2018 SH2 H2 Chemistry P4 Solutions

Answer **all** the questions in the spaces provided.

1 Determination of the average relative formula mass of a mixture of two copper salts

In this experiment, you will determine the average relative formula mass of a mixture of two copper salts by titration.

A solution of the copper salts mixture reacts with excess acidified potassium iodide, producing iodine.

$$2Cu^{2+}(aq) + 4I^{-}(aq) \rightarrow 2CuI(s) + I_2(aq)$$

This iodine is then titrated with aqueous sodium thiosulfate, using starch indicator.

FA 1 is an aqueous solution of the copper salt prepared by dissolving 26.0 g of the salt mixture to make 1.00 dm^3 of solution.
FA 2 is dilute sulfuric acid, H_2SO_4.
FA 3 is aqueous potassium iodide, KI.
FA 4 is 1.50 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.
starch indicator

(a) Preparation of diluted FA 4

1. Pipette 25.0 cm^3 of **FA 4** into the 250 cm^3 graduated flask.
2. Make up the contents of the flask to the 250 cm^3 mark with deionised water.
3. Stopper the flask and mix the contents thoroughly to ensure a homogeneous solution.
   This prepared solution is **diluted FA 4**.

Titration

1. Fill the burette with **diluted FA 4**.
2. Pipette 25.0 cm^3 of **FA 1** into a conical flask.
3. Use the measuring cylinder to add approximately 10.0 cm^3 of **FA 2** to the same conical flask.
4. Use the measuring cylinder to add approximately 20.0 cm^3 of **FA 3** to the mixture in the conical flask. The mixture will appear brown, due to iodine produced in the reaction.
5. Begin your rough titration by adding **diluted FA 4** from the burette until the intensity of the brown colour decreases.
6. Add 10 drops of starch indicator. The mixture will become darker.
7. Continue titrating until the dark colour is discharged. The mixture should appear off-white. This is the end-point.
8. Add **one** drop of starch indicator to check that no traces of dark colour are produced.
9. If the mixture stays off-white, the titration is completed. If some dark colour is produced, because iodine is still present, continue the titration until mixture appears off-white.
10. Record your burette readings and the rough titre in the space below.

11. Carry out as many accurate titrations as you think necessary to obtain consistent results.

12. Make sure any recorded results show the precision of your prac
13. Record in a suitable form below all of your burette readings and of **diluted FA 4** added in each accurate titration.

I	
II	
III	
IV	
V	
VI	
VII	

I All the following data is recorded
initial and final burette readings for **two** (or more) titrations

II Titre values recorded for accurate titrations, **and**
Appropriate headings and units
- initial / start (burette) reading
- final / end (burette) reading
- titre **or** volume used / added *(not "difference")*

unit: / cm^3 (for each heading) **or** cm^3 unit given for each volume recorded.

III: All burette readings are recorded to the nearest 0.05 cm^3. (2 d.p)
(Rough reading do not need 2 d.p)

including 0.00 cm^3 (if this was the initial reading).

*Do **not** award if: (reverse burette readings)*
- *50(.00) is used as an initial burette reading*
- *more than one final burette reading is 50.(00)*

any burette reading is greater than 50.(00)

IV: at least two titres within 0.10 cm^3

[7]

From your accurate titration results, obtain a suitable value for the volume of **diluted FA 4** to be used in your calculations.

Show clearly how you obtained this value.

Candidate calculates the mean correctly.
- Candidate must take the average of two (or more) titres within 0.10 cm^3.
- Working / explanation ***or*** ticks must be shown
- The mean should be quoted to **2 d.p**, and be rounded to nearest 0.01 cm^3.

(e.g. 26.667 cm^3 must be rounded to 26.67 cm^3)

The iodine produced required ..... cm^3 of **diluted FA 4**.
[1]

(b) Calculations

Show your working and appropriate significant figures in the final answer to **each** step of your calculations.

(i) Calculate the number of moles of sodium thiosulfate, $Na_2S_2O_3$, in the volume of **diluted FA 4** obtained in **(a)**.

No of moles of thiosulfate used = $1.50 \times \frac{25.0}{250}$ × mean titre / 1000

Or No of moles of thiosulfate used = 0. 150 × mean titre / 1000
(expressed to 3 or 4 sig fig)

moles of $Na_2S_2O_3$ = ... mol
[1]

(ii) Balance the ionic equation for the reaction of iodine with sodium thiosulfate. State symbols are not required.

…..... I_2 + ... $S_2O_3^{2-}$ → ..... $S_4O_6^{2-}$ + ……. I^-

Hence calculate the number of moles of iodine that reacted with the number of moles of $Na_2S_2O_3$ calculated in **(i)**.

Equation balanced I_2 + **2** $S_2O_3^{2-}$ → $S_4O_6^{2-}$ + **2**I^-
no of moles of I_2 = 0.5 × ans. in (i)

moles of I_2 = ... mol
[1]

(iii) Using your answer to **(ii)**, calculate the number of moles of copper(II) ions in 25.0 cm^3 of **FA 1**.

No of moles of copper(II) ions = 2 × answer (ii)
(expressed to 3 or 4 sig fig)

moles of Cu^{2+} ions = ... mol
[1]

(iv) Using your answer to **(iii)** and the information on page 2, calculate the average relative formula mass of the copper salts in **FA 1**.

Average Mr = [26.0 / ans (iii)] × 25 / 1000

*Average M*r of copper salts = ...
[1]

(v) Write the full electronic configuration of $_{29}Cu$ in CuI.

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$

(vi) Hence, explain why solid CuI appears white in colour.

- d subshell is fully filled, no vacant / half-filled d orbitals
- electrons cannot be promoted from lower energy d orbitals to higher energy d orbitals / d-d transition not possible
- No wavelength of visible spectrum is absorbed, hence solid appears white

[1]

[Total: 14]

2 Qualitative analysis

In this question, you will deduce the two anions present in **FA 1.** Perform the tests described in **Table 2** and record your observations in the table. Test and identify any gases evolved.

If any solution is warmed, a boiling tube MUST be used.

Table 2

	Test	Observations
(a)	To a 1 cm depth of **FA 1** in a test-tube, add aqueous silver nitrate.	Solution remained blue / no ppt
(b)	To a 0.5 cm depth of **FA 1** in a boiling tube, add aqueous sodium hydroxide and add one piece of aluminium foil and warm.	Blue ppt formed, insoluble in excess sodium hydroxide Black ppt formed upon heating Colourless, pungent gas / **gas evolved turns (damp red) litmus blue**
(c)	To a 1 cm depth of **FA 1** in a test-tube, add aqueous barium chloride followed by nitric acid.	White ppt formed **and** insoluble in excess nitric acid
(d)	To a 1 cm depth of **FA 1** in a test-tube, add an equal volume of sulfuric acid followed by $KMnO_4$.	**FA 1** remained blue. Purple $KMnO_4$ remained/ not decolourised

[3]

(e) From your observations, state the anions present in **FA 1**. Explain your answers.

Nitrate NO_3^- + Sulfate SO_4^{2-}

Explanation for NO_3^-

Both Nitrate or Nitrite is reduced by Al foil in NaOH(aq) to give NH_3 gas +
Since $KMnO_4$ remained purple, **FA 1** does not contain reducing agent NO_2^-.

Or
Since no brown NO_2 gas is evolved with nitric acid, **FA 1** contains Nitrate NO_3^-

Explanation for SO_4^{2-}
Since a white ppt is formed with $BaCl_2$, **and** remained in nitric acid, anion in FA1 cannot be SO_3^{2-}.
Since Since $KMnO_4$ remained purple, **FA 1** does not contain reducing agent SO_3^{2-}.

[3]

[Total: 6]

3 Investigation of thermal decomposition of sodium hydrogencarbonate

Sodium hydrogencarbonate, $NaHCO_3$, is used as baking soda in cooking. Baking soda may also contain small amounts of other chemicals.

When baking soda is heated, carbon dioxide is produced. In this experiment, you will investigate the reaction taking place when the sodium hydrogencarbonate in baking soda is thermally decomposed.

FA 5 is baking soda (impure $NaHCO_3$).

(a) Method

Record all your readings in the space below.

1. Weigh the crucible with its lid.
2. Transfer all the **FA 5** from the container into the crucible.
3. Weigh the crucible, lid and **FA 5**.
4. Calculate and record the mass of **FA 5** used.
5. Place the crucible and contents on a pipe-clay triangle.
6. Heat gently, **with the lid off**, for approximately one minute.
7. Heat strongly, **with the lid off**, for a further three minutes.
8. Replace the lid and leave the crucible to cool for about ten minutes.

While the crucible is cooling, you should work on other questions.

9. When it has cooled down, weigh the crucible with its lid and contents.
10. Heat strongly, **with the lid off**, for a further two minutes.
11. Replace the lid and leave the crucible to cool for ten minutes.
12. When it has cooled down, weigh the crucible with its lid and contents.
13. Calculate and record the mass of residue obtained.
14. This residue is **FA 6**. Keep this for use in **3(d)**.

Results

Four weighings recorded and correct headings given **and**
mass of **FA 5** used **and**
mass of residue recorded

Mass	/g
Crucible + lid	a
Crucible + lid + **FA 5** before heating	b
Crucible + lid + content after 1st heating	c
Crucible + lid + content after 2nd heating	d
FA 5 used	b-a = e
Residue OR **FA 6**	d-a = f

If 'mass' not written then 'g' must be with each entry.

Use of lid must be consistent

I	
II	
III	
IV	

II

- All weighings recorded to 3 decimal places
- Mass of **FA 5** and **FA 6** / residue must be correctly subtracted

[4]

(b) Calculations

(i) Given that the percentage purity by mass of **FA 5** is 95.8%, calculate the mass of sodium hydrogencarbonate in the sample of **FA 5** that you weighed out.

Mass $NaHCO_3$ = (95.8 / 100) × mass of **FA 5** used

mass of $NaHCO_3$ in **FA 5** weighed out = ... g

(ii) Calculate the mass of impurity present in your sample of **FA 5**.

Mass impurity = mass of **FA 5** – answer **(i)**
or mass impurity = 4.2/ 100 x mass **FA 5**

mass of impurity = ... g
[1]

(iii) The impurity in **FA 5** does not decompose when it is heated.

This means that the residue, **FA 6**, contains the mass of impurity calculated in **(ii)** together with the solid decomposition product of sodium hydrogencarbonate.

Calculate the mass of the solid decomposition product.

Mass of decomposition solid
= mass of residue (**FA 6)** from table – mass of impurity **(ii) and** expressed to 3 d.p / 3 s.f

or mass of decomposition solid = mass of $NaHCO_3$ – mass **lost** on heating [**(i)** – (mass **FA 5** – mass **FA 6**)]

mass of solid decomposition product = ... g
[1]

(c) Why was the lid put on while the crucible and its contents were cooled?

Lid reduces / stops absorption of water (vapour) by solid / residue / **FA 5** while cooling/ prevents reaction with carbon dioxide.

Do not accept: falling impurities/ loss of solid due to wind/oxidation in the air. [1]

(d) **(i)** Ah Beng carried out the experiment by heating 84.0 g of pure $NaHCO_3$ to constant mass, he obtained 53.0 g of the solid decomposition product.

Ah Hock then suggested the following equation for the thermal decomposition of sodium hydrogencarbonate.

$$NaHCO_3(s) \rightarrow NaOH(s) + CO_2(g)$$

Explain why Ah Hock's suggestion is **incorrect**. Show working in order to explain your answer.

[A_r: C; 12.0; Na; 23.0; H, 1.0; O: 16.0]

If Ah Hock is **correct**., (84 g) $NaHCO_3$ would give **40** g residue / NaOH
or
mole ratio of $NaHCO_3$: NaOH is 1:**1.3** (so not 1:1)
or
Answers could refer to mass lost / moles of CO_2

(ii) Add 1 cm depth of sulfuric acid into a test-tube.
Add some **FA 6** from the crucible to the acid in the test-tube.
Record all your observations.
Use your observations to identify the anion present in **FA 6**.

Any observation below + gas forms white ppt in limewater

Do NOT accept milky / chalky / cloudy

- fizzing / effervescence / bubbling
- solid dissolves / colourless solution **formed**

FA 6 contains carbonate ion / CO_3^{2-}

[2]

(iii) State the assumption you have made in **(ii)**.

Complete decomposition of $NaHCO_3$.
OR
There is no leftover $NaHCO_3$.

[1]

(iv) Steam is one of **three** products obtained when sodium hydrogencarbonate is thermally decomposed.

Use your answer in **(ii)** to complete and balance the equation for the thermal decomposition of sodium hydrogencarbonate. Include state symbols.

2$NaHCO_3(s) \rightarrow H_2O(g) + CO_2(g) +$ **$Na_2CO_3(s)$**

[1]

(v) State whether the balanced equation in **(iv)** agrees with Ah Beng's results.

Show working in order to explain your answer.

(From equation) 84.0 g $NaHCO_3$ should give 0.5 × 106 g residue (= 53.0 g)
And this agrees with Ah Beng's 53.0 g

[1]

[Total: 13]

4 Determination of the enthalpy change of a reaction, ΔH_r

FA 7 is 1.00 mol dm^{-3} sodium hydrogen carbonate, $NaHCO_3$
FA 8 is 2.00 mol dm^{-3} sodium hydroxide, NaOH
FA 9 is 2.00 mol dm^{-3} sulfuric acid, H_2SO_4

An acid-base neutralisation reaction involves reacting the two solutions, to produce water molecules. The equation for this neutralisation reaction is given below.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

You will follow the instructions to perform two experiments, **Experiment A** and **Experiment B**. Record your results in Tables 4.1 and 4.2.

Experiment A
Reaction between **FA 7**, $NaHCO_3$, and **FA 8**, NaOH.

Reaction 1 $NaHCO_3(aq) + NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$ $\Delta H_{reaction1}$

The molar enthalpy change for **reaction 1**, $\Delta H_{reaction1}$, is the enthalpy change when 1.00 mol of $NaHCO_3$ reacts completely with NaOH.

1 Using a measuring cylinder, transfer 30.0 cm^3 of **FA 7** into a Styrofoam cup. Place this cup inside a second Styrofoam cup, which is placed in a 250 cm^3 glass beaker. Place the lid on the cup.

2 Stir and measure the temperature of this **FA 7**, $T_{\mathbf{FA7}}$.

3 Using another measuring cylinder, measure 20.0 cm^3 of **FA 8.**

4 Stir and measure the temperature of this **FA 8**, $T_{\mathbf{FA8}}$.

5 Add **FA 8** from the measuring cylinder to the **FA 7** in the Styrofoam cup. Immediately replace the lid.

6 Using the thermometer, stir the mixture continuously until a maximum temperature is reached. Read and record this temperature T_{max}.

7 Calculate the weighted average initial temperature, $T_{average,}$ of **FA 7** and **FA 8** using the formula given below:

$$T_{average} = \frac{(V_{FA7} \times T_{FA7}) + (V_{FA8} \times T_{FA8})}{(V_{FA7} + V_{FA8})}$$

	Experiment A
$T_{\mathbf{FA7}}$ /°C	
$T_{\mathbf{FA8}}$ /°C	
$T_{average}$ /°C	
T_{max} /°C	
ΔT_{max} /°C	

Table 4.1

Experiment B

Reaction between **FA 7**, $NaHCO_3$, and **FA 9,** H_2SO_4.

Reaction 2 $NaHCO_3(aq) + ½H_2SO_4(aq) \rightarrow ½Na_2SO_4(aq) + H_2O(l) + CO_2(g)$ $\Delta H_{reaction2}$

The molar enthalpy change for **reaction 2**, $\Delta H_{reaction\ 2}$, is the enthalpy change when 1.00 mol of $NaHCO_3$ reacts completely with H_2SO_4.

1 Using a measuring cylinder, transfer 30.0 cm^3 of **FA 7** into a Styrofoam cup. Place this cup inside a second Styrofoam cup, which is placed in a 250 cm^3 glass beaker. Place the lid on the cup.

2 Stir and measure the temperature of this **FA 7**, T_{FA7}.

3 Using another measuring cylinder, measure 20.0 cm^3 of **FA 9.**

4 Stir and measure the temperature of this **FA 9**, T_{FA9}.

5 Add **slowly**, the **FA 9** from the measuring cylinder to the **FA 7** in the Styrofoam cup. Immediately replace the lid.

6 Using the thermometer, stir the mixture continuously until a maximum/minimum temperature is reached. Read and record this temperature T_{max}.

7 Calculate the weighted average initial temperature, $T_{average,}$ of **FA 7** and **FA 9** using the formula given below:

$$T_{average} = \frac{(V_{FA7} \times T_{FA7}) + (V_{FA9} \times T_{FA9})}{(V_{FA7} + V_{FA9})}$$

	Experiment B
T_{FA7} /°C	
T_{FA9} /°C	
$T_{average}$ /°C	
T_{max} /°C	
ΔT_{max} /°C	

Table 4.2

[2]

All temperature readings are recorded to the nearest 0.1 °C.
(for both Tables 4.1 & 4.2)
$\Delta T_{max/min}$ must be correctly calculated i.e. T_{max} – average initial temperature.
(for both Tables 4.1 & 4.2)

(a) For the purpose of calculations, you should assume that the mixture has a density of 1.00 g cm^{-3} and specific heat capacity, c, of 4.18 J g^{-1} K^{-1}.

(i) Use your results from **Table 4.1** to calculate a value for the molar enthalpy change for **reaction 1**, $\Delta H_{reaction1}$.

$q = (20.0 + 30.0)(1.00) \times 4.18 \times \Delta T_{max}$

$n(NaHCO_3) = 1.00 \times 30.0/1000 = 0.03$ mol

$\Delta H_{reaction1} = -q / n(NaHCO_3)$ (this reaction is exothermic)

$\Delta H_{reaction1}$ = ...

[2]

(ii) Use your results from **Table 4.2** to calculate a value for the molar enthalpy change for **reaction 2**, $\Delta H_{reaction2}$.

$q = (20.0 + 30.0)(1.00) \times 4.18 \times \Delta T_{min}$

$n(NaHCO_3) = 1.00 \times 30.0/1000 = 0.03$ mol

$\Delta H_{reaction1} = +q / n(NaHCO_3)$ (this reaction is endothermic)

$\Delta H_{reaction2}$ = ...

[2]

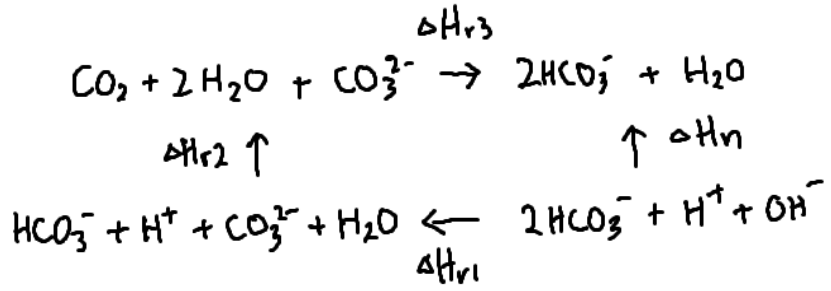
(b) Ionic equations for neutralisation, **reaction 1**, and **reaction 2** are shown below.

$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$	$\Delta H_{neu} = -57.1$ kJ mol^{-1}
$HCO_3^-(aq) + OH^-(aq) \longrightarrow CO_3^{2-}(aq) + H_2O(l)$	$\Delta H_{reaction1}$
$HCO_3^-(aq) + H^+(aq) \longrightarrow H_2O(l) + CO_2(g)$	$\Delta H_{reaction2}$

Carbon dioxide reacts with solutions of carbonate ions according to the following equation.

reaction 3 $CO_2(g) + CO_3^{2-}(aq) + H_2O(l) \longrightarrow 2HCO_3^-(aq)$ $\Delta H_{reaction3}$

Using your calculated answers in **(a),** together with the given value of enthalpy change of neutralisation, ΔH_{neu}, construct an energy cycle to determine a value for the enthalpy change for this reaction, $\Delta H_{reaction3}$.



$\Delta H_{reaction3}$ = (– **(a)(ii)**) + (– **(a)(i)**) + (–57.1)

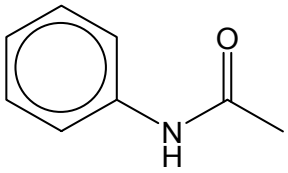
$\Delta H_{reaction3}$ = ...

[3]

[Total: 9]

5 Planning

The structure of *N*-phenylethanamide is shown below:



N-phenylethanamide, $C_6H_5NHCOCH_3$

The preparation of *N*-phenylethanamide from phenylamine is termed acylation. Ethanoic anhydride, $(CH_3CO)_2O$ is commonly used as an acylating agent as its low reactivity relative to ethanoyl chloride allows the reaction rate to be more easily controlled.

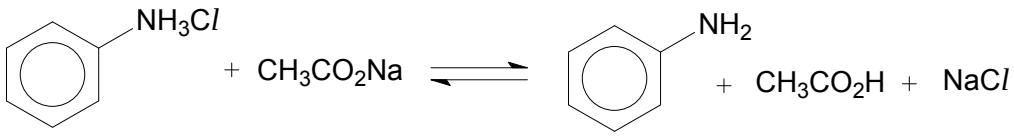
Phenylamine is most conveniently used in the form of the salt phenylammonium chloride, $C_6H_5NH_3Cl$.

The reaction is performed in **two** stages:

Stage **1:**

Phenylamine can be prepared from phenylammonium chloride and sodium ethanoate in an aqueous medium under room temperature.

The reaction is as shown below:



phenylammonium chloride phenylamine

The reaction mixture needs to be continuously stirred for 3 minutes.

Stage **2:**

Phenylamine then reacts with ethanoic anhydride to form solid *N*-phenylethanamide together with ethanoic acid. The addition of ethanoic anhydride to the reaction mixture may be violent.

This mixture is heated under reflux for about 15 minutes, before the crude solid product is removed by filtration and purified by recrystallisation from hot water. The melting point of *N*-phenylethanamide is 114 ^{0}C.

(a) (i) State the type of reaction taking place in stage **1**.

Acid base reaction/ neutralisation

[1]

(ii) Write an equation for the reaction of phenylamine with ethanoic anhydride in stage **2.**

$C_6H_5NH_2 + (CH_3CO)_2O \rightarrow C_6H_5NHCOCH_3 + CH_3CO_2H$.

[1]

(b) The above reaction between phenylamine and ethanoic anhydride in stage **2** gives an 80% yield.

Assuming phenylammonium chloride is completely converted to phenylamine in stage **1**, determine the mass of phenylammonium chloride required to prepare 2 g of *N*-phenylethanamide.
(A_r: C: 12.0, H: 1.0, N: 14.0, O: 16.0, C*l*: 35.5)

Amount of *N*-phenylethanamide = 2/ 135= 0.01481 mol

Amount of phenylamine = 0.01481 mol

Mass of phenylammonium chloride
= 0.01481 x (72 + 8 + 14 + 35.5) x 100/80
= 2.398 g

[2]

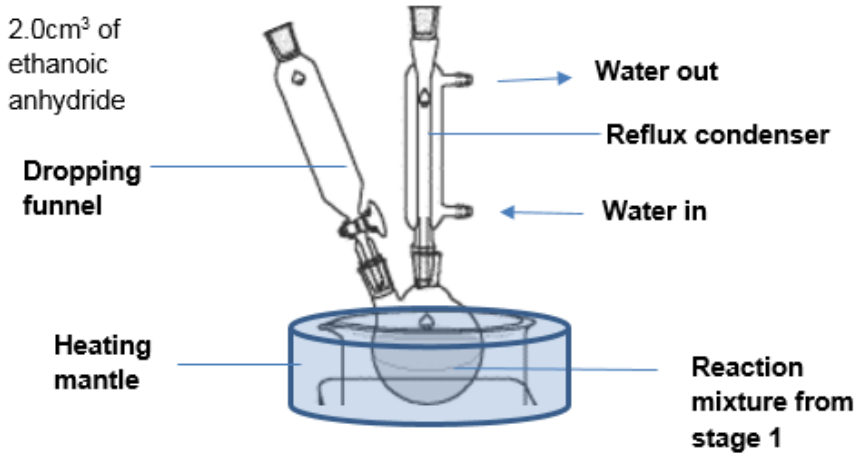
(c) Write a full description of the procedure to carry out stages **1** and **2**.
You do not need to describe the recrystallization process to obtain a pure sample of *N*-phenylethanamide *in stage **2**.*

You may assume that you are provided with :

- 30 cm^3 of water for use as solvent
- 6.0 g of hydrated sodium ethanoate
- 2.0 cm^3 of ethanoic anhydride
- Apparatus normally found in a school or college laboratory

Your plan should include details of:

(i) appropriate quantities of reactants
(ii) appropriate choice of apparatus
(iii) drawing of reflux set-up



1. **Weight out** accurately about **2.40 g** of phenylammonium chloride in a clean and dry weighing bottle.

2. **Measure 30 cm^3 of water** using a 100 cm^3 measuring cylinder.

3. **Dissolve/Add** this completely in 30 cm^3 of water in a clean and dry 250 cm^3 round bottom flask.

4. Weigh out **6.0 g** of hydrated sodium ethanoate in a weighing bottle.

5. Add sodium ethanoate to the solution in (3) and **stir the mixture for three minutes** using a magnetic stirrer.

6. Measure **2.0 cm^3 of ethanoic anhydride** using a 10 cm^3 measuring cylinder.

7. Introduce **2.0 cm^3** of ethanoic anhydride slowly using a **dropping funnel/dropwise using dropper /(using ice bath for the reaction)** and connect the reaction mixture to the reflux condenser as shown in the set up below inside a fumehood.

8. Carry out the **reflux for 15minutes**.

9. **Filter off** the crude product and wash it with a little cold water

(d) Suggest a method to check the purity of *N*-phenylethanamide obtained.

Determine the melting point and verify it against literature data/ value of 114 ^{o}C.
If the melting point is within $\pm$ 2 ^{o}C of literature data/ 114 ^{o}C, it is pure

[Total: 10]