

Year 12 Chemistry

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Chapter 1

Module 5 Equilibrium and Acid Reactions

1.1 Practical Investigation 2.1

Aim: To determine whether chemical reactions are reversible or not

1.1.1 Materials

- 25 mL dropper bottle of 1 mol L^{-1} cobalt(II) chloride hexahydrate
- 25 mL dropper bottle of 0.05 mol L^{-1} potassium chromate
- 25 mL dropper bottle of 0.05 mol L^{-1} potassium dichromate
- 25 mL dropper bottle of 0.1 mol L^{-1} hydrochloric acid
- 25 mL dropper bottle of 0.1 mol L^{-1} sodium hydroxide
- $1 \times 5 \text{ cm}$ piece of magnesium ribbon
- Distilled water
- 1 piece of filter paper ($55 \text{ mm} \times 55 \text{ mm}$)
- 3 watch glasses
- 1 drying oven/incubator
- 4 test tubes
- Test-tube rack
- 4 small labels
- 1 pair brass tongs
- $1 \times (5 \text{ cm} \times 5 \text{ cm})$ piece of sandpaper
- $1 \times (5 \text{ cm} \times 5 \text{ cm})$ piece of steel wool
- Gas lighter
- Dropper
- Video camera
- Safety glasses and gloves

1.1.2 Risk Assessment

Hazard	Precaution
Shattering glassware	Keep beakers, test tubes, and watch glasses in centre of table
Exposure to harmful chemicals	Wear gloves and eye glasses. Handle with caution
Burns from Bunsen burner	Keep on safety flame when not in use

1.1.3 Method

Part A

1. Place a piece of filter paper on a watch glass.
2. Add cobalt chloride drop by drop until the filter paper is covered.
3. Observe the colour of the filter paper.
4. Place the watch glass into a drying oven overnight at 35 °C.
5. Remove the watch glass and filter paper and observe the colour of the filter paper.
6. Add distilled water drop by drop to the same filter paper until it is covered.
7. Observe the colour of the filter paper.
8. Repeat steps 4 and 5.

Part B

1. Label four test tubes A, B, C and D.
2. Add about 1 mL of potassium chromate to test tubes A and B.
3. Add about 1 mL of potassium dichromate to test tubes C and D.
4. Test tubes A and C are reference solutions.
5. Add hydrochloric acid dropwise to test tube B until a colour change occurs.
6. Record your observations.
7. Add sodium hydroxide dropwise to test tube B until another colour change occurs.
8. Record observations.
9. Add sodium hydroxide dropwise to test tube D until a colour change occurs.
10. Record observations.
11. Add hydrochloric acid dropwise to test tube D until another colour change occurs.
12. Record observations

Part C

1. Clean a 5 cm piece of magnesium with sandpaper.
2. Hold the piece of magnesium ribbon with a pair of brass tongs.
3. Light the magnesium ribbon and hold it over a watch glass. Do not look directly at the magnesium while it is alight.
4. Record observations.

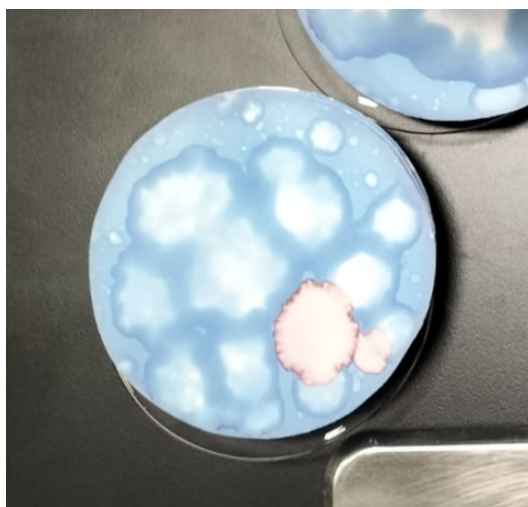
Part D

1. Hold a piece of steel wool with a pair of brass tongs.
2. Light the steel wool and hold it over a watch glass.
3. Record observations

1.1.4 Results

Part A

- When dehydrated, filter paper was blue when saturated with cobalt chloride
- When rehydrated, became pink



Part B

- Potassium chromate \rightarrow initially bright yellow
- Potassium dichromate \rightarrow initially orange
 - HCL \rightarrow greenish-yellow
 - NaOH \rightarrow darker yellow/orange
- $2\text{CrO}_4^{2-} + 2\text{H}^+ \rightleftharpoons \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}$
- $\text{K}_2\text{Cr}_2\text{O}_7 + 2\text{KOH} \rightleftharpoons 2\text{K}_2\text{CrO}_4 + \text{H}_2\text{O}$

Part C

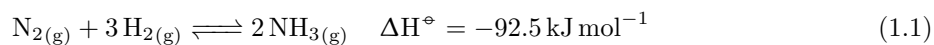
Part D

1.2 Le Chatelier's Principle

"If a system at equilibrium is subject to a change in conditions, then the system will behave in such a way so as to partially counteract the imposed change"

Haber process:

1.2.1 Effect of Concentration



1.2.2 Effect of Pressure

1.2.3 Effect of Partial Pressure

1.2.4 Effect of Volume

Decreasing the volume will increase the pressure. (Boyle's Law) This increases the collision rate between the reactants and favours the forward reaction.

1.2.5 Effect of Temperature

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1.2.6 Summary

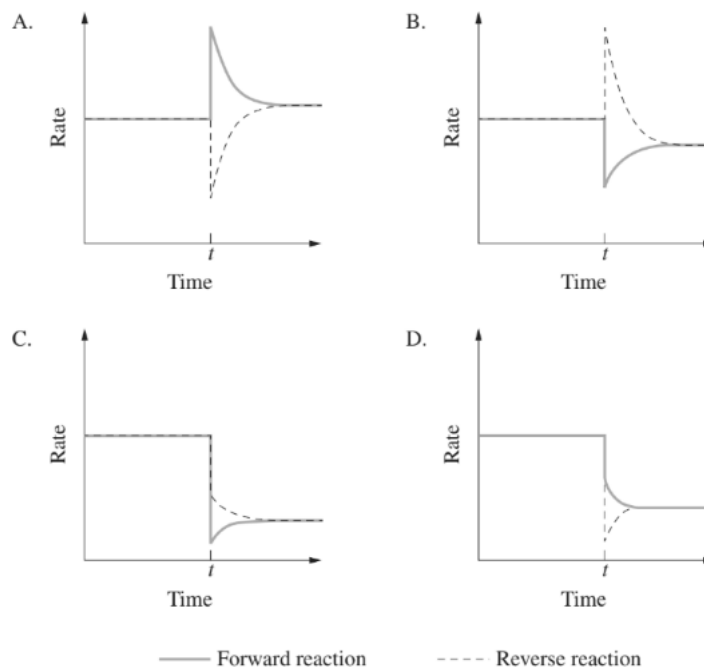
To use Le Chatelier's principle to predict the outcome of a change in conditions, you need to consider the following points.

1. What change is imposed?
2. What is the opposite of the change?
3. Which reaction direction is favoured - the forward or reverse?
4. Does equilibrium shift to the left or right?
5. What happens to the concentrations of each aqueous substance or gas?

- 19 Nitrogen dioxide reacts to form dinitrogen tetroxide in a sealed flask according to the following equation.



Which graph best represents the rates of both the forward and reverse reactions when an equilibrium system containing these gases is cooled at time t ?



D When temperature decreases, the rates of both forward and backward reactions will decrease regardless of which way the endothermic or exothermic reaction goes. (A and B can be eliminated)

This is because all the particles in the system lose kinetic energy, decreasing the rate of collisions hence, decreasing the rate of reaction.

However, since there is a decrease in temperature the exothermic reaction will be favoured in order to counteract the change. In this case, the forward reaction being exothermic is affected less by the drop in temperature as shown in D.

1.3 Practical Investigation 2.3 - Effect of changes to concentration on equilibrium

Aim: To observe the effect of a change in concentration on a system at equilibrium

1.3.1 Materials

- 2 mL of 0.1 molL^{-1} iron(III) chloride solution
- 2 mL of 0.1 molL^{-1} ammonium thiocyanate solution
- 1 mL of 0.1 molL^{-1} calcium fluoride solution
- 20 mL distilled water
- 2x 10 mL measuring cylinders
- 25 mL measuring cylinder
- 4 test tubes
- Test-tube rack
- 4 small labels
- Disposable 1 mL droppers
- Waste bottle
- Digital camera
- Safety glasses

1.3.2 Risk Assessment

Hazard	Precaution
Chemicals may splash onto skin or eyes	Wear safety glasses and wash hands
Chemicals may harm aquatic life	Place in inorganic waste container

1.3.3 Method

1. Pour 1 mL of iron(III) chloride solution into a 10 mL measuring cylinder.
2. Pour 1 mL of ammonium thiocyanate into another 10 mL measuring cylinder.
3. Pour both solutions into the 25 mL measuring cylinder.
4. Add 18 mL of distilled water to the 25 mL measuring cylinder so that the total volume is 20 mL.
5. Label four test tubes A, B, C and D.
6. Pour equal volumes of the solution in the 25 mL measuring cylinder into each of the test tubes.
7. Retain test tube A as the reference solution.
8. Add 1 mL of iron(III) chloride to test tube B.
9. Take a photo to record observations for test tube B relative to test tube A.
10. Add 1 mL of ammonium thiocyanate to test tube C.

11. Take a photo to record observations for test tube C relative to test tube A.
12. Add 1 mL of calcium fluoride to test tube D. (Note: This reacts with the iron(III) ion so there is less iron(III) available to react with the thiocyanate ion.)
13. Take a photo to record observations for test tube D relative to test tube A

1.3.4 Results

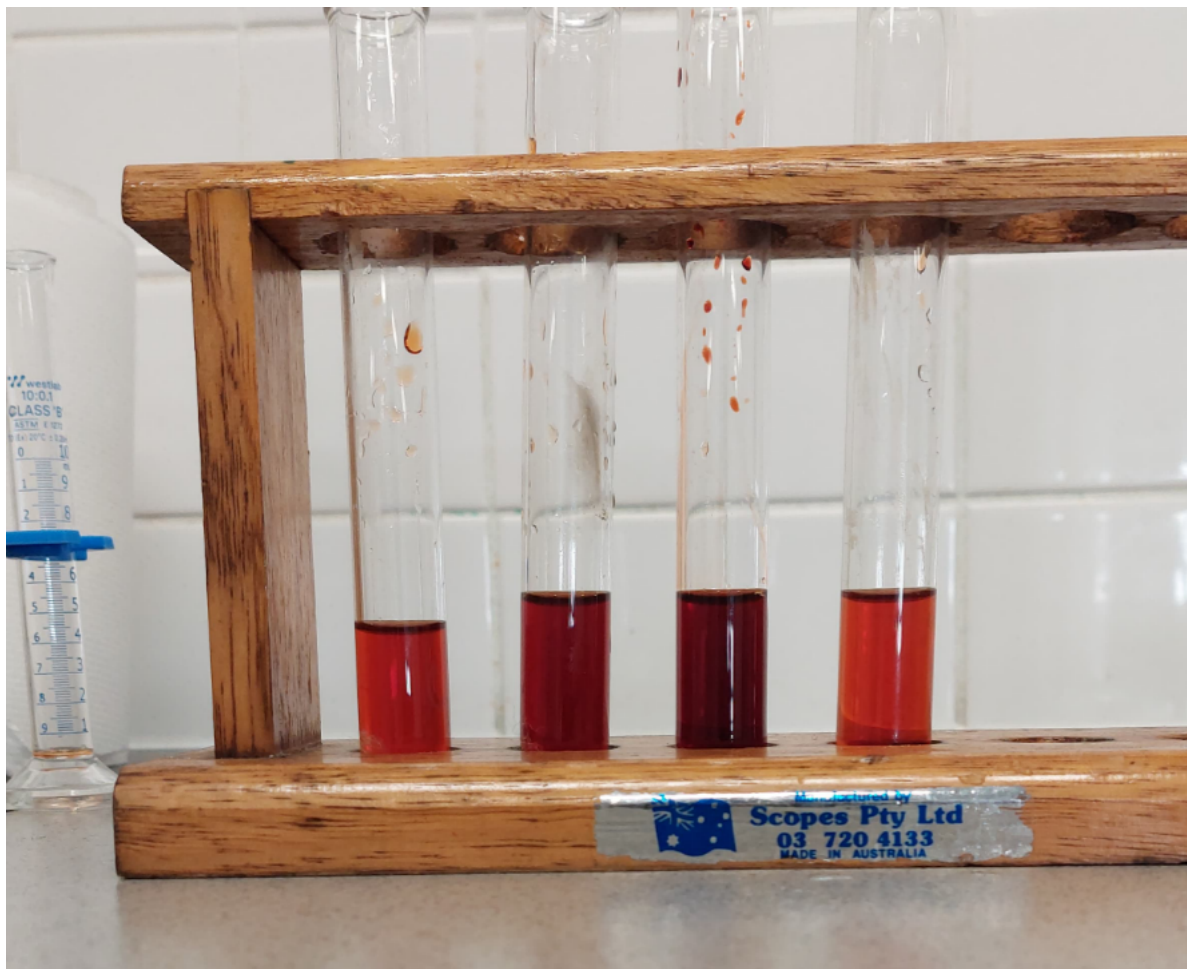


Figure 1.1: Test tubes A, B, C, D

1.3.5 Discussion

Explain each colour change in terms of collision theory.

The test tube B was darker in colour in comparison to test tube A. The increase in moles of reactants allows more successful collisions to occur, increasing the amount of product. The same principle applies to test tube C.

Test tube D was lighter in colour compared to A, due to the calcium fluoride reacting with the iron (III) chloride

1.3.6 Conclusion

Use Le Chatelier's principle to explain what happened in test tubes B, C and D.

Test tube B was darker due to the increase in concentration of the reactant iron (III) chloride causes a shift of the equilibrium towards the products due Le Chatelier's principle

Test tube C was darker due to the increase in concentration of the reactant ammonium thiocyanate causes a shift of the equilibrium towards the products due Le Chatelier's principle

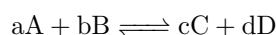
Test tube D was lighter because the calcium fluoride reacted with the iron (III) chloride, lowering the overall concentration of iron (III) chloride. This reduced the amount of reactants available, making the reverse reaction more favourable by Le Chatelier's principle.

1.4 Calculating the Equilibrium Constant

The equilibrium constant can be used to predict the direction of chemical reactions

$$K_{eq} = \frac{[products]}{[reactants]}$$

For reaction:



the equilibrium expression is:

$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The concentration of each chemical species is raised to the power of the number of moles of that species indicated in the chemical equation. (Eg. there are d moles of species D , hence in the equilibrium expression the concentration of species D is raised to the power of d , written as $[D]^d$)

- The value for the equilibrium constant only takes into account the concentration of substances where the concentration can vary
- Solutions and gases can vary in concentration or partial pressure hence are included in K_{eq}
- Solids and pure liquids are NOT included eg. H_2O isn't required when calculating

1.4.1 Reaction Quotient

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Has the same formula as K_{eq} , however applies to any stage of a reaction

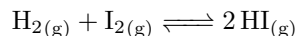
1.4.2 Equilibrium Constant

- Comparing Q to K_{eq} predicts which way the equilibrium will shift
- K_{eq} is where the equilibrium lies
- A large K_{eq} means that there are more products than reactants; ie. equilibrium lies towards completion
- If K_{eq} is close to one, both reactants and products are plentiful at equilibrium

Example Let $Q = 2.1$, $K_{eq} = 0.315$, \therefore products > reactants.

1.4.3 Calculating the equilibrium expression

Consider the reaction between hydrogen and iodine producing hydrogen iodide:



1.5 ICE Tables

	[A]	[B]	[C]
Initial concentration			
Change in concentration			
Equilibrium concentration			

$$\Delta c = c_{eq} - c_u$$

Eg. $\text{A} + \text{B} \rightleftharpoons \text{C} + \text{D}$

	[A]	[B]	[C]	[D]
Initial concentration	0.6	0.6	0	0
Change in concentration	-0.5	-0.5	+0.5	+0.5
Equilibrium concentration	0.1	0.1	0.5	0.5

Eg. $2 \text{X}_{(g)} \rightleftharpoons 3 \text{Y}_{(g)} + 4 \text{Z}_{(g)}$

A sample consisting of 0.500 mol of X is placed into a system with a volume of 0.750 litres.
At equilibrium, the amount of sample X is known to be 0.350 mol.

	X	Y	Z
I	0.5	0	0
C			
E	0.35		

$$\begin{aligned}[X] &= \frac{0.35}{0.75} = 0.467 \\[Y] &= \frac{0.225}{0.75} = 0.3 \\[Z] &= \frac{0.3}{0.75} = 0.4\end{aligned}$$

1.6 Effect of Temperature on the Equilibrium Constant

Although other factors may affect equilibrium, K_{eq} is only affected by temperature. Changing concentration, pressure, or volume will change the concentrations and therefore adjust the reaction point, however the reaction will still equalise to achieve the same K_{eq}

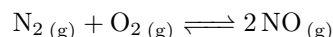
- For a particular reaction, K_{eq} is constant at a given temperature
- Temperature changes the ratio of products and reactants, hence changing K_{eq}
- For $\text{N}_2\text{O}_{2(g)} \rightleftharpoons 2 \text{NO}_{2(g)}$, temperature increases the K_{eq} value and favours the formation of products. The forward reaction is endothermic

	X	Y	Z
I	0.5	0	0
C	-0.15	-0.225	+0.3
E	0.35	0.225	0.3

1.6.1 Example Question

Nitric oxide gas (NO) can be produced from the direct combination of nitrogen gas and oxygen gas in a reversible reaction.

1. Write a balanced chemical equation for this reaction (1 mark)



2. Explain, using collision theory, how an increase in temperature would affect the value for K_{eq} for this system. Refer to the diagram in your answer.

An increase in temperature would favour the forward reaction, hence K_{eq} will increase. More energy allows more collisions to occur

$$K_{eq} = \frac{[p_1][p_2]}{[r_1][r_2]}$$

$$1 = \frac{[1][1]}{[1][1]}$$

If p_2 decreases to [0.5], favouring the forward reaction

$$= \frac{[1.15][0.65]}{[0.85][0.85]}$$

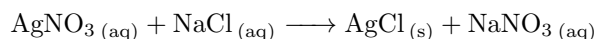
$$= K_{eq}$$

1.7 Applications of the Equilibrium Constant

1.7.1 Use of the Equilibrium Constant for the Dissociation of Ionic Solutions

Different ionic compounds have different solubilities

Example reaction



Complete ionic equation: $\text{Ag}^+ + \text{NO}_3^- + \text{Na}^+ + \text{Cl}^- \longrightarrow \text{AgCl} (\text{s}) + \text{NO}_3^- + \text{Na}^+$

Net ionic equation: $\text{Ag}^+ (\text{aq}) + \text{Cl}^- (\text{aq}) \longrightarrow \text{AgCl} (\text{s})$

Although a precipitate is formed, the reaction rests at a dynamic equilibrium where the rate at which the precipitate is formed is equal to the rate at which the ions are formed.



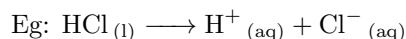
- By general practice, the solid precipitate is written on the left and the ions on the right.
- $K_{sp} = [\text{Ag}^+][\text{Cl}^-]$

- Since the solid is not included in the equilibrium constant, it is referred to as the **solubility product** (K_{sp})
- If the system is not at equilibrium, it is referred to as the **ionic product**
- If ionic product = K_{sp} , then the system is at equilibrium.
- If ionic product < K_{sp} , the forward reaction would be favoured and the solid would dissolve for the system to reach equilibrium.
- If ionic product > K_{sp} , the reverse reaction would be favoured and more precipitate would form for the system to reach equilibrium.

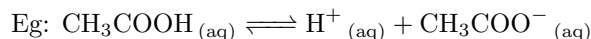
1.7.2 Use of the Equilibrium Constant for the Dissociation of Acids and Bases

Acids

Strong acids dissociate completely in solution. The reaction goes to completion and is not an equilibrium system



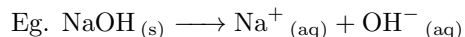
Weak acids do not dissociate completely, instead forming an equilibrium system



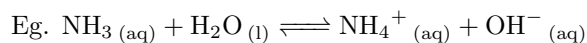
The acid dissociation constant is expressed as $K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$

Bases

Strong bases dissociate completely in solution.



Weak bases do not dissociate completely, only some of the molecules react with water to form ions, they form an equilibrium system



The base dissociation constant is expressed as $K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$

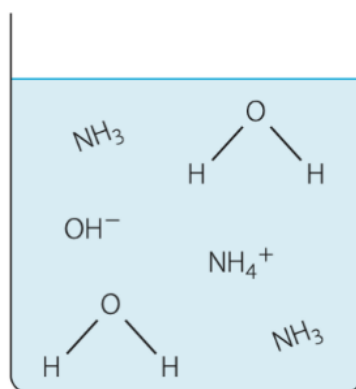
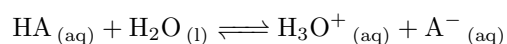


FIGURE 3.3 Some molecules of ammonia react with water to produce ammonium ions and hydroxide ions. This is an example of a dissociation reaction.

Weak Acid

General equation:

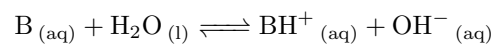


General expression for equilibrium constant:

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

Weak Base

General equation:



General expression for equilibrium constant:

$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

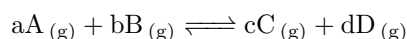
1.7.3 Use of the Equilibrium Constant for Gaseous Systems

In a gaseous system (where all reactants are gases), the partial pressure of each species is related to its concentration ($c = \frac{n}{V}$)

Pressure is due to the collisions between the gases and the walls of the container, therefore all particles within the system contribute to the pressure.

The **partial pressure** is the proportion of the pressure due to collisions for a particular gas species

For the general equation:



, where a , b , c and d are the number of moles

$$\text{Mole fraction of gas A} = \frac{\text{Number of moles of gas A}}{\text{Total number of moles of gas present}}$$

$$\text{Partial pressure of gas A} = \text{Mole fraction of gas A} \times \text{Total pressure of the system}$$

Expression of equilibrium constant in terms of pressure:

$$K_p = \frac{P_A^a \times P_B^b}{P_C^c \times P_D^d}$$

1.8 Beer Lambert Law

Absorbance (for a given wavelength) = Molar absorptivity \times Path length \times Concentration

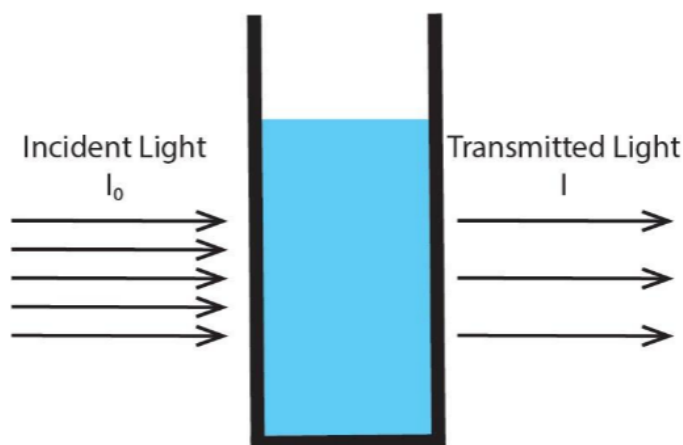
$$A = \epsilon lc$$

- Absorbance has a direct relationship to concentration
- Greater concentration, greater absorbance

Absorbance can be measured using a spectrophotometer (for all wavelengths) or a colourimeter (for the visible spectrum). Using the formula, the absorbance coefficient (ϵ) for a given material can be calculated. The coefficient determines how far light can penetrate a material before it is absorbed, depending on the material and the wavelength being absorbed. Measured in $\text{L mol}^{-1} \text{cm}^{-1}$

A filter is applied to the light and must be the complement to the colour of the solution.

Eg: For reaction $\text{Fe}_3^+_{(\text{aq})} + \text{SCN}^-_{(\text{aq})} \rightleftharpoons \text{FeSCN}_2^+_{(\text{aq})}$ FeSCN is red, \therefore a green filter is needed



$$A = \log_{10} \frac{I_o}{I} = \epsilon lc$$

1.8.1 Worked Example

A solution thickness of 1 cm transmits 30% incident light.

1. Calculate the concentration of the solution given the molar absorptivity of the solution being $4000 \text{ L mol}^{-1} \text{ cm}^{-1}$

$$A = \log_{10} \frac{100}{30} = 0.523$$
$$\therefore C = \frac{A}{\epsilon l} = \frac{0.523}{4000 \times 1}$$
$$= 1.31 \times 10^{-4} \text{ mol L}^{-1}$$

2. Calculate the molar absorptivity of a $1 \times 10^{-4} \text{ mol L}^{-1}$ solution which has an absorbance of 0.20, when the path length is 2.5 cm.

$$A = \epsilon l c$$
$$0.2 = 1 \times 10^{-4} \times 2.5 \times \epsilon$$
$$\epsilon = \frac{0.2}{1 \times 10^{-4} \times 2.5}$$
$$= 800 \text{ L mol}^{-1} \text{ cm}^{-1}$$

3. Which instrument is used in the verification of Lambert's Beer's law?

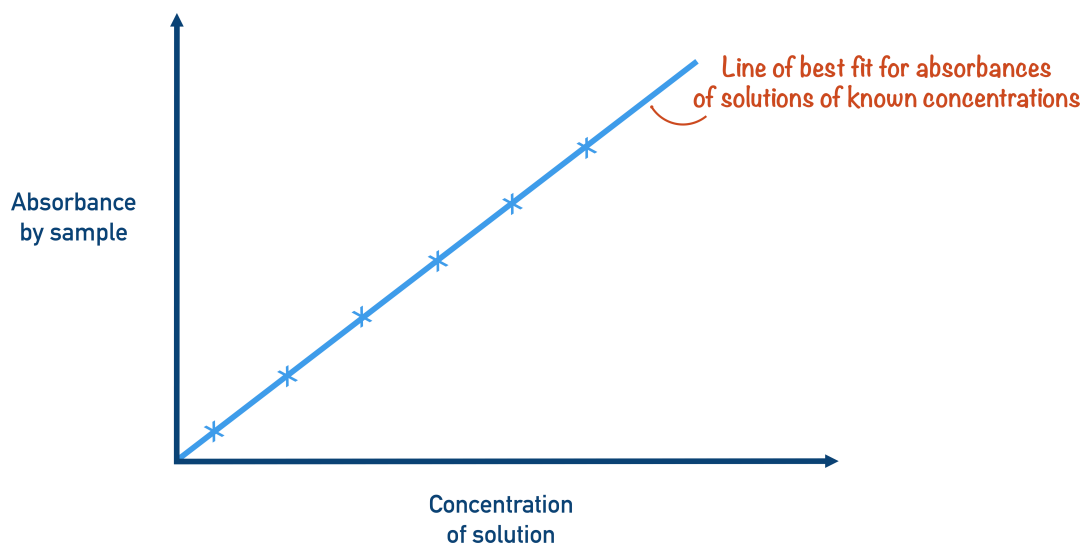
Colourimeter

4. Find out the molar absorptivity of a $1 \times 10^{-4} \text{ mol L}^{-1}$ solution with an absorbance of 0.30, when the path length is 1.5 cm. (3 marks)

$$A = \epsilon l c$$
$$0.3 = \epsilon \times 1.5 \times 1 \times 10^{-4}$$
$$\epsilon = 2000 \text{ L mol}^{-1} \text{ cm}^{-1}$$

1.8.2 Colourimeters

In order to determine concentration of a solution from its absorbance, a calibration curve using known concentrations and absorbances must be used to interpolate or extrapolate datapoints corresponding to a desired value.



Beer Lambert's Law collapses at high and low concentrations, where $c < 20\%$, $c > 80\%$

1.9 Practical Investigation 3.2

Aim: To use colourimetry to determine the equilibrium constant for the reaction of iron(III) ions with thiocyanate ions to form the iron(III) thiocyanate ion

1.9.1 Materials

- 200 mL $0.2 \text{ mol L}^{-1} \text{ Fe(NO}_3)_3$
- 100 mL $0.002 \text{ mol L}^{-1} \text{ KSCN}$
- 500 mL $0.5 \text{ mol L}^{-1} \text{ HNO}_3$
- 60 mL $0.002 \text{ mol L}^{-1} \text{ Fe(NO}_3)_3$
- 150 mL distilled water
- 6 * 100 mL volumetric flasks
- 5 * 150 mL beakers
- 2 * 100 mL beakers
- 1 * 25 mL bulb pipette
- 2 * 10 mL graduated pipettes
- 1 * 10 mL bulb pipette
- 2 pipette bulbs
- 1 disposable pipette
- Waste bottle
- 14 small labels
- 1 colourimeter and set of cuvettes
- Safety glasses

1.9.2 Risk Assessment

Hazard	Precaution
Breaking glassware	Keep glassware on inside of table, do not run with glassware
Spillage of solutions	Handle with caution, clean any spills immediately
Splashing of solution into eyes	Wear safety goggles

1.9.3 Method

1. Label the six volumetric flasks A-F.
2. Use a 25 mL bulb pipette to transfer 25.00 mL of the $0.200 \text{ mol L}^{-1} \text{ Fe(NO}_3)_3$
3. 3 to flask A.
4. Use a graduated 10 mL pipette to transfer 1.00 mL of the 0.002 mol L^{-1}

5. KSCN to flask A.
6. Add HNO₃ to make a final volume of 100.00 mL.
7. Make solutions with known concentration by pushing equilibrium as far as possible to the products. HNO₃ can be used to reduce the concentration of H₃O
8. Rinse the cuvette with distilled water.
9. $\frac{3}{4}$ fill the cuvette with distilled water and wipe the clear sides.
10. Turn on the colourimeter and turn the light to blue or 470 nm.
11. Use the cuvette with distilled water to calibrate the colourimeter. Note: Orientate the cuvette correctly in the colourimeter so that the light passes through the clear sides of the cuvette.
12. Rinse, a 100 mL beaker with standard solution A - it is easier to pour the solution into the cuvette using a beaker than using a volumetric flask.
13. Rinse, then $\frac{3}{4}$ fill the cuvette with standard solution A and measure the absorbance with the colourimeter.
14. Repeat steps 10 and 11 for the other standard solutions (B-F)

1.10 Solution Equilibria - Dissolution of Ionic Compounds

1.10.1 Solubility Revision

- Ionic compounds consist of a positive cation and negative anion
- Due to the polar nature of water molecules, they can arrange themselves around the ions and overcome the ionic bonding, dissolving the substance into an aqueous solution
- When surrounded by water, the ion becomes a **hydrated ion**
- For an ionic compound to dissolve, the energy required has to be less than or similar to the energy released when hydrated

$$\text{Energy to break lattice} \approx \text{Energy released when ions hydrated}$$

Solubility can be measured in g L⁻¹, g/100 g, or g/100 mL

1.10.2 Factors influencing solubility

- Activation energy required to break lattice
- Strength of ionic bonding
- Size of ions
- Charge of ions

1.10.3 Solubility

- The solubility of a solute is the maximum mass in grams that can dissolve in 100 g of the solvent at a given temperature (g L^{-1} can also be used)
- A solution is saturated when no more solute will dissolve at a given temperature
- Heat will increase the solubility of a solute
- Solubility curves show how much of a solute dissolves at a given temperature

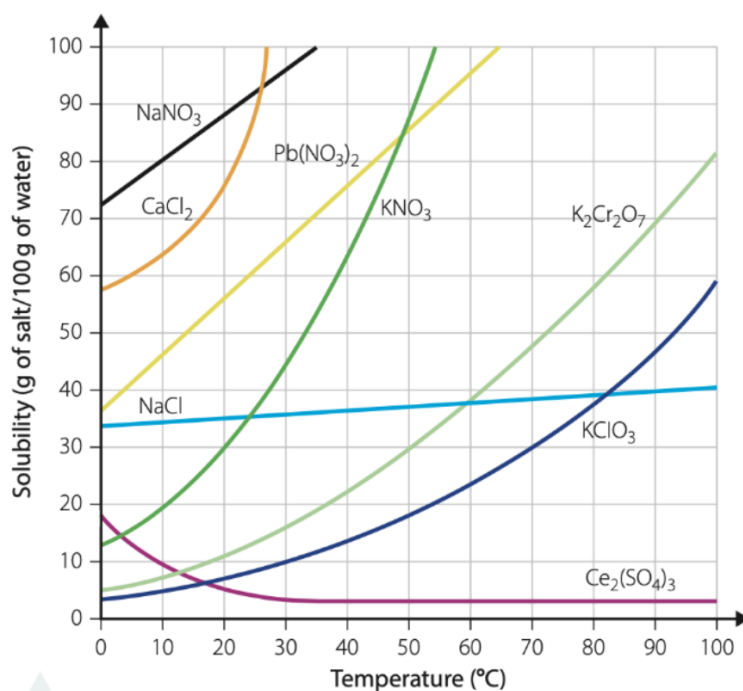


FIGURE 4.8 A selection of solubility curves for various salts

1.10.4 Water of Crystallisation

Water of crystallisation occurs when water molecules are attracted to the ions of a salt

Eg. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ is a substance with five water molecules of crystallisation per unit of copper (II) sulphate

Water or other molecules that form dipole bonds to a metal atom are referred to as ligands

The water of crystallisation can be evaporated by heating the hydrated compound and the product is said to be anhydrous

1.10.5 Toxins in Cycad Fruit

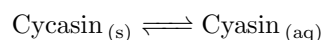
Cycad plants are native Australian trees that produce fruits that contain seeds with cone-like structures. There are several types, however most fruits and seeds are poisonous.

They contain two main toxins; cycasin and beta-methylamino-L-alanine (BMAA)

Cycasin and BMAA are types of azoxy glycosides, a group of toxins known to cause severe gastrointestinal issues in humans. These toxins can cause severe liver disorders and impact the functioning of nerves, leading to ataxia.

Cycasin and BMMA are water-soluble and can be dissolved out of the seeds and into the surrounding water.

Leaching involves placing the fruit in water and leaving it to soak, removing the toxins

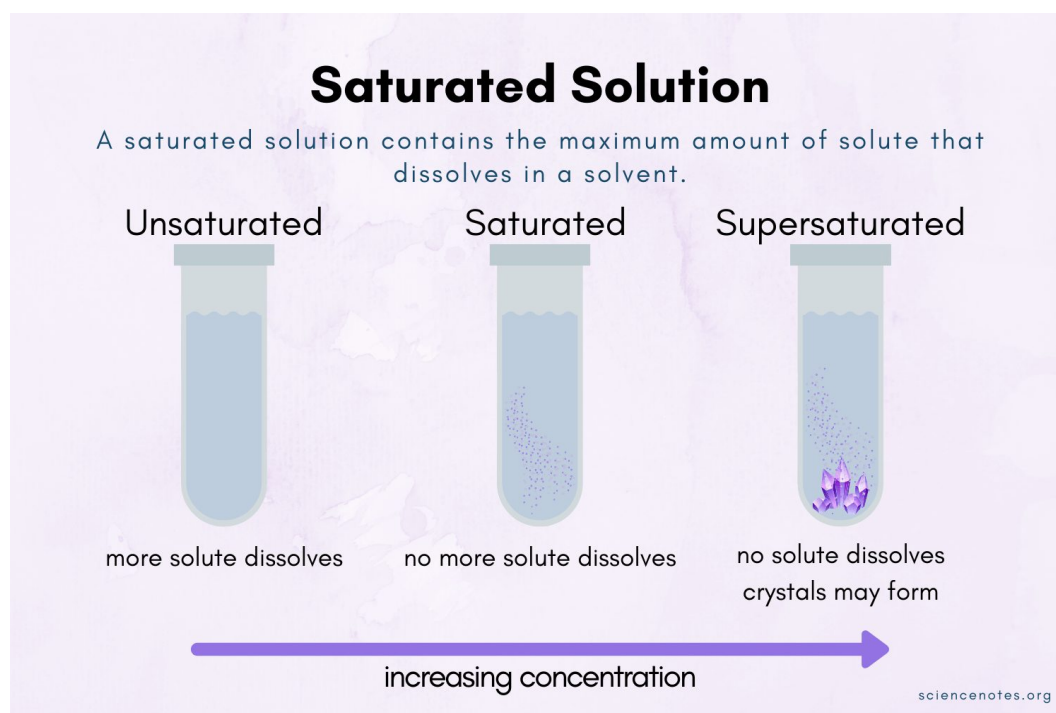


In the above equilibrium, the concentration of toxins in the water is initially zero. By Le Chatelier's Principle, the equilibrium shifts to the right, decreasing the high concentration of solid toxins in the fruit, eventually reaching a dynamic equilibrium. Although some toxins are removed, the system reaches an equilibrium where toxins are still present. To improve the result, running water can be used (ie. an open system) so that the system can never reach an equilibrium. The equilibrium will constantly shift to the right, hence removing most toxins from the seed and fruit.

Other Methods **Cooking** cycad fruit causes the toxins to decompose due to the intense heat. **Fermenting** cycad fruit sees natural processes break toxins down over time.

1.11 Measuring Solubility

- The mass of a substance that dissolves depends upon temperature.
- The solubility of a solute is the maximum mass in grams that can dissolve in 100 g of the solvent at a given temperature.
- Unsaturated solution
- Saturated - solid stays at bottom
- Supersaturated - more solute dissolved than in a saturated solution at the same temperature.
 - If this supersaturated solution is bumped, sugar crystal is added or the side of the glass is scratched, then the extra sugar will precipitate out again.



1.11.1 Solubility Rules

- Solutions of substances dissolved in water are called aqueous solutions. The term "aqueous" comes from the Latin aqua, meaning water.

When ionic substances dissolve in water, they dissociate. This means they separate into their ions, which are then able to move freely and independently of each other through the solution.

Although most ionic compounds are soluble in water, they do not all dissolve to the same extent.

"Soluble" means that a compound dissolves to more than 10 g L⁻¹ (or 1 g/100 mL),

"insoluble" means that it dissolves to less than 1 g L⁻¹, and

"sparingly soluble" means that it dissolves in the range 1 g L⁻¹ to 10 g L⁻¹.

1.12 Solubility Product

1.13 Practical Investigation 4.2 - Deriving the solubility curve for potassium chloride

Aim: To gather data to draw a solubility curve for potassium chloride

1.13.1 Materials

- Potassium chloride
- Distilled water
- 250 mL
- 10 mL measuring cylinder
- Large test tubes
- Bunsen burner
- Tripod
- Gauze mat
- Boss head and clamp
- Retort stand
- -10-110 °C thermometer
- Stirring rod
- Balance
- Weighing bottle
- Matches
- Test-tube rack
- Wire gauze
- Spatula
- Safety glasses

1.13.2 Risk Assessment

Hazard	Precaution
Burning from Bunsen burner	Do not use the Bunsen burner if the gas tube is damaged.
Poison from chemicals	Do not drink chemicals
Spillage of chemicals	Keep beakers in centre of table. Handle with caution

1.13.3 Scientific Diagram

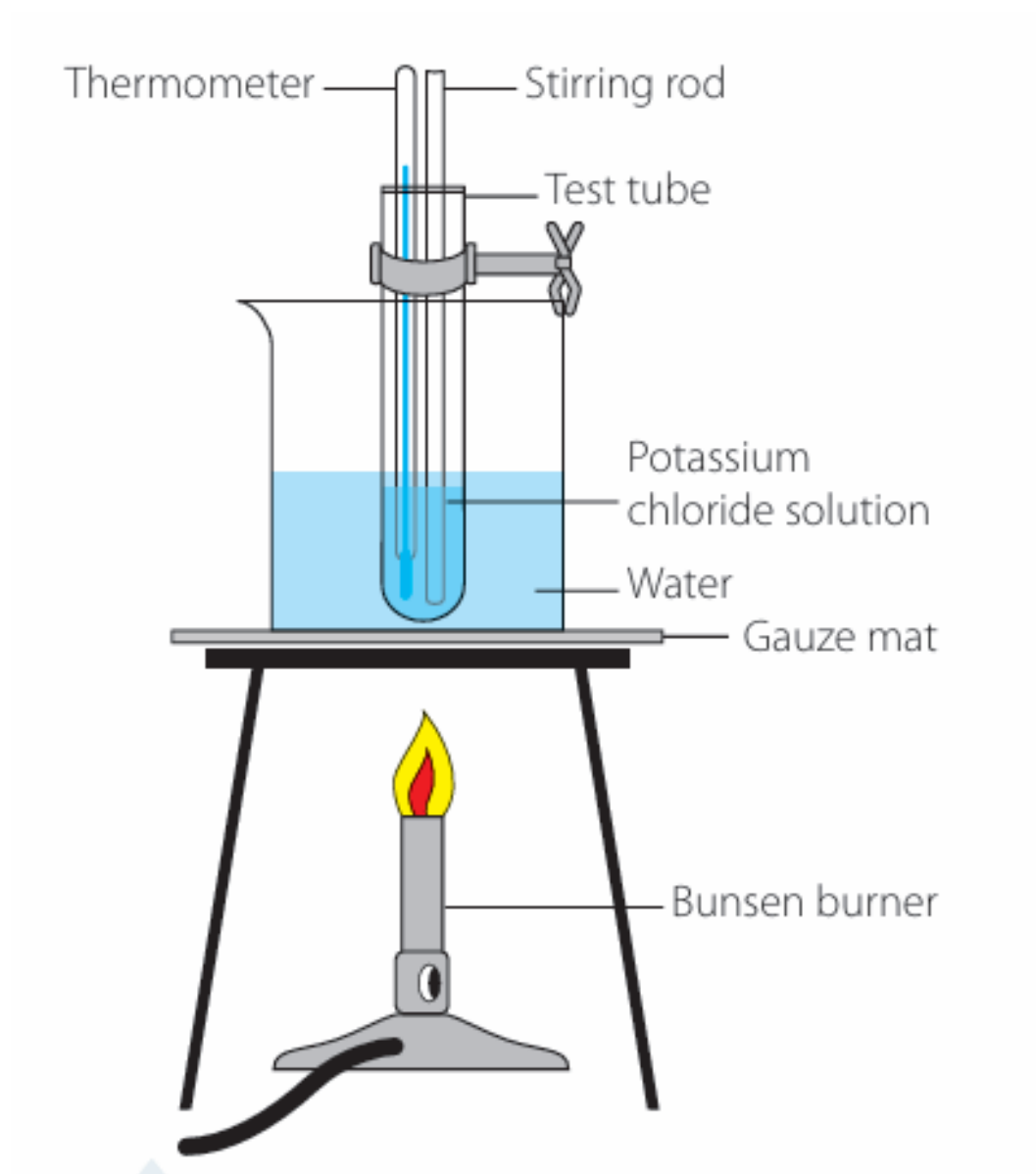


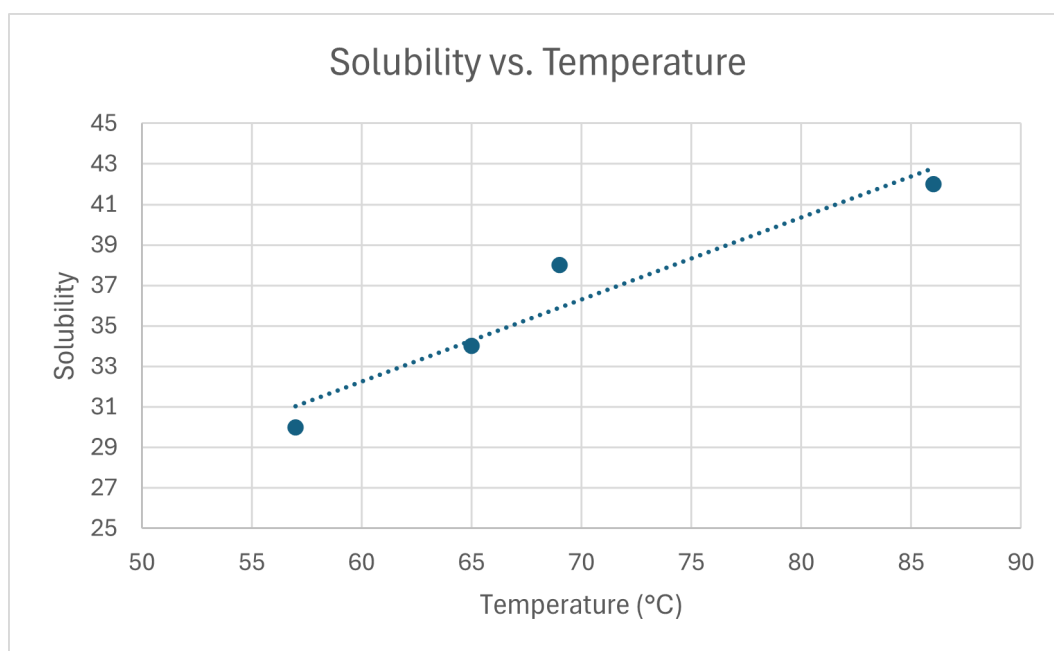
Figure 1.2: Use retort stand instead of beaker clip

1.13.4 Results

Mass of KCl (g)	m_{water} (g)	Temp. recrystallisation (°C)	Solubility (g/100 g H ₂ O)
3.00	10.00	57	30
3.40	10.00	65	34
3.80	10.00	69	38
4.20	10.00	86	42

1.13.5 Analysis of Results

1. Use the class results to plot a graph of solubility against temperature. Plot the temperature along the horizontal axis (0-100°C).



2. From the graph, predict the solubility of potassium chloride at 20 °C, 40 °C, 60 °C and 80 °C.

By technology, trendline forms equation $\text{Solubility} = 0.4056T + 7.9142$

$$\begin{aligned}\text{Solubility at } 20^{\circ}\text{C} &= 0.4056(20) + 7.9142 \\ &= 16.0262\text{g}/100\text{ g H}_2\text{O}\end{aligned}$$

$$\begin{aligned}\text{Solubility at } 40^{\circ}\text{C} &= 0.4056(40) + 7.9142 \\ &= 24.1382\end{aligned}$$

$$\begin{aligned}\text{Solubility at } 60^{\circ}\text{C} &= 0.4056(60) + 7.9142 \\ &= 32.2502\end{aligned}$$

$$\begin{aligned}\text{Solubility at } 80^{\circ}\text{C} &= 0.4056(80) + 7.9142 \\ &= 40.3622\end{aligned}$$

1.13.6 Discussion

1. **Describe what happens to the solubility of potassium chloride as temperature increases**

As temperature increases, solubility of potassium chloride increases, ie. directly proportional relationship

2. **If the theoretical value for the solubility of potassium chloride at 50 °C is 50 g/100 g H₂O, what percentage error does your experiment have?**

$$\text{Percentage error} = \frac{\text{Experimental value} - \text{true value}}{\text{True value}} \times 100\%$$

$$\begin{aligned}\text{Solubility at } 50^{\circ}\text{C} &= 0.4056(50) + 7.9142 \\ &= 28.1942\end{aligned}$$

$$\begin{aligned}\text{Percentage error} &= \frac{28.1842 - 50}{50} \times 100 \\ &= \frac{-21.8158}{50} \times 100 \\ &= -43.6316\%\end{aligned}$$

∴ The experiment had a error value of 43.6316 %

3. **List possible sources of errors in your experiment.**

- Observing the formation of precipitate was subjective as small particles initially formed
- The water bath helped to regulate the temperature of the test tube, however may not have been fully effective

Chapter 2

Module 6 Acid and Base Reactions

2.1 Practical Investigation 5.1 - Preparing and using natural indicators

Aim: To prepare and test natural indicators on a range of substances to determine their acidity or alkalinity

2.1.1 Materials

- Plant material that acts as an indicator (Eg. red cabbage, blueberries, turmeric, petals from violets, geranium, petunias)
- Approx. 5mL of:
 - 0.1 mol L⁻¹ NaOH
 - 0.1 mol L⁻¹ HCl
 - white vinegar
 - household ammonia
 - lemon juice
 - lemonade
 - bicarbonate of soda
 - washing powder
 - antacid tablet
 - salt water
- Distilled water
- 500 mL beaker
- 100 mL beakers
- Test tubes
- Test-tube rack
- 10 mL measuring cylinder
- Knife
- Cutting board

- Mortar and pestle
- Kettle (for warm water)
- Hotplate
- Spatula
- Droppers
- Stirring rod
- Strainer or filter paper and funnel
- Safety glasses

2.1.2 Risk Assessment

Hazard	Precaution
Acids are corrosive, irritate eyes	Handle with caution
Bases are caustic, irritate eyes	Handle with caution
Ammonia is caustic	Use in well ventilated areas

2.1.3 Method

1. For the red cabbage: Finely shred two leaves of cabbage, place in 500 mL beaker and just cover with distilled water (about 200 mL). Slowly boil the cabbage leaves until the water turns a dark reddish-purple and the leaves lose most of their colour.
2. Allow to cool and pour the liquid off into a clean 100 mL beaker. This is the red cabbage indicator. Note: If the colour of the solution is pale, further boiling may be necessary to concentrate the solution.
3. For other plant material: Cut the material into small pieces and place in a mortar and pestle. Grind the material to a paste, add 5-10 mL of warm water and stir.
4. Strain the solution into a beaker to remove any solids.
5. Place 2 mL of each of NaOH and HCl into clean separate test tubes. Add a few drops of one indicator to each test tube until a definite colour is observed. Record the indicator and its colour in your results table.
6. Repeat step 5 with other indicators and record your results in the table.
7. Repeat steps 5 and 6 with other substances. Classify the substances as acidic, basic or neutral.
8. Place 2 mL of HCl in a clean test tube. Choose an indicator that produced a good colour difference between acid and base and add a few drops to the test tube.
9. Add NaOH a few drops at a time to the HCl test tube until the colour no longer changes. Record any colour changes that occur during the addition of NaOH.
10. To the test tube from step 9 add HCl a few drops at a time until the colour no longer changes. Record any colour changes.

2.1.4 Results

2.1.5 Discussion

- 1. Identify which indicators would be most effective in identifying acidic, basic and neutral solutions. Provide a reason for your choice.** Acid Neutral Basic
- 2. Which indicators, if any, were not effective in distinguishing between acidic, basic and neutral solutions? Suggest possible reasons for this.**

The beetroot and blue tea indicators were not as effective compared to the universal indicator. For unknown solutions, universal indicator allows identification of the pH. Blue tea has a lower concentration of anthocyanin compared to the beetroot solution, therefore was less effective as an indicator.

- 3. Using your results, justify whether or not indicator colour change is a reversible reaction.**

It is a reversible reaction

2.1.6 Conclusion

- 1. Explain why indicators give a range of colours in different acid and alkaline solutions.**