# Chapter 1: Important Notes for O-Level Chemistry Practical

## 1.1 Qualitative Analysis

### 1.1.1 Cheat Sheet given during Qualitative Analysis Practical

**Test for anions**

| anion | test | test result |
|---|---|---|
| carbonate ($CO_3^{2-}$) | add dilute acid | effervescence, carbon dioxide produced |
| chloride ($Cl^-$) [in solution] | acidify with dilute nitric acid, then add aqueous silver nitrate | white ppt. |
| iodide ($I^-$) [in solution] | acidify with dilute nitric acid, then add aqueous silver nitrate | yellow ppt. |
| nitrate ($NO_3^-$) [in solution] | add aqueous sodium hydroxide, then aluminium foil; warm carefully | ammonia produced |
| sulfate ($SO_4^{2-}$) [in solution] | acidify with dilute nitric acid, then add aqueous barium nitrate | white ppt. |

**Test for aqueous cations**

| cation | effect of aqueous sodium hydroxide | effect of aqueous ammonia |
|---|---|---|
| aluminium ($Al^{3+}$) | white ppt., soluble in excess giving a colourless solution | white ppt., insoluble in excess |
| ammonium ($NH_4^+$) | ammonia produced on warming | |
| calcium ($Ca^{2+}$) | white ppt., insoluble in excess | |
| copper(II) ($Cu^{2+}$) | light blue ppt., insoluble in excess | light blue ppt., soluble in excess giving a dark blue solution |
| iron(II) ($Fe^{2+}$) | green ppt., insoluble in excess | green ppt., insoluble in excess |
| iron(III) ($Fe^{3+}$) | red-brown ppt., insoluble in excess | red-brown ppt., insoluble in excess |
| lead(II) ($Pb^{2+}$) | white ppt., soluble in excess giving a colourless solution | white ppt., insoluble in excess |
| zinc ($Zn^{2+}$) | white ppt., soluble in excess giving a colourless solution | white ppt., soluble in excess giving a colourless solution |

[Lead(II) ions can be distinguished from aluminium ions by the insolubility of lead(II) chloride.]

**Test for gases**

| gas | test and test result |
|---|---|
| ammonia ($NH_3$) | turns damp red litmus paper blue |
| carbon dioxide ($CO_2$) | gives white ppt. with limewater<br>(ppt. dissolves with excess $CO_2$) |
| chlorine ($Cl_2$) | bleaches damp litmus paper |
| hydrogen ($H_2$) | 'pops' with a lighted splint |
| oxygen ($O_2$) | relights a glowing splint |
| sulfur dioxide ($SO_2$) | turns aqueous acidified potassium manganate(VII) from purple to colourless |

## 1.1.2 TIPS for recording QA Observations

1. For standard cation, anion and gas tests, refer to the **Notes for Qualitative Analysis** given during exam and copy the observations from there.

2. When recording observations, always record the **colour** and **state** of what you see. Be careful with the use of the following words.

| | descriptive words for state | When to use? |
|---|---|---|
| 1 | solid | Use only to describe solid with no solution present. |
| 2 | precipitate | Small solid particles are present in a solution |
| 3 | liquid | Seldom use. Only use when you heat a solid until it melts to a liquid. |
| 4 | solution | A salt/substance dissolved in water |

| | Sample description | Comments |
|---|---|---|
| (√) | white precipitate | use to describe white solid that does not dissolve in the solution.<br>(X) Do not use white solution or (X) cloudy solution. |
| (√) | colourless solution | For solutions that look like plain water.<br>Do not use clear solution. (X) Clear is not a colour. All solutions are clear. |

3. Whenever **bubbles** are seen, a **gas** is formed and you will need to **test for the gas** even if the question did not ask you to test for the gas.

4. Good theory knowledge will help you to predict the gas given out.

| | **reagents** | **possible gases** | **explanation** |
|---|---|---|---|
| **(a)** | dilute acid<br>e.g HC*l*, $H_2SO_4$, $HNO_3$ | carbon dioxide | acid + **metal carbonate** → salt + carbon dioxide + water |
| | | hydrogen | acid + **metal** → salt + hydrogen |
| **(b)** | sodium hydroxide + aluminium + warm gently | ammonia | test for **nitrate** ( Redox reaction) |
| **(c)** | sodium hydroxide + warm gently | ammonia | test for **ammonium** ion<br>alkali + ammonium salt → salt + ammonia + water |
| **(d)** | metal<br>e.g. magnesium | hydrogen | metal + acid → salt + hydrogen<br>reactive metal + water → metal hydroxide + hydrogen [PSCM] |
| **(e)** | hydrogen peroxide + catalyst (e.g. $MnO_2$) | oxygen | $2H_2O_2 \rightarrow 2H_2O + O_2$ |

5. **Thermal Decomposition (when heating solid)**

- When a solid is heated, it may or may not decompose.
- If it decomposes, a **gas** may be evolved at the same time and a **residue** (solid product formed during heating) may be left behind.
- The gas evolved can be tested and identified and this can then help us to identify the substance that was heated.

### (a) HEATING OF METAL CARBONATES

metal carbonate → metal oxide + carbon dioxide gas

Most metal carbonates except potassium carbonate and sodium carbonate when heated give out carbon dioxide and the residue is the corresponding metal oxide

e.g. $\underset{\text{(green solid)}}{CuCO_3} \rightarrow \underset{\text{(black solid)}}{CuO} + CO_2$

$\underset{\text{(white solid)}}{ZnCO_3} \rightarrow \underset{\text{(yellow solid when hot, white solid when cold)}}{ZnO} + CO_2$

### (b) HEATING OF METAL HYDROGEN CARBONATE

metal hydrogen carbonate → metal carbonate + carbon dioxide gas + water

The water vapour formed condenses at the side of the test tube and can be observed as a colourless liquid.

e.g. $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

### (c) HEATING OF METAL NITRATES

metal nitrate → metal nitrite + oxygen gas

metal nitrate → metal oxide + oxygen gas + nitrogen dioxide gas

Some metal nitrates e.g. sodium nitrate, potassium nitrate give out oxygen gas and leave behind the corresponding metal nitrite as residue.

e.g. $2NaNO_3 \rightarrow 2NaNO_2 + O_2$

Most other metal nitrates when heated give out two gases i.e. oxygen and the brown gas nitrogen dioxide and leave behind the corresponding metal oxide as a residue.

e.g. $Zn(NO_3)_2 \rightarrow 2ZnO + 4NO_2 + O_2$

## (d) HEATING OF OXIDES

metal oxide → metal + oxygen

The more reactive metal oxide e.g. aluminium oxide will require a very high temperature (thousands of degree Celsius which is not achievable in the lab) to decompose.

For the temperature of the Bunsen flame, only metal oxides with metals lower in the reactivity series of metals can undergo thermal decomposition.

e.g. $2HgO \rightarrow 2Hg + O_2$
$2Ag_2O \rightarrow 4Ag + O_2$
$2PbO_2 \rightarrow 2PbO + O_2$
$2Pb_3O_4 \rightarrow 6PbO + O_2$

## (e) HEATING OF AMMONIUM COMPOUND

ammonium compound → ammonia gas + other gas

Most ammonium compounds (except ammonium nitrate) when heated give out ammonia gas and the other gases.

e.g. $NH_4Cl \rightarrow NH_3 + HCl$
(Reaction is reversible. The ammonia gas and hydrogen chloride gas will recombine and cool as a white solid at the side of the test tube.)

e.g. $(NH_4)_2CO_3 \rightarrow 2NH_3 + H_2O + CO_2$

## (f) Sample recording when heating solids.

| | Observation |
|---|---|
| Description to changes in solid on heating then cooling | The **(colour)** solid changes to a **(colour)** solid;<br>The **(colour)** solid melts to a **(colour)** liquid;<br>The **yellow** solid changes to a **white** solid when cooled. |
| Description of gas given out | A gas which forms **a white precipitate with limewater** was given out;<br>A gas which relights a glowing splinter was given out.<br>A colourless liquid was found at the side of the test tube;<br>A gas which relights a glowing splinter was given out.<br>A brown gas was also given out;<br>A gas which turns damp red litmus paper blue was given out. |

## 6. Displacement of Metals

e.g. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

**Sample recording:**
A reddish-brown solid coats on the metal.
The blue solution turns to a colourless solution.

e.g. $Mg + ZnSO_4 \rightarrow MgSO_4 + Zn$

**Sample recording:**
A grey solid coats on the metal.
The colourless solution remains colourless.
Bubbles are observed.
A gas which extinguishes a lighted splint with a pop sound was formed.

This observation is because Mg reacts with water to form hydrogen gas.

## 7. Test for Oxidising Agent (O.A)

A reducing agent that shows a colour change or an observable change is usually used to test for the presence of an oxidising agent

| R.A | concentrated hydrochloric acid | acidified potassium iodide solution | acidified iron (II) sulfate solution |
|---|---|---|---|
| in the presence of O.A | Green chlorine gas evolved which turns damp blue litmus paper red and then bleaches it.<br>(Reason $Cl^-$ oxidised to $Cl_2$) | Colourless potassium iodide turns brown<br>(Reason $I^-$ oxidised to $I_2$) | Pale green iron (II) sulfate solution turns yellow<br>(Reason: $Fe^{2+}$ oxidised to $Fe^{3+}$) |

## 8. Test for Reducing Agent (R.A)

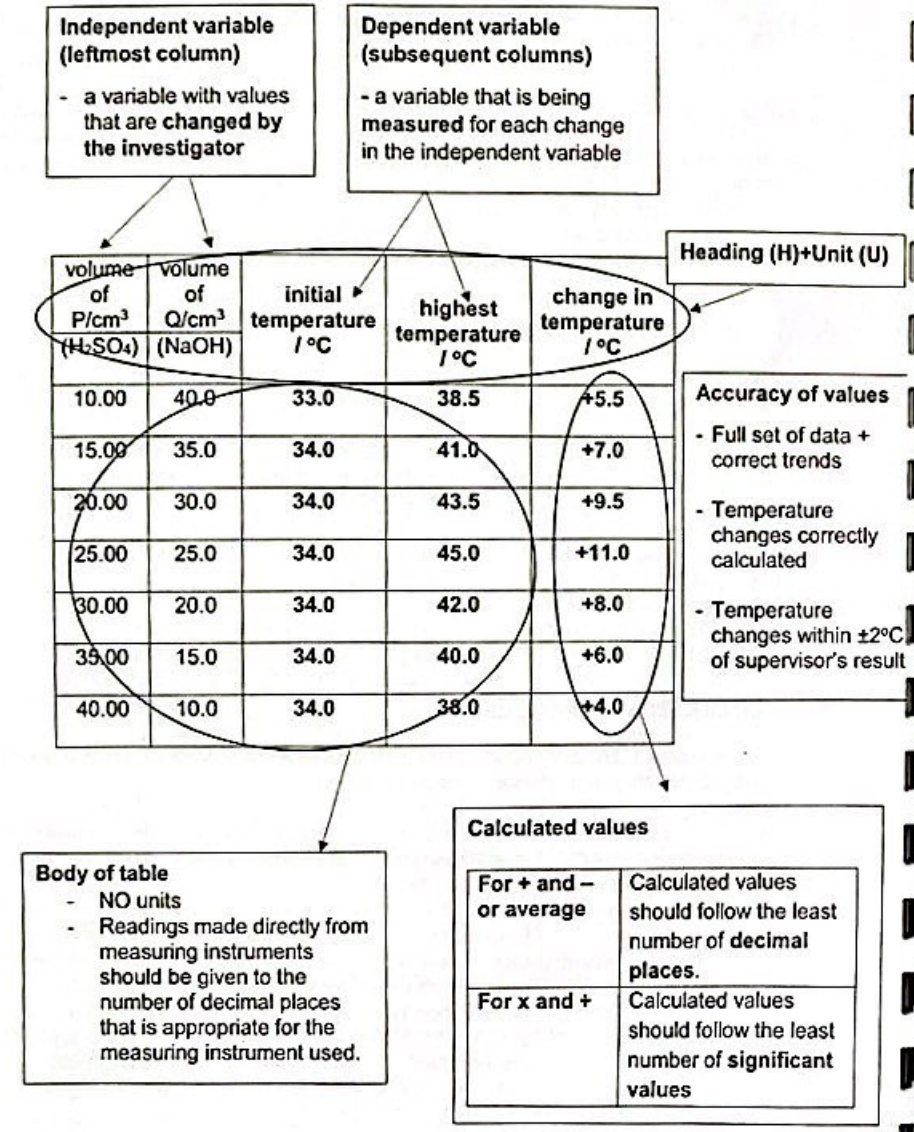
An oxidising agent that shows a colour change or an observable change is usually used to test for the presence of a reducing agent

| O.A | acidified potassium manganate(VII) solution | *acidified potassium dichromate(VI) solution | fresh iron(III) chloride solution | iodine solution |
|---|---|---|---|---|
| in the presence of R.A | Potassium manganate(VII) solution turns from purple to colourless.<br>(reason: $Mn^{7+}$ reduced to $Mn^{2+}$) | Potassium dichromate(VI) solution turns from orange to green.<br>(reason: $Cr^{6+}$ reduced to $Cr^{3+}$) | Yellow iron (III) chloride solution turns green (reason: $Fe^{3+}$ ions being reduced to $Fe^{2+}$ ions) | Iodine solution turns from brown to colourless (reason: iodine being reduced to iodide ions) |

* Not in practical syllabus

## 1.2 Table

### 1.2.1 Important points to note for TABLES



**Independent variable (leftmost column)**

- a variable with values that are changed by the investigator

**Dependent variable (subsequent columns)**

- a variable that is being measured for each change in the independent variable

**Heading (H)+Unit (U)**

| volume of P/cm³ (H₂SO₄) | volume of Q/cm³ (NaOH) | initial temperature / °C | highest temperature / °C | change in temperature / °C |
|---|---|---|---|---|
| 10.00 | 40.0 | 33.0 | 38.5 | +5.5 |
| 15.00 | 35.0 | 34.0 | 41.0 | +7.0 |
| 20.00 | 30.0 | 34.0 | 43.5 | +9.5 |
| 25.00 | 25.0 | 34.0 | 45.0 | +11.0 |
| 30.00 | 20.0 | 34.0 | 42.0 | +8.0 |
| 35.00 | 15.0 | 34.0 | 40.0 | +6.0 |
| 40.00 | 10.0 | 34.0 | 38.0 | +4.0 |

**Accuracy of values**

- Full set of data + correct trends
- Temperature changes correctly calculated
- Temperature changes within ±2°C of supervisor's result

**Body of table**

- NO units
- Readings made directly from measuring instruments should be given to the number of decimal places that is appropriate for the measuring instrument used.

**Calculated values**

| | |
|---|---|
| **For + and – or average** | Calculated values should follow the least number of **decimal places.** |
| **For x and +** | Calculated values should follow the least number of **significant** values |

### 1.2.2 Degree of Accuracy for instruments commonly used in chemistry

| No | Apparatus | Smallest division of instrument | Values can be recorded to the nearest …. | Examples of recording |
|---|---|---|---|---|
| 1 | Ruler | 0.1 cm | 0.1 cm | 11.0 cm<br>11.1 cm |
| 2 | Measuring cylinder (100 cm³) | 1 cm³ | 0.5 cm³ | 14.0 cm³<br>14.5 cm³ |
| 3 | Burette | 0.1 cm³ | 0.05 cm³ | 10.00 cm³<br>10.05 cm³ |
| 4 | Electronic balance | 0.1 g | 0.1 g | 150.0 g<br>151.1 g |
| | | 0.01 g | 0.01 g | 131.00 g<br>131.10 g |
| 5 | Pipette (fixed volume) | | | 25.0 cm³ |
| 6 | Stopwatch (digital) | 0.01 s | 1s<br>( for chemistry experiments involving speed of reaction) | 20 s<br>28 s |
| 7 | Thermometer (-10 °C to 110 °C) | 1 °C | 0.5 °C | 22.0 °C<br>22.5 °C |

## 1.3 Graph

### 1.3.1 Important points to note for graphs

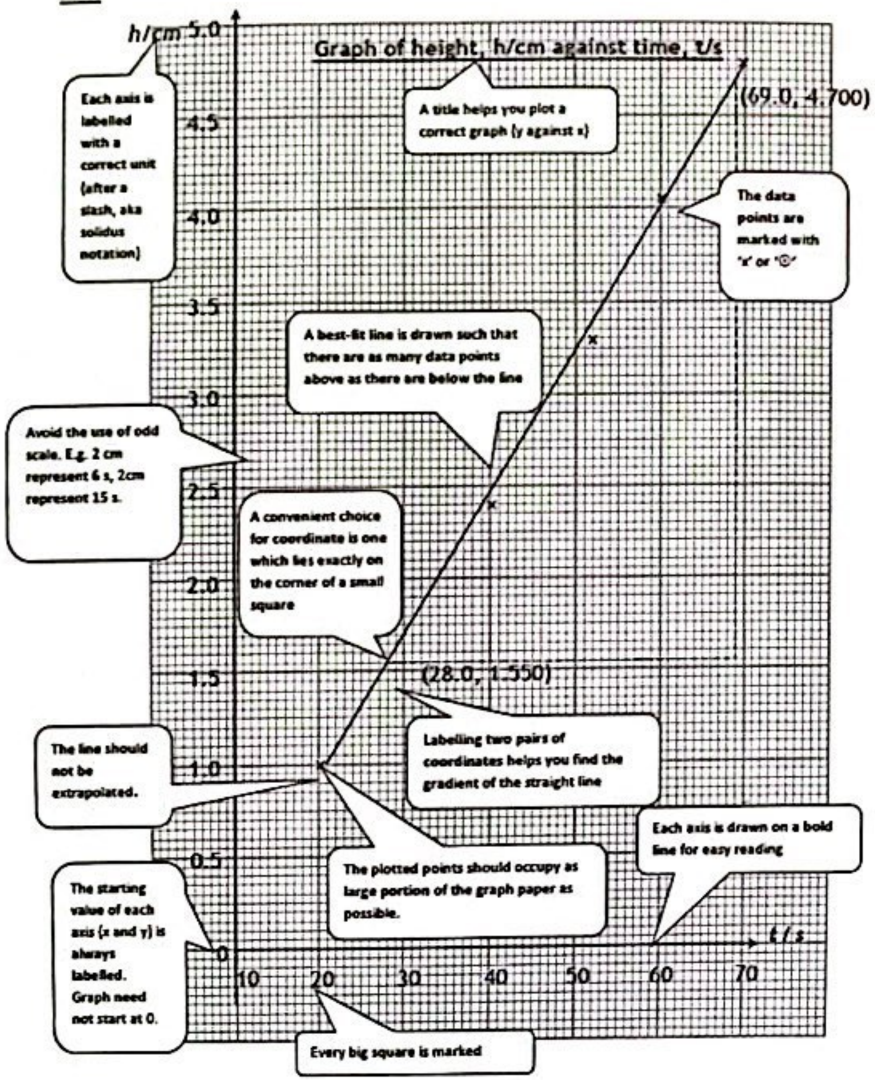
**(a) Clear representation of data:**

1. Graphs should be drawn with a sharp pencil (mechanical pencil) and transparent long ruler
2. Axes should be labelled clearly with quantity and unit
3. Points should be accurately indicated by a small cross.

**(b) Using the graph as an instrument for analysis:**

| Marking Points | Description |
|---|---|
| Axis | Scales for the axes should maximise the use of the graph paper in both the vertical and horizontal directions.<br>This allows as large a graph to be plotted as possible. By doing so, the space between the plotted lines are larger, reducing judgement error involved when the points are joined with a line to describe their relationship. |
| Scale | Use sensible ratio for scale, e.g. 2 cm on graph paper representing 1, 2 or 5 units.<br>Do not use odd scale e,g, 2 cm representing 3 units as this will mean that each division of the graph is a recurring decimal which will make reading of the data points difficult |
| Points | Each point should be plotted to an accuracy within half of the smallest square on the grid.<br>All data values, including the coordinates on the line drawn and the intercept should be read to an accuracy within one half of the smallest square on the grid, and recorded to the number of decimal places corresponding to half of the smallest square on the grid. |
| Line of best fit | 1. A line of best fit, which may be a curve, is drawn when there is a possible relationship between the variables.<br>2. It should have equal number of points on either side of the line over its entire length.<br>3. This line drawn may or may not pass through any of the plotted points.<br>4. The line drawn should be smooth and even to facilitate the accurate reading of the coordinates along the line.<br>5. Use a transparent long ruler to draw the line of best fit so that you can see the other points through the ruler.<br>6. **Extrapolation of best fit line:**<br>(a) In general, if there is no known relationship between the variables, the line of best fit should not be extrapolated as the trend line drawn can only describe the relationship within the range of data collected.<br>(b) If there is a known relationship between the variables, extrapolate the graph to predict values beyond the range of the experiment. |

(c) In summary, the figure below shows an example of how a good graph looks like:

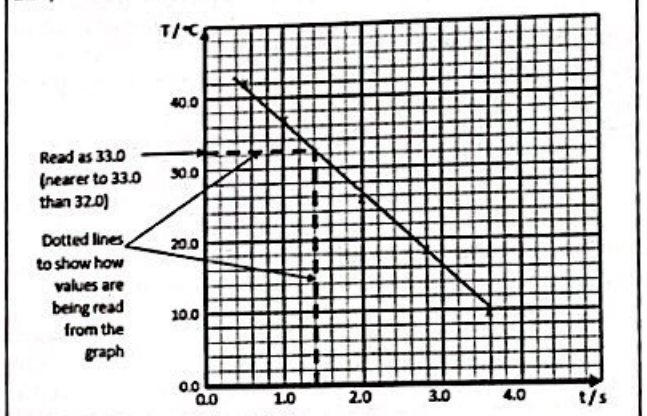


## 1.3.2 Analysis of Graph

(a) **Reading values from a graph**

1. When reading off values from a graph directly, draw dotted perpendicular lines to connect the data point to both the horizontal and vertical axes, for ease of reading.
2. Data values should be read to an accuracy within one half of one of the smallest squares on the grid. See the example below:

Example 5.4 Find T at t = 1.4 s

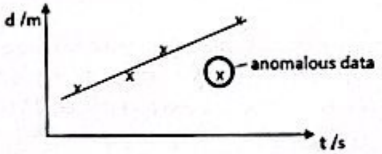


Therefore, T = 33.0°C when t = 1.4s.

(b) **Dealing with anomalous results**

1. When a graph is plotted and there are points that do not follow the general trend or pattern of the graph, these odd data points are known as anomalous results. If time permits, the experiment should be repeated to verify the reliability/validity if the anomalous points.
2. It is important that students indicate the anomalous points on the graph clearly.
3. Failure to do so will result in an erroneous line of best fit. Anomalous points are usually circled and labelled on the graph directly

Example 5.8 Indication of anomalous data on a graph



### 1.3.3 Drawing Conclusions

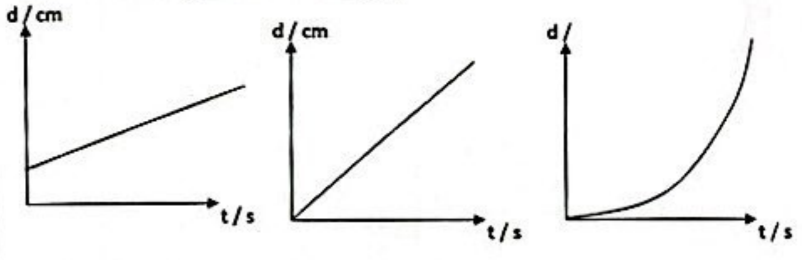
1. Making meaning from the patterns and trends shown in graphs is the most common way to draw conclusion in practical work. The relationship between the independent and the dependent variables can be derived from graphs.

2. If a graph shows a **straight line,** the conclusion may be *"the dependent variable varies linearly with the independent variable".*

3. If the **straight line passes through the origin (0, 0),** it may be further concluded that *"the dependent variable is directly proportional to the independent variable".*

4. If the graph shows a **curve with a general positive gradient,** a statement like *"as the independent variable increases, the dependent variable increases non-linearly"* will suffice as a conclusion.

**Example 5.10 Drawing conclusion from a graph**



(a) d increases linearly with t (b) d is directly proportional to t (c) As t increases, d increases non-linearly

### 1.4 Evaluation and Sources of Error

1. Evaluation may be defined as "judging the validity of conclusion by considering the **quality of the evidence** and the **suitability of the procedure**" (Haywood, 2003).

2. Examples of evaluation question includes:
   - identifying **sources of errors** and **suggesting how the errors may be overcome or reduced,**
   - commenting on how well the evidence supports the conclusion;
   - identifying and explaining the presence of anomalous results;
   - considering extraneous variables that are not reflected in the plan and suggesting how these may be controlled;
   - comparing the results collected with theoretical calculation or alternative source of date (such as a graph from the textbook) to assess the accuracy of the collected evidence;
   - an assessment of the adequacy of the range of data obtained.

3. There are two common types of errors that affect results:

   (a) random errors

   **Random errors** arise when an observer estimates the reading on an instrument. They are **inevitable** but can be **reduced by taking a large number of readings and averaging them.**

   (b) systematic errors.

   **Systematic errors** cause readings taken by an observer to consistently deviate from the true values. This could be attributed to existing zero errors in the instruments or caused by environmental factors such as hot weather on a particular day. Averaging does not reduce systematic errors. **They can be corrected if the observer can identify the causes of errors.**

4. Some **guiding questions** that may help you to **identify experimental errors.**

**(a) Which experimental procedures were difficult to carry out?**

**(b) Which measurements were difficult to observe and record accurately?**
e.g to record the time taken for a complete reaction by observing the exact moment where a colour change occur, the time measured may be inaccurate as there is a time lapse between seeing the colour change, and stopping the stopwatch.

**(c) Other than the measured variables, what other factors could have affected the results?**
It is sometimes impossible to ensure that all control variables are kept constant in an experiment. For example, in an experiment that investigates the rate of heat loss of a liquid over a period of time, the loss of heat to the environment can by minimize by using

a Styrofoam cup but it cannot be eliminated completely. The room temperature may vary during the experiment due to changing weather condition. The presence of wind may also affect the rate of cooling. All these factors could all potentially affect the rate of cooling.

5. After identifying the errors of the experiment, students could then be asked to suggest ways to improve the reliability/validity of the experimental results.

6. The **improvements** suggested must be able to address the sources of errors identified. They can include
   - **modification** to experimental procedures,
   - more rigorous **control of variables**,
   - the use of **more accurate measuring instruments** and/or **procedures**.

Examples of sources of errors and how they affect the results collected.

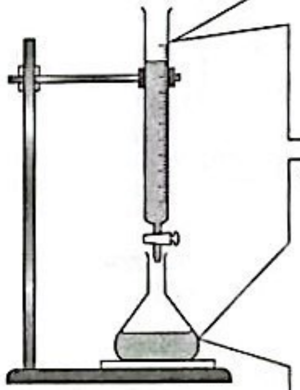
| **Sources of error** | **How they affect the results** |
|---|---|
| 1. Heat loss to surrounding | temperature measured is lower than expected |
| 2. When evaporating to dryness, some of the salt formed may be lost due to "sputtering". | mass of salt obtained lower than expected |
| 3. The salt formed by crystallisation may not be dry | mass of salt measured is higher than the actual mass of the salt |
| 4. Difficulty in comparing colour of solution with universal indicator scale | inaccurate pH measurements |
| 5. Slow run of solution from burette, more heat loss to surrounding before taking temperature readings | temperatures measured smaller than expected |
| 6. The large piece of magnesium ribbon does not allow for fast reaction with acid; more heat loss to surrounding due to slow reaction | temperature rise measured is lower than expected |
| 7. No rinsing of measuring cylinder/ thermometer between change of solutions (Accepts only if rinsing was not observed) | incorrect temperature measurements |
| 8. human reaction time (e.g. in the starting and stopping of stopwatch) | inaccurate time measurements |
| 9. Inconsistent swirling of solution mixture | inaccurate time measurements |
| 10. Mixture of substance A and substance B used for preparing solution P may not be homogeneous | inaccurate results |

## 1.5 Volumetric Analysis (Titration)

### 1.5.1 What is titration use for?

1. Titration, also known as volumetric analysis is use to determine the concentration of an unknown solution.

2. In titration, we slowly add a solution of known/standard concentration (called the **titrant**) from a burette into a solution of known volume and unknown concentration (called the **analyte**) in a conical flask until the end-point is reached.



The **titrant** is added from the burette

- The concentration of **titrant** is known.
- The volume of **titrant** needed for reaction to complete is noted when the colour of indicator changes.

The **analyte** is contained in the conical flask

- 25.0 $cm^3$ of the **analyte** is accurately measured and placed in the conical flask.
- The concentration of the **analyte** is usually unknown and can be determined using simple mole calculations.

$$xA + yB \rightarrow C + D$$

$$\frac{M_A V_A}{M_B V_B} = \frac{x}{y}$$

$M_A$ = concentration of solution A in mol/$dm^3$
$M_B$ = concentration of solution B in mol/$dm^3$
$V_A$ = volume of solution A in $cm^3$
$V_B$ = volume of solution B in $cm^3$

3. There are 2 main types of titration.
   - Acid – Base titration
   - Redox titration

| | |
|---|---|
| **Step 3:** Perform the titration | 1. During titration, the titrant needs to be added to the analyte slowly especially nearing the end-point.<br>2. The contents in the conical flask need to be swirled continuously.<br>3. You may use deionised water to wash down the analyte from the side of the conical flask.<br>4. When nearing the end point, add the titrant dropwise. The titration is complete when one drop of the titrant causes the indicator to change colour permanently.<br>5. Titrations should be repeated to obtain titres that are very close to one another. The average of the two most consistent results will then be calculated to be used for further calculations. |

### 1.5.3 Recoding titration values in a table

1. Sample titration table and values.

| Titration | 1 | 2 | 3 |
|---|---|---|---|
| Final burette **reading** /cm$^3$ | 19.50 | 19.40 | 19.40 |
| Initial burette **reading** /cm$^3$ | 0.00 | 0.00 | 0.00 |
| **Volume of solution used/ cm$^3$** | 19.50 | 19.40 | 19.40 |
| Best titration results | | ✓ | ✓ |

How marks are awarded for titration results (5m)?

| | |
|---|---|
| **Table** | records initial burette readings, final burette readings and volume added with correct headings and units in a titration table [1] |
| **Decimal place** | all burette readings for all accurate titres in titration table are recorded to nearest 0.05cm$^3$ [1] |
| **Accuracy** | • for the average titre (of consistent readings) within **0.20 cm$^3$** of Supervisor's average value score 2 marks [2]<br>• for the average titre (of consistent readings) within **0.30 cm$^3$** of Supervisor's average value score 1 mark. [1] |
| **Concordance** | Choose the two best results which should be within 0.20 cm$^3$ and tick (✓) them in the table. |

### 1.5.4 Calculating average titre

Use the two best results to obtain the average volume of titrant used. This volume will be used in stoichiometric calculations in other parts of the question

e.g. Average volume of solution Q = $\frac{19.40 + 19.40}{2}$

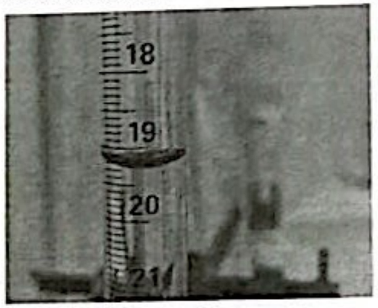
$= 19.40\ cm^3$

Note: Working must be shown.

### 1.5.5 Possible errors in accuracy

1. **Wrong reading of burette**

**e.g. Is this 19.20 cm$^3$ or 20.80 cm$^3$**



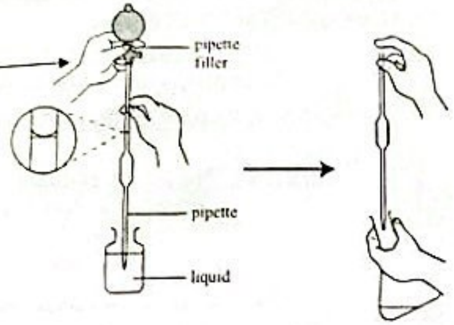
2. **Inaccurate readings along the way**

- Ensure no parallax error,
- no bubble in pipette and burette,
- pipette and burette are vertical
- You may want to use a white tile as backing to help you read more accurately.

3. **Rinsing Error**

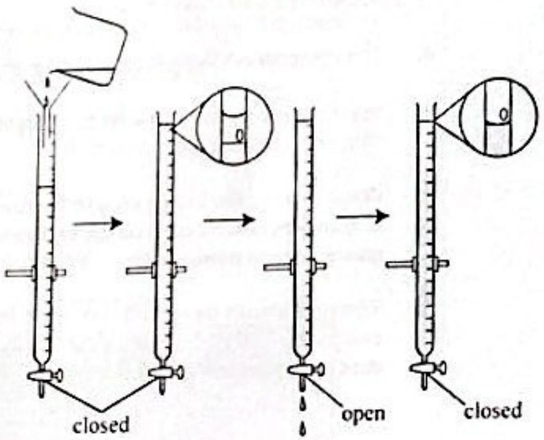
**Burette and Pipette– Must rinse with water + the solution**
**Conical flask – Only rinse with water. DO NOT rinse with the solution.**

### 1.5.2 How to perform titration?

| | |
|---|---|
| **Step 1:** Filling the conical flask. | 1. Wash the pipette with **water + deionised water + analyte.** |
| | 2. Wash the conical flask **water + deionised water** ONLY |
| | 3. Use a pipette filler and pipette, draw out 25.0 cm³ of the analyte and transfer it to the conical flask<br><br> <br>How to use the pipette to draw out 25.0 cm³ of analyte?<br>**(a)** Insert the pipette gently into the pipette filler.<br>**(b)** Press "A" to evacuate the air.<br>**(c)** Press "S" to draw the solution up the pipette.<br>**(d)** Press "E" to evacuate the solution from the pipette.<br>**(e)** The correct volume of solution is drawn when the bottom of the meniscus is in line with the brown ring of the pipette. Avoid parallax error when taking readings.<br>*Note:*<br>**(a)** There will always be a small amount of solution left at the tip of the pipette. Just tap lightly 2 times and it will suffice. You need not tap so hard to try to expel all the solution.<br>**(b)** Record the volume of your analyte that you had drawn using a pipette correct to **1 decimal place.** (e.g. 25.0 cm³)<br>**(c)** Make sure that there is **no air bubble** in the pipette as the air bubble will significantly affect the accuracy of your results. |

**Safe use of pipette:**

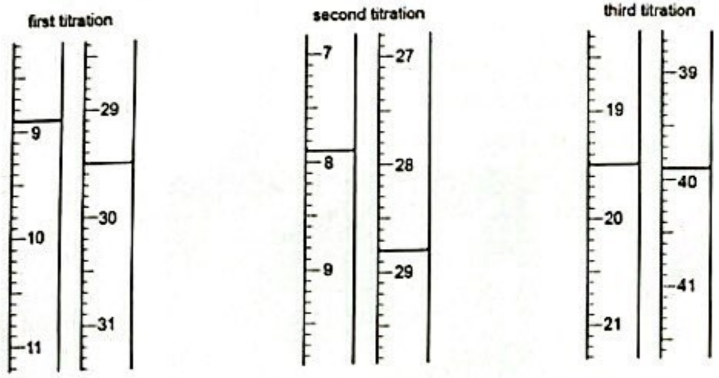
Make sure that your hand is holding near the top of the pipette (above the bulb of the pipette) when inserting the pipette filler.

| | |
|---|---|
| **Step 2:** Filling up the burette | <br> <br>Instructions:<br>**1.** Wash the burette with **tap water + deionised water + titrant.**<br>**2.** Rinse a filter funnel with **distilled water + titrant**<br>**3.** Clamp the burette to a retort stand with the tap close. Ensure that the burette is of a suitable height and it is vertical to the ground.<br>**4.** Using the filter funnel, pour the titrant into the burette until the level is above the 0.00 cm³ mark. If the solution level is above your eye level, place the retort stand and burette on the bench so as to read more comfortably. **Remove the filter funnel from the burette when not in use.**<br>**5.** If you see any bubbles in the burette, open the tap and let some of the solution flow out with the bubble. The air bubble will significantly affect the accuracy of your results.<br>Note:<br>**1.** The starting solution level in the burette need not be 0.00 cm³ because the volume of solution used is the difference between the initial reading and the final reading. However. Starting with 0.00 cm³ may help to eliminate subtraction errors.<br>**2.** Record your burette readings to the nearest 0.05 cm³ (e.g. 22.10 cm³, 18.65 cm³) |

### 1.5.6 Dry Practice

When reading the burette, note that the smaller number is on top while the bigger number is below.

(i) Complete the table below.



| Titration | 1 | 2 | 3 |
|---|---|---|---|
| | | | |
| | | | |
| | | | |
| | | | |

(ii) From the titration results, obtain an average volume of solution to be used in your calculations. Show clearly how you obtained this volume.

### 1.6 Planning experiments

1. **For investigative** experiments.
   - (a) State clearly the **approach** you are going to take to investigate the problem.
   - (b) State the **variable** that you **change** and the **variables that are kept constant**.
   - (c) Write out the **procedures step** by step in chronological order.
   - (d) State what values you are collecting and summarise it in a **table**.
   - (e) Explain how you will **process the data** collected.
   - (f) Make a **conclusion** for the investigation.

2. TIPS when writing Approach

   Consider
   - (a) how you would like to do the experiment?
   - (b) the measurements you will take?
   - (c) how the measurements will help you to investigate the problem?

3. TIPS when writing **procedure**
   - (a) Write the procedures starting with a **verb**. Each step should only have one instruction.
   - (b) State the **measuring instruments** you need to measure certain physical quantities.

     e.g. Use a measuring cylinder to measure 50 $cm^3$ of solution A.
   - (c) State the **apparatus** and **materials** such as conical flask, burette, filter funnel, litmus paper you would use in the procedure.

4. TIPS that will help you in **DATA processing**
   - (a) Use **unknown variables** for the physical quantities measured **if further calculations are needed.**

     e.g. Use an electronic balance to measure the mass of dry residue, ***x***.
   - (b) Write down the calculations in terms of the unknown variables in order to obtain the value of certain quantities.

     e.g. Number of moles of carbon dioxide = v/24 mol.

5. TIPS in writing **Conclusion**

   State how certain information such as gradient from a graph drawn is able to provide a conclusion.