Basic principles of Chemistry

PRACTICAL
Introduction/Rationale

Chemistry is a science. Chemistry practical all over the world is emphasized to all candidates sitting for a Chemistry paper. There are about seven main basic universal emphasis for all chemistry candidates sitting for a chemistry paper;

(i) Titration/volumetric analysis
(ii) Thermochemistry (energy changes)
(iii) Chemical kinetic (rates of reaction)
(iv) Qualitative analysis (organic/inorganic)
(v) Solubility and solubility curves
(vi) Flame test
(vii) Physical / general chemistry

(a) Titration/volumetric analysis

Titration is determining the end point of the burette contents that react with fixed (usually 25.0cm³ from a pipette) conical flask contents. As evidence of a titration actually done examining body require the candidate to record their burette readings before and after the titration. For KCSE candidates burette readings must be recorded in a titration table in the format provided by the Kenya National Examination Council. As evidence of all titration actually done Kenya National Examination Council require the candidate to record their burette readings before and after the titration to complete the titration table in the format provided.

Sample Titration table format

<table>
<thead>
<tr>
<th>Final burette reading (cm³)</th>
<th>24.0</th>
<th>24.0</th>
<th>24.0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial burette reading (cm³)</td>
<td>0.0</td>
<td>0.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of solution used(cm³)</td>
<td>24.0</td>
<td>24.0</td>
<td>24.0</td>
</tr>
</tbody>
</table>

Calculate the average volume of solution used

\[
\frac{24.0 + 24.0 + 24.0}{3} = 24.0 \text{ cm}^3
\]

As evidence of understanding the degree of accuracy of burettes, all readings must be recorded to a decimal point. As evidence of accuracy in carrying out the titration, candidates value should be within 0.2 of the school value.
The school value is the **teachers** readings presented to the examining body/council based on the **concentrations** of the solutions s/he presented to her/his candidates. Bonus mark is awarded for averaged reading within 0.1 school value as Final answer. Calculations involved after the titration require candidates **thorough practice mastery** on the:

(i) relationship among the mole, molar mass, mole ratios, concentration, molarity.
(ii) mathematical application of 1st principles.

Very useful information which candidates forget appear usually in the beginning of the paper as:

“You are provided with…”

All calculation must be to the 4th **decimal point** unless they divide fully to a lesser decimal point.

**Never round off answers.**

**b) Thermochemistry/energy changes**

Energy is the capacity to do work which is measured in Joules (J) or (kJ). Chemical/physical changes take place with **absorption** (Endothermic) or **evolution/production** (Exothermic) of heat.

Practically:

(i) endothermic changes show absorption of heat by a **fall** / **drop** in temperature and has a **+ΔH**

(ii) exothermic changes show evolution/production of heat by a **rise** in temperature and has a **-ΔH**

(iii) temperature is measure using a **thermometer**.

(iv) a school thermometer is either coloured (alcohol) or colourless (mercury)

(v) For accuracy, candidates in the same practical session should use the same type of thermometer.

(vi) fall / drop (+ΔH) in temperature is movement of thermometer **level downward.**

(vii) rise (-ΔH) in temperature is movement of thermometer **level upwards.**

Physical changes changes mainly involve melting/freezing/fusion and boiling/vapourization.

Chemical changes changes mainly involve displacement, dissolving, neutralization

**a). Energy changes in physical processes**

Melting/freezing/fusion/solidification and boiling/vaporization/evaporation are the two physical processes.

Melting / freezing point of pure substances is fixed / constant.

The boiling point of pure substance depends on **external atmospheric pressure.**

Melting/fusion is the physical change of a **solid** to **liquid.** Freezing/fusion is the physical change of a **liquid** to **solid.**

Melting/freezing/fusion/solidification are therefore two **opposite** but **same** reversible physical processes. i.e
Boiling/vaporization/evaporation is the physical change of a liquid to gas/vapour. Condensation/liquidification is the physical change of gas/vapour to liquid. Boiling/vaporization/evaporation and condensation/liquidification are therefore two opposite but same reversible physical processes. i.e

\[ A (s) \overset{\text{equilibrium}}{\rightleftharpoons} A(l) \]

Practically

(i) Melting/liquidification/fusion involves heating a solid to weaken the strong bonds holding the solid particles together.

Solids are made up of very strong bonds holding the particles very close to each other (Kinetic Theory of matter)

On heating these particles gain energy/heat from the surrounding heat source to form a liquid with weaker bonds holding the particles close together but with some degree of freedom.

Melting/fusion is an endothermic (+\(\Delta H\)) process that require/absorb energy from the surrounding.

(ii) Freezing/fusion/solidification involves cooling a liquid to reform/rejoin the very strong bonds to hold the particles very close to each other as solid and thus lose their degree of freedom (Kinetic Theory of matter).

Freezing/fusion/solidification is an exothermic (-\(\Delta H\)) process that require particles holding the liquid together to lose energy to the surrounding.

(iii) Boiling/vaporization/evaporation involves heating a liquid to completely break/free the bonds holding the liquid particles together.

Gaseous particles have high degree of freedom (Kinetic Theory of matter). Boiling/vaporization/evaporation is an endothermic (+\(\Delta H\)) process that require/absorb energy from the surrounding.

(iv) Condensation/liquidification is reverse process of boiling/vaporization/evaporation.

It involves gaseous particles losing energy to the surrounding to form a liquid. It is an exothermic (+\(\Delta H\)) process.

The quantity of energy required to change one mole of a solid to liquid or to form one mole of a liquid from liquid at constant temperature is called molar enthalpy/latent heat of fusion. e.g.

\[ H_2O(s) \rightarrow H_2O(l) \quad \Delta H = +6.0kJ \text{ mole}^{-1} \quad \text{(endothermic process)} \]
\[ H_2O(l) \rightarrow H_2O(s) \quad \Delta H = -6.0kJ \text{ mole}^{-1} \quad \text{(exothermic process)} \]

The quantity of energy required to change one mole of a liquid to gas/vapour or to form one mole of a liquid from gas/vapour at constant temperature is called molar enthalpy/latent heat of vapourization. e.g.

\[ H_2O(l) \rightarrow H_2O(g) \quad \Delta H = +44.0kJ \text{ mole}^{-1} \quad \text{(endothermic process)} \]
\[ H_2O(g) \rightarrow H_2O(l) \quad \Delta H = -44.0kJ \text{ mole}^{-1} \quad \text{(exothermic process)} \]
To determine the boiling point of water

Procedure:
Measure 20cm³ of tap water into a 50cm³ glass beaker. Determine and record its temperature. Heat the water on a strong Bunsen burner flame and record its temperature after every thirty seconds for four minutes.

Sample results

<table>
<thead>
<tr>
<th>Time (seconds)</th>
<th>0</th>
<th>30</th>
<th>60</th>
<th>90</th>
<th>120</th>
<th>150</th>
<th>180</th>
<th>210</th>
<th>240</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature (°C)</td>
<td>25.0</td>
<td>45.0</td>
<td>85.0</td>
<td>95.0</td>
<td>96.0</td>
<td>96.0</td>
<td>96.0</td>
<td>97.0</td>
<td>98.0</td>
</tr>
</tbody>
</table>

Questions
1. Plot a graph of temperature against time (y-axis)
   Sketch graph of temperature against time

Temperature °C

Time in Seconds
2. From the graph show and determine the boiling point of water
Note:
Water boils at 100°C at sea level/one atmosphere pressure/101300Pa but boils at below 100°C at higher altitudes.
The sample results above are from Kiriari Girls High School-Embu County on the slopes of Mt Kenya in Kenya. Water here boils at 96°C.
3. Calculate the molar heat of vaporization of water.(H= 1.0, O= 16.0)
Working:

Mass of water = density x volume => (20 x 1) /1000 = 0.02kg

Quantity of heat produced
= mass of water x specific heat capacity of water x temperature change
=>0.02kg x 4.2 x (96 – 25) = -5.964kJ

Heat of vaporization of one mole H₂O
= Quantity of heat
Molar mass of H₂O

=> -5.964kJ
18 = -0.3313 kJ mole⁻¹

To determine the melting point of candle wax

Procedure

Weigh exactly 5.0 g of candle wax into a boiling tube. Heat it on a strongly Bunsen burner flame until it completely melts.
Insert a thermometer and remove the boiling tube from the flame. Stir continuously. Determine and record the temperature after every 30 seconds for four minutes.

Sample results
<table>
<thead>
<tr>
<th>Time (seconds)</th>
<th>0</th>
<th>30</th>
<th>60</th>
<th>90</th>
<th>120</th>
<th>150</th>
<th>180</th>
<th>210</th>
<th>240</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature</td>
<td>93.0</td>
<td>85.0</td>
<td>78.0</td>
<td>70.0</td>
<td>69.0</td>
<td>69.0</td>
<td>69.0</td>
<td>67.0</td>
<td>65.0</td>
</tr>
</tbody>
</table>

Questions
1. Plot a graph of temperature against time(y-axis)
b) Energy changes in chemical processes
(i) Standard enthalpy/heat of displacement $\Delta H^\circ_d$
(ii) Standard enthalpy/heat of neutralization $\Delta H^\circ_n$
(iii) Standard enthalpy/heat of solution/dissolution $\Delta H^\circ_s$

(i) Standard enthalpy/heat of displacement $\Delta H^\circ_d$

The molar standard enthalpy/heat of displacement may be defined as the energy/heat change when one mole of substance is displaced/removed from its solution at standard conditions.

Some displacement reactions
(i) $\text{Zn(s)} + \text{CuSO}_4(aq) \rightarrow \text{Cu(s)} + \text{ZnSO}_4(aq)$
   Ionically: $\text{Zn(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Cu(s)} + \text{Zn}^{2+}(aq)$

(ii) $\text{Fe(s)} + \text{CuSO}_4(aq) \rightarrow \text{Cu(s)} + \text{FeSO}_4(aq)$
   Ionically: $\text{Fe(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Cu(s)} + \text{Fe}^{2+}(aq)$

(iii) $\text{Pb(s)} + \text{CuSO}_4(aq) \rightarrow \text{Cu(s)} + \text{PbSO}_4(s)$
   This reaction stops after some time as insoluble PbSO$_4$(s) coat/cover unreacted lead.

(iv) $\text{Cl}_2(g) + 2\text{NaBr(aq)} \rightarrow \text{Br}_2(aq) + 2\text{NaCl(aq)}$
   Ionically:
   $\text{Cl}_2(g) + 2\text{Br}^-(aq) \rightarrow \text{Br}_2(aq) + 2\text{Cl}^-(aq)$

To determine the molar standard enthalpy/heat of displacement ($\Delta H^\circ_d$) of copper:

Procedure
Place 20cm³ of 0.2M copper(II)sulphate(VI)solution into a 50cm³ plastic beaker/calorimeter.

Determine and record the temperature of the solution $T_1$.

Put all the Zinc powder provided into the plastic beaker. Stir the mixture using the thermometer.

Determine and record the highest temperature change to the nearest 0.5°C - $T_2$.

Repeat the experiment to complete table 1 below

Sample results Table 1

<table>
<thead>
<tr>
<th>Experiment</th>
<th>I</th>
<th>II</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final temperature of solution ($T_2$)</td>
<td>30.0 °C</td>
<td>31.0 °C</td>
</tr>
<tr>
<td>initial temperature of solution ($T_1$)</td>
<td>25.0 °C</td>
<td>24.0 °C</td>
</tr>
<tr>
<td>Change in temperature ($\Delta T$)</td>
<td>5.0</td>
<td>6.0</td>
</tr>
</tbody>
</table>

Questions

1. (a) Calculate:
   (i) average $\Delta T$
   
   Average $\Delta T = \frac{\text{change in temperature in experiment I and II}}{2}$
   
   => $\frac{5.0 + 6.0}{2} = 5.5^\circ C$

   (ii) the number of moles of solution used
   
   Moles used = molarity x volume of solution
   
   = $\frac{0.2 \times 20}{1000} = 0.004$ moles

   (iii) the enthalpy change $\Delta H$ for the reaction
   
   Heat produced $\Delta H = \text{mass of solution} \times \text{specific heat capacity} \times \Delta T$
   
   => $\frac{20 \times 4.2 \times 5.5}{1000} = \frac{462}{1000}$ Joules = $-0.462$ kJ

   (iv) State two assumptions made in the above calculations.
   
   Density of solution = density of water = 1gcm⁻³
   
   Specific heat capacity of solution = Specific heat capacity of water = 4.2 kJ⁻¹kg⁻¹K
   
   This is because the solution is assumed to be infinite dilute.

2. Calculate the enthalpy change for one mole of displacement of Cu²⁺ (aq) ions.

   Molar heat of displacement $\Delta H_d = \frac{\text{Heat produced \hspace{1mm} \Delta H}}{\text{Number of moles of fuel}}$

   => $\frac{0.462 \hspace{1mm} \text{kJ}}{0.004} = -115.5 \hspace{1mm} \text{kJ\hspace{1mm}mole}^{-1}$

3. Write an ionic equation for the reaction taking place.

   Zn(s) + Cu²⁺(aq) -> Cu(s) + Zn²⁺(aq)

4. State the observation made during the reaction.
Blue colour of copper(II)sulphate(VI) fades/becomes less blue/colourless. Brown solid deposits are formed at the bottom of reaction vessel/beaker.

5. Illustrate the above reaction using an energy level diagram.

![Energy level diagram](image)

8. The enthalpy of displacement $\Delta H_d$ of copper(II)sulphate (VI) solution is $12kJmol^{-1}$. Calculate the molarity of the solution given that 40cm$^3$ of this solution produces 2.204kJ of energy during a displacement reaction with excess iron filings.

Number of moles = \( \frac{\text{Heat produced} \ \Delta H}{\text{Molar heat of displacement} \ \Delta H_d} \)

\[
=> \frac{2.204 \text{ kJ}}{126 \text{ moles}} = 0.0206 \text{ moles}
\]

Molarity of the solution = \( \frac{\text{moles} \times 1000}{\text{Volume of solution used}} \)

\[
= \frac{0.0206 \text{ moles} \times 1000}{40} = 0.5167 \text{ M}
\]

Graphical determination of the molar enthalpy of displacement of copper

Procedure:
Place 20cm$^3$ of 0.2M copper(II)sulphate (VI) solution into a calorimeter/50cm$^3$ of plastic beaker wrapped in cotton wool/tissue paper.
Record its temperature at time $T=0$. Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds.
Place all the (1.5g) Zinc powder provided after 1 ½ minutes.
Stir the solution with the thermometer carefully and continue recording the temperature after every 30 seconds for five minutes.
Determine the highest temperature change to the nearest 0.5°C.

<table>
<thead>
<tr>
<th>Time (°C)</th>
<th>0.0</th>
<th>30.0</th>
<th>60.0</th>
<th>90.0</th>
<th>120.0</th>
<th>150.0</th>
<th>180.0</th>
<th>210.0</th>
<th>240.0</th>
<th>270.0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature</td>
<td>25.0</td>
<td>25.0</td>
<td>25.0</td>
<td>25.0</td>
<td>xxx</td>
<td>36.0</td>
<td>35.5</td>
<td>35.0</td>
<td>34.5</td>
<td></td>
</tr>
</tbody>
</table>

Plot a graph of temperature (y-axis) against time

**Questions**

1. **Show and determine the change in temperature $\Delta T$**
   From a well constructed graph $\Delta T = T_2 - T_1$ at 150 second by **extrapolation**
   $$\Delta T = 36.5 - 25.0 = 11.5^\circ C$$

2. **Calculate the number of moles of copper(II) sulphate(VI) used given the molar heat of displacement of Cu$^{2+}$ (aq)ions is 125kJ/mole**
   Heat produced $\Delta H = \text{mass of solution(m)} \times \text{specific heat capacity (c)} \times \Delta T$
   $$\Rightarrow \ 20 \times 4.2 \times 11.5 = \frac{966}{1000} \text{ Joules} = -0.966 \text{ kJ}$$
   Number of moles = $$\frac{0.966 \text{ kJ}}{125 \text{ moles}} = -0.007728 \text{ moles}$$
   Molar heat of displacement $\Delta H_d$
   $$\Rightarrow \ 0.966 \text{ kJ} = -7.728 \times 10^{-3} \text{ moles}$$
3. What was the concentration of copper(II)sulphate(VI) in moles per litre.
Molarity = \( \frac{\text{moles} \times 1000}{\text{Volume used}} \)
\[ \Rightarrow \frac{7.728 \times 10^{-3} \text{moles} \times 1000}{20} = 0.3864 \text{M} \]

4. The actual concentration of copper (II) Sulphate (VI) solution was 0.4M. Explain the differences between the two.
Practical value is lower than theoretical. Heat/energy loss to the surrounding and that absorbed by the reaction vessel decreases \( \Delta T \) hence lowering the practical number of moles and molarity against the theoretical value.

(c) Standard enthalpy/heat of neutralization \( \Delta H^0_n \)
The molar standard enthalpy/heat of neutralization \( \Delta H^0_n \) is defined as the energy/heat change when one mole of a \( H^+ (H_3O^+) \) ions react completely with one mole of \( OH^- \) ions to form one mole of \( H_2O \)/water.
Neutralization is thus a reaction of an acid /\( H^+ (H_3O^+) \) ions with a base/alkali/ \( OH^- \) ions to form salt and water only.
Strong acids/bases/alkalis are completely/fully/wholly dissociated to many free ions \( H^+ /H_3O^+ \) and \( OH^- \) ions).

(ii) for strong acid/base/alkali neutralization, no energy is used to dissociate/ionize since molecule is wholly/fully dissociated/ionized into free \( H^+ H_3O^+ \) and \( OH^- \) ions. The overall energy evolved is comparatively higher / more than weak acid-base/ alkali neutralizations.
For strong acid-base/alkali neutralization, the enthalpy of neutralization is constant at about \( 57.3 \) kJmole\(^{-1} \) irrespective of the acid-base used.
This is because ionically:
\[
\text{OH}^- (aq) + \text{H}^+ (aq) \rightarrow \text{H}_2\text{O}(l)
\]
for all wholly/fully /completely dissociated acid/base/alkali

Weak acids/bases/alkalis are partially dissociated to few free ions \( H^+ (H_3O^+ \) and \( OH^- \) ions) and exist more as molecules.
Neutralization is an exothermic\((-\Delta H)\) process.
The energy produced during neutralization depend on the amount of free ions \( H^+ H_3O^+ \) and \( OH^- \) ions existing in the acid/base/alkali reactant:
(i) for weak acid-base/alkali neutralization, some of the energy is used to dissociate/ionize the molecule into free \( H^+ H_3O^+ \) and \( OH^- \) ions therefore the overall energy evolved is comparatively lower/lesser/smaller than strong acid / base/ alkali neutralizations.
Practically \( \Delta H^0 \) can be determined as in the examples below:

**To determine the molar enthalpy of neutralization \( \Delta H_n \) of Hydrochloric acid**

**Procedure**

Place 50cm³ of 2M hydrochloric acid into a calorimeter/200cm³ plastic beaker wrapped in cotton wool/tissue paper.

Record its temperature \( T_1 \).

Using a clean measuring cylinder, measure another 50cm³ of 2M sodium hydroxide. Rinse the bulb of the thermometer in distilled water.

Determine the temperature of the sodium hydroxide \( T_2 \).

Average \( T_2 \) and \( T_1 \) to get the initial temperature of the mixture \( T_3 \).

Carefully add all the alkali into the calorimeter/200cm³ plastic beaker wrapped in cotton wool/tissue paper containing the acid.

Stir vigorously the mixture with the thermometer.

Determine the highest temperature change to the nearest 0.5°C \( T_4 \) as the final temperature of the mixture.

Repeat the experiment to complete table 1.

**(ii) enthalpy change \( \Delta H \) of neutralization.**

\[
\Delta H = (m) \times (c) \times \Delta T
\]

\( \Rightarrow (50 + 50) \times 4.2 \times 13.5 = \underline{5670} \text{ Joules} = 5.67 \text{ kJ} \)

**(iii) the molar heat of neutralization the acid.**

\[
\Delta H_n = \frac{\text{Enthalpy change } \Delta H}{\text{Number of moles}} \quad \Rightarrow \quad 5.67 \text{ kJ} = 56.7 \text{ kJ mole}^{-1}
\]

**(c) Write the ionic equation for the reaction that takes place**

\( \text{OH}^- (aq) + \text{H}^+ (aq) \rightarrow \text{H}_2\text{O(l)} \)

**(d) The theoretical enthalpy change is 57.4 kJ. Explain the difference with the results above.**

The theoretical value is higher

Heat/energy loss to the surrounding/environment lowers \( \Delta T/T_6 \) and thus \( \Delta H_n \)

Heat/energy is absorbed by the reaction vessel/calorimeter/plastic cup lowers \( \Delta T \) and hence \( \Delta H_n \)

**Sample results**
13

All high school revision materials are available on www.kusoma.co.ke

Experiment | I | II |
---|---|---|
Temperature of acid $T_1^\circ$C | 22.5 | 22.5 |
Temperature of base $T_2^\circ$C | 22.0 | 23.0 |
Final temperature of solution $T_4^\circ$C | 35.5 | 36.0 |
Initial temperature of solution $T_3^\circ$C | 22.25 | 22.75 |
Temperature change( $T_5^\circ$C) | 13.25 | 13.75 |

(a) Calculate $T_6$ the average temperature change
\[ T_6 = \frac{13.25 + 13.75}{2} = 13.5^\circ C \]

(b) Why should the apparatus be very clean?
Impurities present in the apparatus reacts with acid/base lowering the overall temperature change and hence $\Delta H^\circ_n$.

(c) Calculate the:
(i) number of moles of the acid used
\[ \text{number of moles} = \text{molarity} \times \text{volume} \]
\[ \Rightarrow 2 \times \frac{50}{1000} = 0.1 \text{moles} \]

(e) Compare the $\Delta H_n$ of the experiment above with similar experiment repeated with neutralization of a solution of:
(i) potassium hydroxide with nitric(V) acid
The results would be the same/similar.
Both are neutralization reactions of strong acids and bases/alkalis that are fully /wholly dissociated into many free $H^+$ / $H_3O^+$ and $OH^-$ ions.

(ii) ammonia with ethanoic acid
The results would be lower/ $\Delta H_n$ would be less.
Both are neutralization reactions of weak acids and bases/alkalis that are partially /partly dissociated into few free $H^+$ / $H_3O^+$ and $OH^-$ ions. Some energy is used to ionize the molecule.

(f) Draw an energy level diagram to illustrate the energy changes

Theoretical examples
1. The molar enthalpy of neutralization was experimentary shown to be 51.5kJ per mole of 0.5M hydrochloric acid and 0.5M sodium hydroxide. If the volume of
sodium hydroxide was 20cm³, what was the volume of hydrochloric acid used if the reaction produced a 5.0°C rise in temperature?

Working:
Moles of sodium hydroxide = molarity x volume

\[ \frac{0.5 \text{ M} \times 20 \text{cm}^3}{1000} = 0.01 \text{ moles} \]

Enthalpy change \( \Delta H = \frac{\Delta H_n}{\text{Moles sodium hydroxide}} \) => 51.5 = 0.515kJ

Mass of base + acid = Enthalpy change \( \Delta H \) in Joules
Specific heat capacity x \( \Delta T \)

\[ \frac{0.515 \text{kJ} \times 1000}{4.2 \times 5} = 24.5238 \text{g} \]

Mass/volume of HCl = Total volume – volume of NaOH

\[ \Rightarrow 24.5238 - 20.0 = 4.5238 \text{cm}^3 \]

Graphically \( \Delta H_n \) can be determined as in the example below:

Procedure
Place 8 test tubes in a test tube rack.
Put 5cm³ of 2M sodium hydroxide solution into each test tube. Measure 25cm³ of 1M hydrochloric acid into 100cm³ plastic beaker.
Record its initial temperature at volume of base =0.
Put one portion of the base into the beaker containing the acid.
Stir carefully with the thermometer and record the highest temperature change to the nearest 0.5°C.
Repeat the procedure above with other portions of the base to complete table 1 below

<table>
<thead>
<tr>
<th>Volume of acid(cm³)</th>
<th>25.0</th>
<th>25.0</th>
<th>25.0</th>
<th>25.0</th>
<th>25.0</th>
<th>25.0</th>
<th>25.0</th>
<th>25.0</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of alkali(cm³)</td>
<td>0</td>
<td>5.0</td>
<td>10.0</td>
<td>15.0</td>
<td>20.0</td>
<td>25.0</td>
<td>30.0</td>
<td>35.0</td>
</tr>
<tr>
<td>Final temperature(°C)</td>
<td>22.0</td>
<td>24.0</td>
<td>26.0</td>
<td>28.0</td>
<td>28.0</td>
<td>27.0</td>
<td>26.0</td>
<td>25.0</td>
</tr>
<tr>
<td>Initial temperature(°C)</td>
<td>22.0</td>
<td>22.0</td>
<td>22.0</td>
<td>22.0</td>
<td>22.0</td>
<td>22.0</td>
<td>22.0</td>
<td>22.0</td>
</tr>
<tr>
<td>Change in temperature</td>
<td>0.0</td>
<td>2.0</td>
<td>4.0</td>
<td>6.0</td>
<td>6.0</td>
<td>5.0</td>
<td>4.0</td>
<td>3.0</td>
</tr>
</tbody>
</table>

Complete the table to determine the change in temperature.
Plot a graph of volume of sodium hydroxide against temperature change.
From the graph show and determine:
(i) the highest temperature change \( \Delta T \)
\[ \Delta T = T_2 - T_1 \text{ : highest temperature}-T_2 \text{ (from extrapolating a correctly plotted graph)} \]

less lowest temperature at volume of base=0-T_1

\[ \Rightarrow 28.7 - 22.0 = 6.7 \, ^\circ \text{C} \]

(ii) the volume of sodium hydroxide used for complete neutralization
From correctly plotted graph = 16.75 cm³

(iii) Calculate the number of moles of the alkali used
Moles NaOH = molarity x volume \( V_n \)

\[ \Rightarrow 2 \times 16.75 = 0.0335 \text{ moles} \]

(iv) Calculate \( \Delta H \) for the reaction.
\[ \Delta H = \text{mass of solution mixture} \times c \times \Delta T \]

\[ \Rightarrow (25.0 + 16.75) \times 4.2 \times 6.7 \]

\[ = \frac{1174.845 \, \text{J}}{1000} = 1.174845 \, \text{kJ} \]
(iii) Calculate the molar enthalpy of the alkali:

\[ \Delta H_n = \text{Heat change} \implies 1.174845 \text{ kJ} \]

\[ \text{number of moles} = 0.0335 \text{ moles} \]

\[ = 35.0699 \text{ kJ mole}^{-1} \]

(i) Standard enthalpy/heat of solution/dissolution \( \Delta H^0_s \)

The standard enthalpy of solution \( \Delta H^0_s \) is defined as the energy change when one mole of a substance is dissolved in excess distilled water to form an infinite dilute solution. An infinite dilute solution is one which is too dilute to be diluted further. Practically the heat of solution is determined by dissolving a known mass/volume of a solute in known mass/volume of water/solvent and determining the temperature change.

To determine the heat of dissolution of ammonium nitrate(V)

Place 100cm³ of distilled water into a plastic cup/beaker/calorimeter

Put all the 5.0g of ammonium nitrate(v)/potassium nitrate(V)/ammonium chloride into the water.

Stir the mixture using the thermometer and record the temperature change after every ½ minute to complete table 1.

Continue stirring throughout the experiment.

---

**Sample results: Table 1**

<table>
<thead>
<tr>
<th>Time (minutes)</th>
<th>0</th>
<th>½</th>
<th>1</th>
<th>1½</th>
<th>2</th>
<th>2½</th>
<th>3</th>
<th>3½</th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature °C</td>
<td>22.0</td>
<td>21.0</td>
<td>20.0</td>
<td>19.0</td>
<td>19.0</td>
<td>19.5</td>
<td>20.0</td>
<td>20.5</td>
</tr>
</tbody>
</table>

Plot a graph of temperature (y-axis) against temperature

(a) From the graph show and determine:
(i) the highest temperature change $\Delta T$

$\Delta T = T_2 - T_1 :$ highest temperature $- T_2$ (from extrapolating a correctly plotted graph)

less lowest temperature at volume of base $= 0 - T_1$

$\Rightarrow 18.7 - 22.0 = 3.3 \, ^{\circ}C \quad \text{(not -3.3 \, ^{\circ}C)}$

(b) Calculate the total energy change $\Delta H$ during the reaction

$\Delta H = \text{mass of water} \times c \times \Delta T$

$\Rightarrow \Delta H = 100 \times 4.2 \times 3.3 \, ^{\circ}C = +\frac{1386 \, J}{1000} = +1.386 \, kJ$

(c) Calculate the number of moles of ammonium nitrate ($v$) used

$\text{Moles} = \frac{\text{mass}}{\text{molar mass}} \Rightarrow \frac{5.0}{80} = 0.0625 \, \text{moles}$

(d) What is the molar heat of dissolution of ammonium nitrate (V)

$\Delta H = \text{Heat change} = +1.386 \, kJ = +22.176 \, kJ \, \text{mole}^{-1}$

Number of mole $0.0625 \, \text{moles}$

(e) What would happen if the distilled water is heated before experiment was performed.

The ammonium nitrate (V) would take less time to dissolve.

Increase in temperature reduces lattice energy causing endothermic dissolution to be faster.

(e) Illustrate the above process on an energy level diagram

Graphically $\Delta H_{r}$ can be represented in an energy level diagram

Endothermic process

$NH_4(NO_3(s)) \rightarrow NH_4^+(aq) + NO_3^-(aq)$

Lattice energy (endothermic $+\Delta H$)

Hydration energy (exothermic $-\Delta H$)

$\Delta H = +22.176kJ$
c) Chemical Kinetic/Rate of reaction
The rate of a chemical reaction can be defined as the time taken for a known amount of reactants to form known amount of products.
Some reactions are too slow to be determined e.g weathering others are instantaneous.
The SI unit of time is seconds. Minutes and hours are also common.
Time is determined using a stop watch/clock.
Candidates using stop watch/clock should learn to:
(i) Press start button concurrently with starting off determination of a reaction using one hand each.
(ii) Press stop button when the reaction is over.
(iii) Record all times in seconds unless specified.
(iv) Press reset button to begin another timing
(v) Ignore time beyond seconds for stop clock/watch beyond this accuracy
(vi) Avoid accidental pressing of any button before recording
It can be very frustrating repeating a whole procedure.
The following factors theoretically and practically alter/influence/affect/determine the rate of a chemical reaction:
(a) Concentration
(b) Temperature

(a) Concentration
An increase in concentration increases the rate the rate of reaction by reducing the time taken to completion.
Theoretically, increase in concentration is a decrease in distance between reacting particles which increases their collision frequency.
Practically decreasing concentration is diluting/adding water
To demonstrate the effect of concentration on reaction rate
You are provided with
(i) sodium thiosulphate containing 40gdm⁻³ solution labeled A
(ii) 2M hydrochloric acid labeled solution B
You are required to determine the rate of reaction between solution A and B

Procedure
Measure 40cm³ of solution A into 100 cm³ glass beaker. Place it on top of a pen-mark “X”. Measure another 40cm³ of solution B. Simultaneously put solution B into solution A and start off a stop watch/clock. Determine the time taken for the pen-mark “X” to be invisible/obscurred from above. Repeat the procedure by measuring 35cm³ of solution B and adding 5cm³ of water. Complete the table 1 below by using other values os solution B and water.
### Sample results: Table 1

<table>
<thead>
<tr>
<th>Volume of solution A (cm³)</th>
<th>40</th>
<th>40</th>
<th>40</th>
<th>40</th>
<th>40</th>
<th>40</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of solution B (cm³)</td>
<td>40</td>
<td>35</td>
<td>30</td>
<td>25</td>
<td>20</td>
<td>15</td>
</tr>
<tr>
<td>Volume of water (cm³)</td>
<td>0</td>
<td>5</td>
<td>10</td>
<td>15</td>
<td>20</td>
<td>25</td>
</tr>
<tr>
<td>Time taken for x to be invisible</td>
<td>14</td>
<td>17</td>
<td>21</td>
<td>25</td>
<td>50</td>
<td>120</td>
</tr>
<tr>
<td>1/t</td>
<td>0.0714</td>
<td>0.0588</td>
<td>0.0351</td>
<td>0.04</td>
<td>0.02</td>
<td>0.0083</td>
</tr>
</tbody>
</table>

Calculate 1/t in each case

Plot a graph of 1/t (y-axis) against volume of solution B.

![Graph of 1/t vs Volume of solution B](image)

### Sample questions

(i) Explain the shape of the graph

(Straight line graph from the origin)
Decrease in concentration decreases the rate of reaction. The higher the concentration of solution B the less time taken for mark x to be obscured/invisible due to increased collision frequency between the reacting particles.

(ii) From the graph determine the time taken for the mark to be invisible at 37cm$^3$
At 37cm$^3$ then $1/t \Rightarrow 1/37 = 0.027$
From a well plotted graph:

$$1/t = 0.027 \Rightarrow 16.2602 \text{ seconds}$$

(ii) From the graph determine the volume of solution B at 100 seconds
100 seconds $\Rightarrow 1/t = 1/1000 = 0.01$
From a well plotted graph:

At $1/t = 0.01 \Rightarrow$ the volume of B $= 17.0cm^3$

(iii) State another factor that would alter the rate of the above reaction.
Temperature

(iii) State another factor that would not alter the rate of the above reaction.
Surface area
Pressure
Catalyst

(b) Temperature
An increase in temperature increases the rate of reaction.
An increase of 10°C/10K practically doubles the rate of a chemical reaction/reduces time of completion by $\frac{1}{2}$.
An increase in temperature increase the kinetic energy of reacting particles increasing their collision frequency
Practically, increase in temperature involves heating the reactants
The results and presentation should be as in the effect of concentration.
Increased temperature reverses the table I time results
i.e less time as temperature increases.

d) Qualitative analysis
Process of identifying unknown compounds
Compounds may be:
(i) Inorganic
(ii) Organic

Inorganic analysis:
This involve mainly identification of ionic compounds containing cations and anions.
Cations present in an ionic compounds are identified by adding a precipitating reagent that forms a precipitate unique to the cation/s in the compound.
The main precipitating reagents used are:
2M NaOH and/or 2M NH₃(aq)
When using 2M sodium hydroxide:
(i) No white precipitate is formed if K⁺ and Na⁺ ions are present
(ii) No white precipitate is formed if NH₄⁺ ions are present but a clourless gas with pungent smell of urine is produced which may not be recognized in a school laboratory examination setting.
(iii) White precipitate that dissolves / soluble in excess if Zn²⁺ Pb²⁺ Al³⁺ ions are present.
(iv) White precipitate that do not dissolves/insoluble in excess if Ba²⁺ Mg²⁺ Ca²⁺ ions are present.
(v) Blue precipitate that do not dissolves /insoluble in excess if Cu²⁺ ions are present. 
(vi) Green precipitate that do not dissolves/insoluble in excess if Fe²⁺ ions are present.
(vii) Brown precipitate that do not dissolves/insoluble in excess if Fe³⁺ ions are present.

When using 2M aqueous ammonia
(i) No white precipitate is formed if K⁺ , NH₄⁺ Na⁺ ions are present
(ii) White precipitate that dissolves / soluble in excess if Zn²⁺ ions are present. 
(iii) White precipitate that do not dissolves/insoluble in excess if Ba²⁺ Mg²⁺ Ca²⁺ Pb²⁺ Al³⁺ ions are present.
(iv) Blue precipitate that dissolves /soluble in excess to form a deep/royal blue solution in excess if Cu²⁺ ions are present.
(v) Green precipitate that do not dissolves/insoluble in excess if Fe²⁺ ions are present.
(vi) Brown precipitate that do not dissolves/insoluble in excess if Fe³⁺ ions are present.

Anions present in an ionic compounds are identified by adding a specific precipitating reagent that forms a precipitate unique to the specific anion/s in the compound.
(i) Lead(II)nitrate(V) solution
Lead forms insoluble PbSO₄, PbSO₃, PbCO₃, PbS, PbI₂, PbCl₂
PbS is a black precipitate, 
PbI₂ is a yellow precipitate.
All the others are white precipitates.
(a) If a Lead(II)nitrate(V) solution is added to a substance/ solution/ compound :
(i) A yellow ppt shows presence of I⁻ ions
(ii) A black ppt shows presence of S²⁻ ions
(iii) A white ppt shows presence of SO₄²⁻, SO₃²⁻, CO₃²⁻, Cl⁻

(b) If the white precipitate is added dilute nitric(V) acid:
(i) It dissolves to show presence of SO₃²⁻, CO₃²⁻
(ii) It persist/remains to show presence of SO₄²⁻, Cl⁻
(c) If the white precipitate in b(i) is added acidified potassium manganate(VII)/
dichromate(VI)
   (i) acidified potassium manganate(VII) is decolorized/orange colour of acidified
potassium dichromate(VI) turns to green to show presence of $\text{SO}_3^{2-}$
   (ii) acidified potassium manganate(VII) is not decolorized/orange colour of acidified
potassium dichromate(VI) does not turn to green/remains orange to show absence of
$\text{SO}_3^{2-}$/presence of $\text{CO}_3^{2-}$
(c) If the white precipitate in b(ii) is boiled:
   (i) It dissolves to show presence of $\text{Cl}^-$
   (ii) It persists/remains to show presence of $\text{SO}_4^{2-}$

(ii) Barium(II) nitrate(V)/Barium chloride solution
Barium(II) nitrate(V)/Barium chloride solution precipitates $\text{BaSO}_4$, $\text{BaSO}_3$, $\text{BaCO}_3$,
from $\text{SO}_4^{2-}$, $\text{SO}_3^{2-}$, $\text{CO}_3^{2-}$ ions.
Inorganic qualitative analysis require continuous practice discussion

**Sample presentation of results**
You are provided with solid Y (aluminium (III)sulphate(VI)hexahydrate). Carry out the
following tests and record your observations and inferences in the space provided.

1(a) Appearance

<table>
<thead>
<tr>
<th>Observations</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>White crystalline solid</td>
<td>Coloured ions $\text{Cu}^{2+}$, $\text{Fe}^{2+}$, $\text{Fe}^{3+}$ absent</td>
</tr>
</tbody>
</table>

(b) Place about a half spatula full of the solid into a clean dry boiling tube. Heat gently
then strongly.

<table>
<thead>
<tr>
<th>Observations</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Colourless droplets formed on the cooler part of the test tube</td>
<td>Hydrated compound/compound containing water of crystallization</td>
</tr>
<tr>
<td>Solid remains a white residue</td>
<td></td>
</tr>
</tbody>
</table>

(c) Place all the remaining portion of the solid in a test tube. Add about 10cm$^3$ of
distilled water. Shake thoroughly. Divide the mixture into five portions.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid dissolves to form a colourless solution</td>
<td>Polar soluble compound $\text{Cu}^{2+}$, $\text{Fe}^{2+}$, $\text{Fe}^{3+}$ absent</td>
</tr>
</tbody>
</table>
(i) To the first portion, add three drops of sodium hydroxide then add excess of the alkali.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference (1mark)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt, soluble in excess</td>
<td>Zn$^{2+}$, Pb$^{2+}$, Al$^{3+}$</td>
</tr>
</tbody>
</table>

(ii) To the second portion, add three drops of aqueous ammonia then add excess of the alkali.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference (1mark)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt, insoluble in excess</td>
<td>Pb$^{2+}$, Al$^{3+}$</td>
</tr>
</tbody>
</table>

(iii) To the third portion, add three drops of sodium sulphate(VI)solution.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference (1mark)</th>
</tr>
</thead>
<tbody>
<tr>
<td>No white ppt</td>
<td>Al$^{3+}$</td>
</tr>
</tbody>
</table>

(iv) I. To the fourth portion, add three drops of Lead(II)nitrate(IV)solution. **Preserve**

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference (1mark)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt</td>
<td>CO$_3^{2-}$, SO$_4^{2-}$, SO$_3^{2-}$, Cl$^-$</td>
</tr>
</tbody>
</table>

II. To the portion in (iv) I above, add five drops of dilute hydrochloric acid.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference (1mark)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt persist/remains</td>
<td>SO$_4^{2-}$, Cl$^-$</td>
</tr>
</tbody>
</table>

III. To the portion in (iv) II above, heat to boil.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference (1mark)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt persist/remains</td>
<td>SO$_4^{2-}$</td>
</tr>
</tbody>
</table>

**Organic analysis:**
This involves mainly identification of the functional group:

(i) \(-\ C=\text{C} / = \text{C} = \text{C} = / - \text{C} - \text{C}\)

(ii) \(\text{R}---\text{OH}\)

(iii) \(\text{R}---\text{COOH} / \text{H}^+\)

These functional groups can be identified by:

(i) Burning—a substance which “catches fire” must reduce in amount. Candidates should not confuse burning with flame coloration/test.

(ii) Decolorization of bromine water/chlorine water/acidified \(\text{KMnO}_4\) to show presence of

\(-\ C=\text{C} / - \text{C} = \text{C} - \text{and} \ R - \text{OH}\)

(iii) Turning orange acidified \(\text{K}_2\text{Cr}_2\text{O}_7\) to green to show presence as in above.

(iv) Turning blue litmus paper red. red litmus paper remaining red show presence of \(\text{H}^+\) ions

(d) Flame test

The colour change on a clear colourless Bunsen flame is useful in identifying some cations / metals.

A very clean metallic spatula is recommended since dirt obscures /changes the correct coloration distinct flame coloration of some compounds

<table>
<thead>
<tr>
<th>Compound</th>
<th>Colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>Barium/barium salts</td>
<td>orange</td>
</tr>
<tr>
<td>Sodium/ sodium salts</td>
<td>yellow</td>
</tr>
<tr>
<td>Potassium/potassium salts</td>
<td>Purple/lilac</td>
</tr>
<tr>
<td>Lithium/Lithium salts</td>
<td>Deep red/crimson</td>
</tr>
<tr>
<td>Calcium/ calcium salts</td>
<td>red</td>
</tr>
</tbody>
</table>
(e) Physical chemistry
Chemistry is a science subject that incorporate many scientific techniques. Examining body/council, require tabulated results/data from the candidate. This tabulated results is usually then put in a graph. The general philosophy of methods of presentation of chemistry practical data is therefore availability of evidence showing:
(i) Practical done (complete table)
(ii) Accuracy of apparatus used (decimal point)
(iii) Accuracy/care in doing experiment to get collect trend (against teachers results)
(iv) Graphical work (use of mathematical science)
(v) Calculations (Scientific mathematical integration)

(f) Sample practicals
Name …………………………………………………..Class………………Index No………………

Candidate’s signature………………………………………………

Date done…………………..Date marked………………Date revised…………………..

233/3
CHEMISTRY Paper 3
PRACTICAL.
Pre-KCSE Practice 1: 2013
MARKS SCHEME

Instruction to Candidate
Write your name and index number in the spaces provided above.
Sign and write the date of examination in the spaces provided above
Answer all questions in the spaces provided.
Mathematical tables and electronic calculators may be used.
All working must be clearly shown where necessary.
This paper consist of 8 printed pages.
Candidates should check the question paper to ascertain that all the pages are printed
and indicated and that no questions are missin

For examiners use only

<table>
<thead>
<tr>
<th>Question</th>
<th>Maximum score</th>
<th>Candidates score</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>20</td>
<td>20</td>
</tr>
<tr>
<td>2</td>
<td>10</td>
<td>10</td>
</tr>
<tr>
<td>3</td>
<td>10</td>
<td>10</td>
</tr>
<tr>
<td>Total score</td>
<td>40</td>
<td>40</td>
</tr>
</tbody>
</table>
1. You are provided with:
   (i) solution L containing 5.0g per litre of a dibasic organic acid $\text{H}_2\text{X}.2\text{H}_2\text{O}$.
   (ii) solution M which is acidified potassium manganate(VII)
   (iii) solution N a mixture of sodium ethanedioate and ethanedioic acid
   (iv) 0.1M sodium hydroxide solution P
   (v) 1.0M sulphuric(VI)
You are required to:
   (i) standardize solution M using solution L
   (ii) use standardized solution M and solution P to determine the % of sodium ethanedioate in the mixture.

Procedure 1
Fill the burette with solution M. Pipette 25.0cm³ of solution L into a conical flask. Heat this solution to about 70°C **(but not to boil)**. Titrate the hot solution L with solution M until a permanent pink colour just appears. Shake thoroughly during the titration. Repeat this procedure to complete table 1.

<table>
<thead>
<tr>
<th>Table 1</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading (cm³)</td>
</tr>
<tr>
<td>20.0</td>
</tr>
<tr>
<td>Initial burette reading (cm³)</td>
</tr>
<tr>
<td>Volume of N used (cm³)</td>
</tr>
</tbody>
</table>

(a) Calculate the average volume of solution L used
(1mk)

$$\frac{20.0 + 20.0 + 20.0}{3} = 20.0 \text{cm}^3$$

(b) Given that the concentration of the dibasic acid is 0.05molesdm⁻³. Determine the value of x in the formula $\text{H}_2\text{X}.2\text{H}_2\text{O}$ (H=1.0, O=16.0)
(1mark)

Molar mass $\text{H}_2\text{X}.2\text{H}_2\text{O} = \frac{\text{mass / litre}}{\text{moles / litre}} = \frac{5.0\text{g/litre}}{0.05\text{molesdm}^{-3}} = 100 \text{ g}$

$\text{H}_2\text{X}.2\text{H}_2\text{O} = 100$

$$X = 100 - (2 \times 1) + 2 \times (2 \times 1) + (2 \times 16) = 100 - 34 = 62$$
(c) Calculate the number of moles of the dibasic acid H₂X₂H₂O. \( \text{Moles} = \text{molarity} \times \text{pipette volume} \) \( \Rightarrow \)
\[
\frac{0.05 \times 25}{1000} = \frac{0.00125}{1.25 \times 10^{-3}} \text{ moles}
\]

(d) Given the mole ratio manganate(VII)\((\text{MnO}_4^-)\): acid \(\text{H}_2\text{X}\) is 2:5, calculate the number of moles of manganate(VII) \((\text{MnO}_4^-)\) in the average titre. \( \text{Moles} \ \text{H}_2\text{X} = \frac{2}{5} \ \text{moles of MnO}_4^- \)
\[
\Rightarrow \frac{2}{5} \times \frac{0.0125}{1.25 \times 10^{-2}} \text{ moles} = \frac{0.0005}{5.0 \times 10^{-4}} \text{ moles}
\]

(e) Calculate the concentration of the manganate(VII) \((\text{MnO}_4^-)\) in moles per litre. \( \text{Moles per litre/molarity} = \frac{\text{moles} \times 1000}{\text{average burette volume}} \)
\[
\Rightarrow \frac{0.0005}{5.0 \times 10^{-4}} \text{ moles} \times 1000 = 0.02083 \text{ moles l}^{-1} / \text{M}
\]

Procedure 2
With solution M still in the burette, pipette 25.0cm³ of solution N into a conical flask. Heat the conical flask containing solution N to about 70°C. Titrate while hot with solution M. Repeat the experiment to complete table 2.

<table>
<thead>
<tr>
<th>Table 2</th>
<th>(2marks)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading (cm³)</td>
<td>12.5</td>
</tr>
<tr>
<td>Initial burette reading (cm³)</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of N used (cm³)</td>
<td>12.5</td>
</tr>
</tbody>
</table>

(a) Calculate the average volume of solution L used
\( \text{(1mk)} \)
\[
\frac{12.5 + 12.5 + 12.5}{3} = 12.5 \text{ cm³}
\]

(b) Calculations:
(i) How many moles of manganate(VII) ions are contained in the average volume of solution M used?
\( \text{Moles} = \frac{\text{molarity of solution M} \times \text{average burette volume}}{1000} \)
\[
\Rightarrow \frac{0.02083 \text{ moles l}^{-1} / \text{M} \times 12.5}{1000} = \frac{0.00026}{2.6 \times 10^{-4}} \text{ moles}
\]
(ii) The reaction between manganate(VII) ions and ethanedioate ions that reacted with is as in the equation:
2MnO₄⁻ (aq) + 5C₂O₄²⁻ (aq) + 16H⁺ (aq) → 2Mn²⁺(aq) + 10CO₂(g) + 8H₂O(l)

Calculate the number of moles of ethanedioate ions that reacted with manganate (VII) ions in the average volume of solution M.

1 mark

From the stoichiometric/ionic equation:

mole ratio MnO₄⁻ (aq): C₂O₄²⁻ (aq) = 2:5

=> moles C₂O₄²⁻ = \( \frac{5}{2} \) moles MnO₄⁻

=> \( \frac{5}{2} \) x 0.00026 / 2.5 x 10⁻³ moles

= \( \frac{0.00065}{6.5 \times 10^{-4}} \) moles

(iii) Calculate the number of moles of ethanedioate ions contained in 250 cm³ solution N.

25 cm³ pipette volume → 0.00065 / 6.5 x 10⁻⁴ moles

250 cm³ → 0.0065 / 6.5 x 10⁻³ moles x 250 = \( \frac{0.0065}{6.5 \times 10^{-3}} \) moles

Procedure 3

Remove solution M from the burette and rinse it with distilled water. Fill the burette with sodium hydroxide solution P. Pipette 25 cm³ of solution N into a conical flask and add 2-3 drops of phenolphthalein indicator. Titrate this solution N with solution P from the burette. Repeat the procedure to complete table 3.

Table 3

<table>
<thead>
<tr>
<th></th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading (cm³)</td>
<td>12.5</td>
<td>12.5</td>
<td>12.5</td>
</tr>
<tr>
<td>Initial burette reading (cm³)</td>
<td>0.0</td>
<td>0.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of N used (cm³)</td>
<td>12.5</td>
<td>12.5</td>
<td>12.5</td>
</tr>
</tbody>
</table>

(a) Calculate the average volume of solution L used

\( \frac{12.5 + 12.5 + 12.5}{3} = 12.5 \) cm³

(1 mark)

(b) Calculations:

(i) How many moles of sodium hydroxide solution P were contained in the average volume?

1 mark

\[ \text{Moles} = \text{molarity of solution P} \times \text{average burette volume} \]

\[ = \frac{0.1 \text{ moles}}{1000} \times 24.9 = \frac{0.00249}{2.49 \times 10^{-3}} \text{ moles} \]

(ii) Given that NaOH solution P reacted with the ethanedioate ions from the acid only and the equation for the reaction is:

2NaOH (aq) + H₂C₂O₄ (aq) → Na₂C₂O₄(g) + 2H₂O(l)
Calculate the number of moles of ethanedioic acid that were used in the reaction. (1 mk)

*From the stoichiometric equation, mole ratio*

NaOH(aq): H₂C₂O₄ (aq) = 2:1

=> moles H₂C₂O₄ = 1/2 moles NaOH

=> 1/2 x 0.00249 / 2.49 x 10⁻³ moles

= 0.001245 / 1.245 x 10⁻³ moles.

(iii) How many moles of ethanedioic acid were contained in 250cm³ of solution N? (1 mark)

25cm³ pipette volume -> 0.001245 / 1.245 x 10⁻³ moles x 250

250cm³ -> 0.01245 / 1.245 x 10⁻² moles

(iii) Determine the % by mass of sodium ethanedioate in the mixture (H= 1.0, O=16.0, C=12.0 and total mass of mixture = 2.0 g in 250cm³ solution) (1 mark)

Molar mass H₂C₂O₄ = 90.0 g

Mass of H₂C₂O₄ in 250cm³ =

moles in 250cm³ x molar mass H₂C₂O₄

=> 0.01245 / 1.245 x 10⁻² moles x 90.0

= 1.1205 g

% by mass of sodium ethanedioate

= (Mass of mixture - mass of H₂C₂O₄) x 100 %

Mass of mixture

=> 2.0 - 1.1205 g = 43.975 %

2. You are provided with 5.0 g solid B. You are to determine the molar mass of solid B.

**Procedure**

Place 100 cm³ of liquid L into a plastic beaker. Determine its temperature and record it at time = 0 in Table 2 below. Stir continuously using the thermometer and record the highest temperature change to the nearest 0.5°C after every 30 seconds. After 120 seconds, add all solid B. Continue stirring and recording the temperature to complete Table 2.

<table>
<thead>
<tr>
<th>Table 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Time (seconds)</td>
</tr>
<tr>
<td>Temperature (°C)</td>
</tr>
</tbody>
</table>

(2 mark)
(a) Plot a graph of temperature against time (x-axis) (3 marks)

(b) From the graph show and determine (2 marks)

(i) the highest temperature change $\Delta T$

$$\Delta T = T_2 - T_1 \quad \Rightarrow \quad 13.4 - 20 = 6.6^\circ C$$

Note $\Delta T$ is not - 6.6$^\circ C$

(ii) the temperature of the mixture at 130 seconds

From extrapolation at 130 seconds = 19.2$^\circ C$

(iii) the time when all the solid first dissolved

From extrapolation of the lowest temperature = 220 Seconds
(d) Calculate the heat change for the reaction. (Assume density of liquid L is 1.0gcm$^{-3}$) specific heat capacity is 4.2Jkg$^{-1}$K$^{-1}$. (1mark)

$$\Delta H = \text{mass of liquid L} \times c \times \Delta T \Rightarrow 100 \times 4.2 \times 6.6 = +\frac{2772}{1000} \text{J} = +2.772 \text{kJ}$$

(e) Given the molar enthalpy of dissolution of Solid B in liquid L is +22.176kJ mole$^{-1}$, determine the number of moles of B used. (1mark)

$$\text{Moles of B} = \frac{\Delta H}{\Delta H_s} = \frac{+2.772 \text{kJ} + 22.176 \text{kJ mole}^{-1}}{0.125 \text{moles}}$$

(f) Calculate the molar mass of B. (1mark)

$$\text{Molar mass of B} = \frac{\text{Mass used}}{\text{Moles used}} \Rightarrow \frac{5.0}{0.125 \text{moles}} \Rightarrow 40 \text{g}$$

3(a) You are provided with solid Y. Carry out the following tests and record your observations and inferences in the space provided.

(i) Appearance

<table>
<thead>
<tr>
<th>Observations</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>White crystalline solid</td>
<td>Coloured Fe$^{2+}$, Fe$^{3+}$, Cu$^{2+}$ ions absent</td>
</tr>
</tbody>
</table>

(ii) Place about a half spatula full of the solid into a clean dry boiling tube. Heat gently then strongly.

<table>
<thead>
<tr>
<th>Observations</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Colourless droplets forms on the cooler parts of test tube</td>
<td>Hydrated compound/salt</td>
</tr>
<tr>
<td>Solid remain white</td>
<td></td>
</tr>
</tbody>
</table>

(ii) Place all the remaining portion of the solid in a test tube. Add about 10cm$^3$ of distilled water. Shake thoroughly. Divide the mixture into five portions.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid dissolves to form a colourless solution absent</td>
<td>Coloured Fe$^{2+}$, Fe$^{3+}$, Cu$^{2+}$ ions absent</td>
</tr>
</tbody>
</table>

I. To the first portion add three drops of universal indicator. (1mark)

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH= 4</td>
<td>weakly acidic solution</td>
</tr>
</tbody>
</table>
II. To the second portion, add three drops of aqueous ammonia then add excess of the alkali.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt, insoluble in excess</td>
<td>Al(^{3+}), Pb(^{2+})</td>
</tr>
</tbody>
</table>

III. To the third portion, add three drops of sodium sulphide solution.

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>No black ppt</td>
<td>Al(^{3+})</td>
</tr>
</tbody>
</table>

IV. To the fourth portion, add three drops of acidified Lead(II)nitrate(IV)solution. Heat to boil

<table>
<thead>
<tr>
<th>Observation</th>
<th>Inference</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt, persist/remains on boiling</td>
<td>SO(_4^{2-})</td>
</tr>
</tbody>
</table>

(b) You are provided with solid P. Carry out the following tests and record your observations and inferences in the space provided.

(i) Place a portion of solid P on a clean metallic spatula and introduce it on a Bunsen flame.

Solid burns with a yellow sooty flame  \(\text{C} \equiv \text{C}  //  \text{C} \equiv \text{C}\) bonds  

(ii) Add all the remaining solid to about 10cm^3 of water in a test tube and shake well. Divide the mixture into 4 portions. 

Solid dissolves to form a colorless solution  
Polar organic compound

I. To the 1st portion, test with litmus papers

Red litmus paper remain red  \(\text{H}^+\) ions  
Blue litmus paper turn blue
II. To the 2nd portion, add a little sodium hydrogen carbonate (1/2 mark)

<table>
<thead>
<tr>
<th>Effervescence/fizzing/bubbles</th>
<th>H⁺ ions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Colourless gas produced</td>
<td></td>
</tr>
</tbody>
</table>

III. To the 3rd portion, and three drops of solution M. Warm (1/2 mark)

<table>
<thead>
<tr>
<th>Acidified KMnO₄ is decolorized</th>
<th>( R - OH, C = C ) # C ( \equiv ) C bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td>// solution M is decolorized</td>
<td></td>
</tr>
</tbody>
</table>

IV. To the 4th portion, add three drops of bromine water (1/2 marks)

| Bromine water is decolorized   | \( C | = C // C \equiv C \) bonds |
|--------------------------------|--------------------------------|
|                                |                               |
Requirements for each Candidates:

- 0.05 M Oxalic acid labeled **Solution L**
- 0.01M Potassium manganate (VII) labeled **Solution M**
- 0.03M Oxalic acid labeled **Solution N**
- 0.1M Sodium hydroxide labeled **Solution P**
- 1.0M sulphuric(VI)acid.
- 15 0cm³ distilled water labeled **Liquid L**
- 50cm³ burette
- 25cm³ pipette
- Two clean conical flasks
- Pipette filler
- -10 -110°C Thermometer
- Stop watch/clock
- 200cm³ clean beaker
- **5.4g ammonium chloride/8.0g Ammonium nitrate(V)** labeled **Solid B** weighed accurately
- About 2.0 g of hydrated Aluminium sulphate labeled **Solid Y**
- About 2.0g of Citric acid labeled **Solid P**
- One boiling tube
- Six clean dry test tubes
- Pair of litmus papers(red and blue)
- Clean metallic spatula

Access to bench reagents/apparatus

- Means of heating
- 2M aqueous ammonia
- 0.1M acidified Lead(II)nitrate(V)
- Bromine water
- 0.1M sodium sulphide
- About 0.1g Sodium hydrogen carbonate
- Universal indicator solution
- pH chart
You are provided with:
- Solution A containing an oxidizing agent A;
- Solution B, 0.05M aqueous sodium thiosulphate;
- Solution C containing a reducing agent C;
- Aqueous Potassium iodide;
- Solution D, starch solution.

You are required to determine the:
- Concentration of solution A
- Rate of reaction between the oxidizing agent A and the reducing agent C.

**Procedure 1**

1. Using a pipette and pipette filter, place 25.0 cm$^3$ of solution A into a 250 ml conical flask.
2. Measure 10 cm$^3$ of aqueous potassium iodide and add it to solution A in the conical flask.
   Shake the mixture. Add 10 cm$^3$ of 2 M sulphuric(VI) acid to the mixture and shake.
3. Fill a burette with solution B and use it to titrate the mixture in the conical flask until it just turns orange yellow. Add 2 cm$^3$ of solution D to the mixture in the conical flask. Shake thoroughly. Continue titrating until the mixture just turns colourless. Record your results in Table 1 below.
4. Repeat the procedure and complete Table 1. **Retain the remainder** of solution A and D for use in Procedure II.

**Table 1**

<table>
<thead>
<tr>
<th></th>
<th>I</th>
<th>II</th>
<th>III</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final burette reading</td>
<td>20.0</td>
<td>20.0</td>
<td>20.0</td>
</tr>
<tr>
<td>Initial burette reading</td>
<td>0.0</td>
<td>0.0</td>
<td>0.0</td>
</tr>
<tr>
<td>Volume of solution B used (cm$^3$)</td>
<td>20.0</td>
<td>20.0</td>
<td>20.0</td>
</tr>
</tbody>
</table>
(a) Calculate the:

(i) average volume of solution B used (1mk)
\[
\frac{20.0 + 20.0 + 20.0}{3} = 20.0 \text{ cm}^3
\]

(ii) number of moles of sodium thiosulphate (1mk)

\[
\text{Moles} = \text{molarity} \times \text{burette volume} \times 1000
\]
\[
=> \frac{0.05 \times 20.0}{1000} = 0.001 / 1.0 \times 10^{-3} \text{ moles}
\]

(b) Given that one mole of A reacts with six moles of sodium thiosulphate, calculate the:

(i) number of moles of A that were used (1mk)

\[
\text{Mole ratio A:B} = 1:6
\]
\[
=> \text{Moles A} = \frac{0.001 / 1.0 \times 10^{-3}}{6} = 0.00016 / 1.6 \times 10^{-4} \text{ moles}
\]

(ii) concentration of solution A in moles per litre (2mk)

\[
\text{Molarity of solution A} = \frac{\text{moles} \times 1000}{\text{Pipette volume}}
\]
\[
=> \frac{0.00016 / 1.6 \times 10^{-4} \text{ moles} \times 1000}{20} = 0.008 / 8.0 \times 10^{-4} \text{ M}
\]

Procedure II

1. Label six test tubes as 1, 2, 3, 4, 5 and 6 and place them on a test tube rack.
2. Using a clean burette, measure the volumes of distilled water as shown in table 2 into the labeled test tubes.
3. Using a burette, measure the volumes of solution A shown in table 2 into each of the test tubes.
4. Clean the burette and rinse it with about 5cm³ of solution C
5. Using the burette, measure 5cm³ of solution C and place it into a 100ml beaker.
6. Using a 10ml measuring cylinder, measure 5cm³ of solution D and add it to the beaker containing solution C. Shake the mixture.
7. Pour the contents of test tube number 1 to the mixture in the beaker and immediately start off stop watch/clock. Swirl the contents of the beaker. Record the time taken for a blue colour to appear in table 2.
8. Repeat steps 5 to 7 using the contents of test-tube 2, 3, 4, 5, and 6.

9. Complete table 2 by computing \( \text{Rate} = \frac{1}{\text{Time}} \) (S\(^{-1}\)).

**Table 2 (Sample results)**

<table>
<thead>
<tr>
<th>Test-tube number</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of distilled water (cm(^3))</td>
<td>0</td>
<td>2</td>
<td>3</td>
<td>5</td>
<td>6</td>
<td>7</td>
</tr>
<tr>
<td>Volume of solution A (cm(^3))</td>
<td>10</td>
<td>8</td>
<td>7</td>
<td>5</td>
<td>4</td>
<td>3</td>
</tr>
<tr>
<td>Time (seconds)</td>
<td>40.0</td>
<td>60.0</td>
<td>70.0</td>
<td>90.0</td>
<td>100.0</td>
<td>110.0</td>
</tr>
<tr>
<td>Rate = ( \frac{1}{\text{time}} ) (S(^{-1}))</td>
<td>0.025 (2.5 \times 10^{-2})</td>
<td>0.0167 (1.67 \times 10^{-2})</td>
<td>0.0143 (1.43 \times 10^{-2})</td>
<td>0.0111 (1.11 \times 10^{-2})</td>
<td>0.01</td>
<td>(1.11 \times 10^{-2})</td>
</tr>
</tbody>
</table>

Plot a graph of rate (y-axis) against volume of solution A (3mk).

**Sketch graph of rate against time**

![Graph showing rate vs volume of solution A](image)
(b) What time would be taken for the blue colour to appear if the experiment was repeated using 4cm³ of distilled water and 6cm³ of solution A? (2mk)

From a correctly plotted graph
\[ \frac{1}{t} \text{ at } 6\text{cm}³ = 0.0125 \sqrt{\text{ }} \implies t = \frac{1}{0.0125} = 80\text{ seconds} \sqrt{\text{ }}

2. You are provided with solid E. Carry out the experiments below. Write your observations and inferences in the spaces provided
(a) Place all solid in a boiling tube. Add 20cm³ of distilled water and shake until all the solid dissolves. Label the solution as solution E. Use solution E for experiment (i) and (ii)
(i) To 2cm³ of solution E in a test tube in each of experiment I, II, III, and IV add:

I. two drops of aqueous sodium sulphate(VI)

<table>
<thead>
<tr>
<th>Observation (1mk)</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White precipitate</td>
<td>Pb²⁺ Ba²⁺ Ca²⁺</td>
</tr>
</tbody>
</table>

II. five drops of aqueous sodium chloride

<table>
<thead>
<tr>
<th>Observation (1mk)</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td>White ppt</td>
<td>Ca²⁺ Ba²⁺</td>
</tr>
</tbody>
</table>

III. two drops of barium chloride

<table>
<thead>
<tr>
<th>Observation (1mk)</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td>No white ppt</td>
<td>SO₄²⁻ SO₃²⁻ CO₃²⁻</td>
</tr>
</tbody>
</table>

IV. two drops of lead(II) nitrate(V)

<table>
<thead>
<tr>
<th>Observation (1mk)</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td>No white ppt</td>
<td>Cl⁻</td>
</tr>
</tbody>
</table>
(ii) To 2cm³ of solution E in a test tube, add 5 drops of aqueous sodium hydroxide. Add the piece of aluminium foil provided to the mixture and shake. Warm the mixture and test any gas produced with both blue and red litmus papers.

**Observations (1mk)**

<table>
<thead>
<tr>
<th>Blue litmus paper remain blue</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Red limus paper turn blue</td>
<td>NO₃⁻</td>
</tr>
<tr>
<td>Effervescence /fizzing/ bubbles</td>
<td></td>
</tr>
</tbody>
</table>

**Note:**

Solid E is Calcium nitrate(V) / Barium nitrate(V)

3. You are provided with solid F. Carry out the following test. Write your observations and inferences in the spaces provided.

(a) Place all of solid F in a boiling tube. Add about 20cm³ of distilled water and shake until all the solid dissolves. Label the solution as solution F.

Add about half of the solid sodium hydrogen carbonate provided to 2cm³ of solution F

**Observations (1mk)**

<table>
<thead>
<tr>
<th>No effervescence/fizzing</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>H⁺ absent</td>
</tr>
</tbody>
</table>

(b)(i) Add about 10cm³ of dilute hydrochloric acid to the rest of solution F in the boiling tube. Filter the mixture. Wash the residue with about 2cm³ of distilled water. Dry the residue between filter papers. Place about one third of the dry residue on a metallic spatula and burn it in a Bunsen burner flame.

**Observations (1mk)**

<table>
<thead>
<tr>
<th>Solid burns with a yellow sooty flame</th>
<th>Inferences (1mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>C≡C</td>
</tr>
<tr>
<td></td>
<td>// C≡C</td>
</tr>
</tbody>
</table>

(ii) Place all the remaining residue into a boiling tube. Add about 10cm³ of distilled water and shake thoroughly. Retain the mixture for the tests in (c)

**Observations (½mk)**

<table>
<thead>
<tr>
<th>Solid dissolves to a colourless solution</th>
<th>Inferences (½mk)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Polar compound</td>
</tr>
</tbody>
</table>
(c) Divide the mixture into two portions:

(i) to the first portion, add the rest of the solid sodium hydrogen carbonate

**Observations (1mk)**

- Effervescence/fizzing

**Inferences (1mk)**

- $H^+$

(ii) to the second portion, add two drops of bromine water.

**Observations (1mk)**

- Bromine water decolorized

**Inferences (1mk)**

- $\text{C} \equiv \text{C} // \text{C} \equiv \text{C} \text{ bonds}$

**************END***************

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- Marking schemes to most of the revision papers.
- K.C.S.E past papers.
- University past papers.
- K.A.S.N.E.B past papers e.g. C.P.A
- Primary school schemes of work.
- Primary school revision papers to all classes.
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- K.C.P.E past papers.