_3^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{SO}_3^{2-}_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$

2. Amphiprotic Substances

- a) $\text{H}_2\text{PO}_4^-_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})} \longrightarrow \text{H}_3\text{PO}_{4(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$
- $\text{H}_2\text{PO}_4^-_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \longrightarrow \text{HPO}_4^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$
- b) $\text{HCO}_3^-_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})} \longrightarrow \text{H}_2\text{CO}_{3(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$
- $\text{HCO}_3^-_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \longrightarrow \text{CO}_3^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$
- c) $\text{HS}^-_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})} \longrightarrow \text{H}_2\text{S}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$
- $\text{HS}^-_{(\text{aq})} + \text{OH}^-_{(\text{aq})} \longrightarrow \text{S}^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$

3. Acidic & Basic Salts

- a) $\text{CH}_3\text{COO}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{OH}^-_{(\text{aq})} + \text{CH}_3\text{COOH}_{(\text{aq})}$
- b) $\text{NH}_4^+_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{NH}_{3(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$
- c) $\text{NO}_2^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HNO}_{2(\text{aq})} + \text{OH}^-_{(\text{aq})}$
- d) $\text{HC}_2\text{O}_4^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{C}_2\text{O}_4^{2-}_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$
- e) $\text{CN}^-_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HCN}_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$

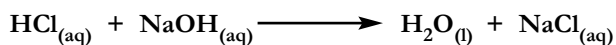
Worksheet 4

Part B (continued)

4. Titrations

a) 23.10 omitted

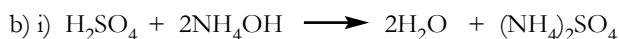
$$\text{Average titre} = (22.50 + 22.45 + 22.50) / 3 = 22.48 \text{ mL}$$



$$\frac{\text{Ca} \times \text{Va}}{a} = \frac{\text{Cb} \times \text{Vb}}{b}$$

$$\begin{aligned} \text{Cb} &= b \times \text{Ca} \times \text{Va} / (\text{Vb} \times a) \\ &= 1 \times 0.09255 \times 22.48 / 25.00 \times 1 \\ &= 0.08322 \end{aligned}$$

$$\therefore c(\text{NaOH}) = 0.08322 \text{ mol L}^{-1}$$



ii) $\frac{\text{Ca} \times \text{Va}}{a} = \frac{\text{Cb} \times \text{Vb}}{b}$

$$\begin{aligned} \text{Ca} &= a \times \text{Cb} \times \text{Vb} / (\text{Va} \times b) \\ &= 1 \times 0.05025 \times 28.32 / (25.00 \times 2) \\ &= 0.02846 \end{aligned}$$

$$\therefore c(\text{H}_2\text{SO}_4) = 0.08322 \text{ mol L}^{-1}$$

iii) L is best choice because strong acid-weak base titration has end-point at acidic pH. L changes colour at 4.2.

c) i) Can be obtained in a pure, dry state. Is stable and does not react with gases in air or absorb moisture.

ii) moles required:

$$n = C \times V = 0.02500 \times 0.5 = 0.01250 \text{ mol}$$

mass required:

$$\begin{aligned} m &= n \times \text{MM} \quad (\text{MM} = 90.04 \text{ g}) \\ &= 0.01250 \times 90.04 \end{aligned}$$

$$\text{mass} = 1.126 \text{ g}$$

iii) Weigh out chemical into clean, dry beaker.

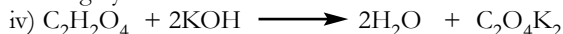
Add enough pure water, and stir, to dissolve it completely.

Transfer solution into 500 mL volumetric flask.

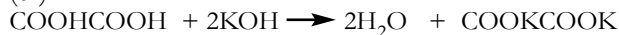
Rinse beaker with small amounts of extra water and add washings to flask.

Fill flask to mark with pure water. Use a dropper at the end to fill exactly to the mark.

Insert stopper and invert repeatedly to mix solution thoroughly.



(or)

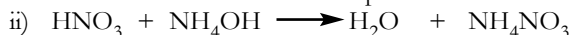


$$\frac{\text{Ca} \times \text{Va}}{a} = \frac{\text{Cb} \times \text{Vb}}{b}$$

$$\begin{aligned} \text{Cb} &= b \times \text{Ca} \times \text{Va} / (\text{Vb} \times a) \\ &= 2 \times 0.02500 \times 31.45 / 25.00 \times 1 \\ &= 0.06290 \end{aligned}$$

$$\therefore c(\text{KOH}) = 0.06290 \text{ mol L}^{-1}$$

d) i) use nitric acid because it is strong acid. Weak acid-weak base titrations have indistinct end-point... best avoided.



$$\begin{aligned} \text{Cb} &= b \times \text{Ca} \times \text{Va} / (\text{Vb} \times a) \\ &= 1 \times 0.7438 \times 12.76 / 10.00 \times 1 \\ &= 0.9491 \end{aligned}$$

$$\therefore c(\text{NH}_4\text{OH}) = 0.9491 \text{ mol L}^{-1}$$

Worksheet 5

Part A

- | | |
|---|----------------------|
| a) alcohols | b) -OH |
| c) $\text{C}_n\text{H}_{2n+1}\text{OH}$ | d) polar |
| e) hydrogen | f) alkanes |
| g) COOH | h) hydrogen |
| i) m.p. & b.p. is even higher than | |
| j) alkanols with alkanolic acids | |
| k) water | l) alkanol |
| m) -yl | n) alkanolic acid |
| o) -oate | p) reflux |
| q) pressure | r) condensed |
| s) reflux condenser | t) Sulfuric acid |
| u) equilibrium | v) odours and tastes |
| w) fruits | x) fats and oils |
| y) artificial flavourings | z) solvents |
| aa) shampoo/cosmetics/plastics | |

Part B

1. Names of Esters

- | | |
|----------------------|----------------------|
| a) ethyl propanoate | b) propyl ethanoate |
| c) pentyl methanoate | d) methyl pentanoate |
| e) butyl hexanoate | f) octyl ethanoate |

2. Condensed Structural Formulas

- | | |
|---|---|
| a) i) $\text{CH}_3\text{CH}_2\text{OH}$ | ii) $\text{CH}_3\text{CH}_2\text{COOH}$ |
| iii) $\text{CH}_3\text{CH}_2\text{COOCH}_2\text{CH}_3$ | |
| b) i) $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ | ii) CH_3COOH |
| iii) $\text{CH}_3\text{COO}(\text{CH}_2)_2\text{CH}_3$ | |
| c) i) $\text{CH}_3(\text{CH}_2)_3\text{CH}_2\text{OH}$ | ii) HCOOH |
| iii) $\text{HCOO}(\text{CH}_2)_4\text{CH}_3$ | |
| d) i) CH_3OH | ii) $\text{CH}_3(\text{CH}_2)_3\text{COOH}$ |
| iii) $\text{CH}_3(\text{CH}_2)_3\text{COOCH}_3$ | |
| e) i) $\text{CH}_3(\text{CH}_2)_2\text{CH}_2\text{OH}$ | ii) $\text{CH}_3(\text{CH}_2)_4\text{COOH}$ |
| iii) $\text{CH}_3(\text{CH}_2)_4\text{COO}(\text{CH}_2)_3\text{CH}_3$ | |
| f) i) $\text{CH}_3(\text{CH}_2)_6\text{CH}_2\text{OH}$ | ii) CH_3COOH |
| iii) $\text{CH}_3\text{COO}(\text{CH}_2)_7\text{CH}_3$ | |

3. Names from Structures

- | | |
|-----------------------|---------------------------|
| a) butyl methanoate. | butanol + methanoic acid |
| b) butyl propanoate. | butanol + propanoic acid |
| c) pentyl penanoate. | pentanol + pentanoic acid |
| d) methyl pentanoate. | methanol + pentanoic acid |
| e) heptyl hexanoate. | heptanol + hexanoic acid. |

Practice Questions

Part A Multiple Choice

- | | | | | |
|------|------|-------|-------|-------|
| 1. D | 5. C | 9. C | 13. B | 17. C |
| 2. B | 6. B | 10. A | 14. A | 18. A |
| 3. A | 7. B | 11. D | 15. D | 19. C |
| 4. A | 8. D | 12. C | 16. B | 20. D |

Part B Longer Response

In some cases there may be more than one correct answer possible. The following "model" answers are correct, but not necessarily perfect.

21.

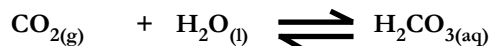
- A chemical which changes colour depending on the pH of the solution it is in.
- Soil testing, to help gardening or agriculture.
- Put Hydrangea flowers through a blender with a small amount of water and ethanol. Filter the mixture. The liquid will act as an acid-base indicator.

22.

- a) $2\text{HCl}_{(\text{aq})} + \text{MgO}_{(\text{s})} \longrightarrow \text{H}_2\text{O}_{(\text{l})} + \text{MgCl}_{2(\text{aq})}$
 b) It has neutralized an acid, forming water and a salt. Therefore it is a base.
 c) $\text{CO}_{2(\text{g})} + 2\text{NaOH}_{(\text{aq})} \longrightarrow 2\text{H}_2\text{O}_{(\text{l})} + \text{Na}_2\text{CO}_{3(\text{aq})}$
 Carbon dioxide has neutralized a base, therefore it is acidic.

23.

- a) Carbon dioxide reacts with water as follows:



In the sealed bottle, this system is in equilibrium. When the lid is removed, the pressure of CO_2 above the liquid drops. The equilibrium shifts left (attempting to increase the pressure by making more gas) so bubbles form as CO_2 comes out of solution.

- b) If exothermic, then heat is a "product" of the reaction as written above. By Le Chatelier's Principle, higher temperature should shift equilibrium left. Therefore, a warm bottle of drink will form bubbles faster when opened.

24.

- a) volcanic eruptions, hot springs, geysers.
 b) example: smelting of lead sulfide ore:
 $\text{PbS} + \text{O}_2 \longrightarrow \text{Pb} + \text{SO}_2$
 c) "Acid Rain" can acidify lakes and streams, killing the living things. Forests can die-back due to soil acidity.

- d) $\text{S} + \text{O}_2 \longrightarrow \text{SO}_2$
 $n(\text{S}) = m/\text{MM} = 1.00 \times 10^6 / 32.07$
 $= 31,182 \text{ mol}$

$$\therefore n(\text{SO}_2) = 31,182$$

$$m(\text{SO}_2) = n \times \text{MM} \quad \text{MM} = 64.07 \text{ g}$$

$$= 31,182 \times 64.07$$

$$= 2.00 \times 10^6 \text{ g} \quad (= 2.00 \text{ tonne})$$

25.

- a) $\text{H}_2\text{A} + 2\text{H}_2\text{O}_{(\text{l})} \rightleftharpoons 2\text{H}_3\text{O}^+_{(\text{aq})} + \text{A}^{2-}_{(\text{aq})}$
 acid base conj.acid conj.base

- b) if $[\text{H}_2\text{A}] = 0.0250$, then $[\text{H}_3\text{O}^+] = 0.0500 \text{ mol L}^{-1}$
 $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$
 $= 1.30$

- c) H_2A must be a weak acid and has only partially ionized. The $[\text{H}_3\text{O}^+]$ is lower and pH higher than predicted.

26.

Acids may be added to preserve the food. By lowering the pH, it becomes more difficult for microbes to grow in the food and cause it to spoil. Ethanoic acid (vinegar) is commonly used for this.

Acids are also added to flavour the food by adding the sourness that improves some flavours such as fruit drinks or jams. Citric acid is often used this way.

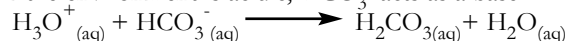
27.

- a) An acid produces hydrogen ions in solution.
 b) $\text{HCl}_{(\text{g})} \longrightarrow \text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$
 c) Acids are proton donors. Bases are proton acceptors.
 d) $\text{HCl}_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{Cl}^-_{(\text{aq})} + \text{H}_3\text{O}^+_{(\text{aq})}$
 The HCl molecule has transferred a proton to the water molecule, therefore it is an acid.

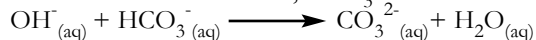
28.

- a) $\text{CO}_3^{2-}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \longrightarrow \text{HCO}_3^-_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
 Water acts as an acid and donates a proton to the carbonate ion. This forms a hydroxide ion, which explains why the solution is basic.

- b) If the environment is acidic, HCO_3^- acts as a base:



If the environment is basic, HCO_3^- acts as an acid:



29.

- a) A weak acid-strong base titration has an end point about $\text{pH} = 8-10$, so phenolphthalein is best.
 b) $\text{CH}_3\text{COOH} + \text{KOH} \longrightarrow \text{H}_2\text{O} + \text{CH}_3\text{COOK}$
 c) The first titre should be discarded because it does not agree closely with the others. The remaining 3 should be averaged. Average = $(26.70 + 26.75 + 26.65) / 3 = 26.70 \text{ mL}$
 d) $\frac{\text{Ca} \times \text{Va}}{a} = \frac{\text{Cb} \times \text{Vb}}{b}$

$$\text{Ca} = a \times \text{Cb} \times \text{Vb} / (\text{Va} \times b)$$

$$= 1 \times 0.008263 \times 26.70 / (25.00 \times 1)$$

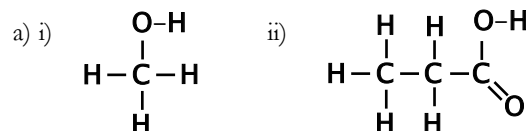
$$= 0.008825$$

$$\therefore c(\text{CH}_3\text{COOH}) = 0.08322 \text{ mol L}^{-1}$$

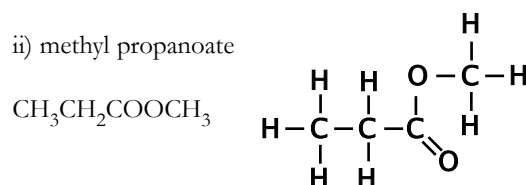
30.

- a) A buffer can maintain a constant pH despite addition of acid or base.
 b) A (roughly equal) mixture of a weak acid and its conjugate base, such as ethanoic acid plus sodium ethanoate (which provides ethanoate ions).
 c) Our blood is buffered by a mixture of bicarbonate ions and carbonate ions. The blood pH remains quite constant, despite constant changes occurring as gases dissolve, food is absorbed, etc.

31.



- ii) methyl propanoate



- b) The mixture needs heating, but a closed flask could explode due to pressure build up. An open flask is needed, but volatile chemicals will then evaporate away. The reflux system is open, but condenses vapours and returns them to the reaction flask.

- c) concentrated sulfuric acid.

- d) Esters generally have sweet, fruity odours.

**FOR MAXIMUM MARKS SHOW
 FORMULAS & WORKING,
 APPROPRIATE PRECISION & UNITS
 IN ALL CHEMICAL PROBLEMS**

