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# Chemistry
Unit A21: Practical Manual

# Questions Solutions

# Practical 1.1

Investigating the effect of changing concentration on the rate of a chemical reaction (spec ref 4.3.7)

1. The graph is a straight line graph through the origin, indicating the order of reaction with respect to sodium thiosulfate is first order.

# Practical 1.2

Investigating the effect of changing concentration on the rate of a chemical reaction (spec ref 4.3.7)

1. The graph is a straight line which indicates the order of reaction with respect to iodine is zero order.

2. The sodium hydrogencarbonate quenches/stops the reaction in each sample by reacting with acid.

   $2NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2CO_2 + 2H_2O$

3. Using starch enables a sharp endpoint to be observed in each titration.

# Practical 1.3

Investigating the effect of changing concentration on the rate of a chemical reaction (spec ref 4.3.7)

1. The graph is a straight line graph through the origin, indicating the order of reaction with respect to iodide is first order.

2. The maximum amount of iodine produced in each case is $8 \times 10^{-5}$ mol. The amount of thiosulfate added is $2 \times 10^{-5}$ mol, which will react with $1 \times 10^{-5}$ mol of iodine. The percentage of the reaction studied is $(1 \times 10^{-5})/(8 \times 10^{-5}) \times 100 = 12.5\%$

3. The rate constant can be determined by calculating the gradient of the straight line graph. The units are $mol^{-1} dm^3 s^{-1}$.
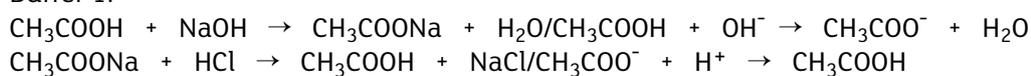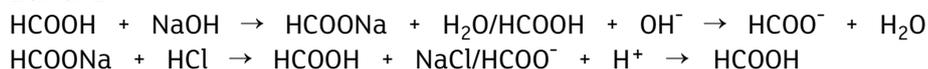
# Practical 2.1
## Making buffer solutions and investigating their pH values (spec ref 4.5.5)

In each buffer system, the amount of acid/alkali added is small relative to the amount of the buffer system. When acid is added, it is removed by the alkali or conjugate base in the buffer system. When alkali is added, it is removed by the acid/conjugate acid of the buffer system.
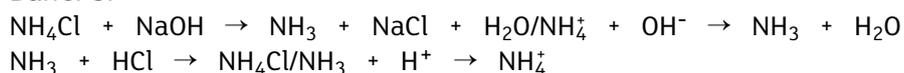
Buffer 1:
$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O/CH_3COOH + OH^- \rightarrow CH_3COO^- + H_2O$
$CH_3COONa + HCl \rightarrow CH_3COOH + NaCl/CH_3COO^- + H^+ \rightarrow CH_3COOH$

Buffer 2:
$HCOOH + NaOH \rightarrow HCOONa + H_2O/HCOOH + OH^- \rightarrow HCOO^- + H_2O$
$HCOONa + HCl \rightarrow HCOOH + NaCl/HCOO^- + H^+ \rightarrow HCOOH$

Buffer 3:
$NH_4Cl + NaOH \rightarrow NH_3 + NaCl + H_2O/NH_4^+ + OH^- \rightarrow NH_3 + H_2O$
$NH_3 + HCl \rightarrow NH_4Cl/NH_3 + H^+ \rightarrow NH_4^+$

2. Molar mass of potassium methanoate = 1 + 12 + (2 × 16) + 39 = 84
   Number of moles = 1.42/84 = 0.0169 mol
   [HCOOK] = 0.0169 ÷ 0.0500 = 0.338 mol dm$^{-3}$
   $K_a = [H^+(aq)] [HCOO^-(aq)]/ [HCOOH(aq)] = 1.78 \times 10^{-4}$ mol dm$^{-3}$
   $[H^+] = K_a$[weak acid]/[salt] = $(1.78 \times 10^{-4})(0.111)/(0.338) = 5.85 \times 10^{-5}$ mol dm$^{-3}$
   pH = - log $(5.58 \times 10^{-5})$ mol dm$^{-3}$ = 4.23

3. $[H^+] = 10^{-pH} = 10^{-4.38} = 4.17 \times 10^{-5}$ mol dm$^{-3}$
   $K_a = [H^+(aq)] [CH_3COO^-(aq)]/[CH_3COOH(aq)] = 1.8 \times 10^{-5}$ mol dm$^{-3}$
   [salt] = $K_a$ [weak acid]/$[H^+]$ = $(1.8 \times 10^{-5})(1.0)/(4.17 \times 10^{-5})$ = 0.432 mol dm$^{-3}$
   Molar mass of $CH_3COONa$ = 12 + 3 + 12 + (2 × 16) + 23 = 82
   Mass of sodium ethanoate required for 1 dm$^3$ = 0.432 × 82 = 35 g
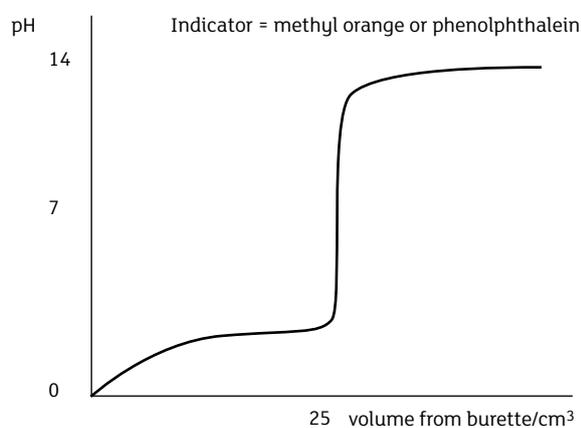   Mass for 100cm$^3$ = 3.5 g

# Practical 2.2

## Determining the shape of a titration curve
## (spec ref: 4.5.6)

1. Sketch each of the following pH curves on the grids shown and name a suitable indicator.
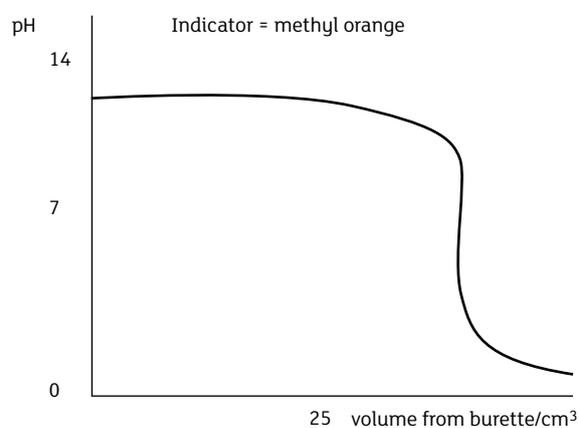
**(1)** Flask 25 cm$^3$ 0.10 mol dm$^{-3}$ $H_2SO_4$

Burette 50 cm$^3$ 0.20 mol dm$^{-3}$ NaOH

pH

Indicator = methyl orange or phenolphthalein

14

7

0

25   volume from burette/cm$^3$

**(2)** Flask 20 cm$^3$ 1.00 mol dm$^{-3}$ $NH_3$

Burette 50 cm$^3$ 0.50 mol dm$^{-3}$ HCl

pH

Indicator = methyl orange

14

7

0

25   volume from burette/cm$^3$

**(3)** Flask 20 cm$^3$ 0.20 mol dm$^{-3}$ $CH_3COOH$

Burette 50 cm$^3$ 0.20 mol dm$^{-3}$ NaOH

pH

Indicator = phenolphthalein

14

7

0

25   volume from burette/cm$^3$
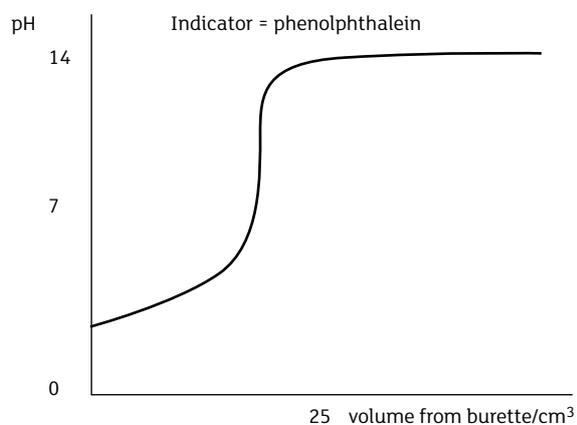
**(4)** Flask 25 cm$^3$ 0.50 mol dm$^{-3}$ $NH_3$

Burette 50 cm$^3$ 1.00 mol dm$^{-3}$ methanoic acid

pH

Indicator = none suitable

14

7

0

25   volume from burette/cm$^3$
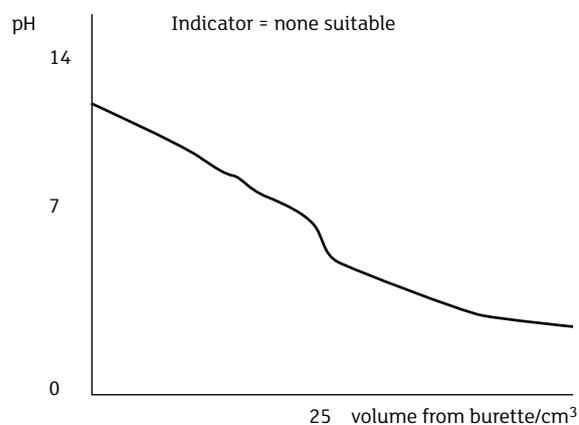
# Practical 2.3
## Determining the pH of salts (spec ref: 4.5.8)

1. Sodium chloride: Ions present are those of a strong acid and base so will not react with hydrogen or hydroxide ions, resulting in a neutral solution.
   Sodium ethanoate: Ions present are those of a strong base and weak acid. Ethanoate ions react with hydrogen ions, resulting in an alkaline solution as $[H^+] < [OH^-]$
   $$CH_3COO^- + H^+ \rightarrow CH_3COOH$$

   Ammonium chloride: Ions present are those of a weak base and strong acid. Ammonium ions react with hydroxide ions, resulting in an acidic solution as $[H^+] > [OH^-]$.
   $$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

   Ammonium ethanoate: Ions present are those of a weak acid and weak base. Ammonium ions react with hydroxide ions and ethanoate ions react with hydrogen ions, resulting in a solution which is approximately neutral.
   $$CH_3COO^- + H^+ \rightarrow CH_3COOH$$
   $$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

2. Potassium cyanide: pH greater than 7. Ions present are those of a strong base and weak acid. Cyanide ions react with hydrogen ions, resulting in an alkaline solution as $[H^+] < [OH^-]$.

   Sodium methanoate: pH greater than 7. Ions present are those of a strong base and weak acid. Methanote ions react with hydrogen ions, resulting in an alkaline solution as $[H^+] < [OH^-]$.

   Methylammonium chloride: pH less than 7. Ions present are those of a strong acid and weak base. Methylammonium ions react with hydroxide ions, resulting in an acidic solution as $[H^+] > [OH^-]$.

   Ammonium sulfate: pH less than 7. Ions present are those of a strong acid and weak base. Ammonium ions react with hydroxide ions, resulting in an acidic solution as $[H^+] > [OH^-]$.

   Potassium sulfate: pH equal to 7. Ions present are those of a strong acid and strong base, so will not react with hydrogen ions or hydroxide ions.

   Sodium carbonate: pH greater than 7. Ions present are those of a strong base and weak acid. Carbonate ions react with hydrogen ions, resulting in an alkaline solution as $[H^+] < [OH^-]$.

   Sodium fluoride: pH equal to 7. Ions present are those of a strong acid and strong base, so will not react with hydrogen ions or hydroxide ions.

# Practical 3.1

Preparation of 2,4-dinitrophenylhydrazones
(spec ref: 4.7.8)

1. A pure sample will have a sharp melting point over a small range of temperature (+/- 1 °C).
2. Melting point determined by carbonyl compound used.
3. Equation in the form:
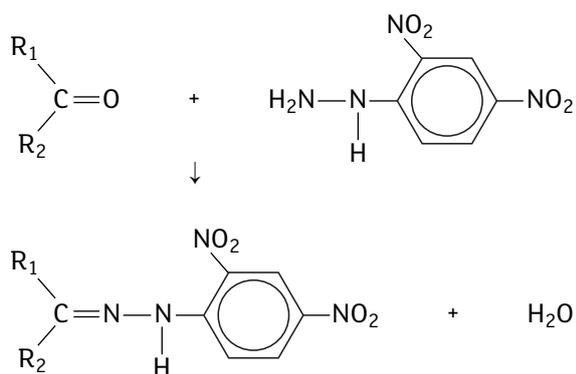


# Practical 3.2

Using Fehling's solution and Tollens' reagent to distinguish between aldehydes and ketones
(spec ref 4.7.9)

1. Orange to green.

2. $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

3. A yellow/orange solid will be observed with an aldehyde and a ketone.

# Practical 4.1
Preparing a carboxylic acid from an alcohol
(spec ref 4.8.3)

1. The addition of concentrated sulfuric acid is exothermic.

2. Anti-bumping granules promote smooth boiling.

3. A vinegary smell is observed, universal indicator paper turns red and effervescence is observed with sodium carbonate.

4. The reaction mixture should be set up for distillation only and not refluxed. The receiving flask should be placed in an iced water bath so that the ethanal condenses and is collected.
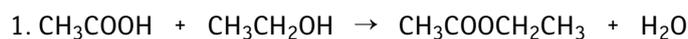
# Practical 4.2
Carrying out test tube reactions of a carboxylic acid with sodium carbonate, sodium hydroxide and aqueous ammonia and measuring pH changes (spec ref 4.8.5)

1.
$$2CH_3COOH + Na_2CO_3 \rightarrow 2CH_3COONa + H_2O + CO_2$$
$$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$$
$$CH_3COOH + NH_3 \rightarrow CH_3COONH_4$$

2. $CO_2 + Ca(OH)_2 \rightarrow CaCO_3 + H_2O$

3. Add a spatula measure of phosphorus(V) chloride to the suspected carboxylic acid; effervescence is observed along with steamy/white fumes which turn damp universal indicator paper red.
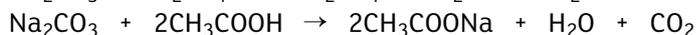
# Practical 5.1

## Preparing a liquid ester from a carboxylic acid and an alcohol (spec ref 4.9.5)

1. $CH_3COOH \ + \ CH_3CH_2OH \ \rightarrow \ CH_3COOCH_2CH_3 \ + \ H_2O$

2. Catalyst and dehydrating agent.

3. It removes the concentrated sulfuric acid and can also react with unreacted ethanoic acid.
   $Na_2CO_3 \ + \ H_2SO_4 \ \rightarrow \ Na_2SO_4 \ + \ H_2O \ + \ CO_2$
   $Na_2CO_3 \ + \ 2CH_3COOH \ \rightarrow \ 2CH_3COONa \ + \ H_2O \ + \ CO_2$

4. Add a few drops of water; the layer which increases in volume is the aqueous layer.

5. $2CH_3CH_2OH \ \rightarrow \ (CH_3CH_2)_2O \ + \ H_2O$
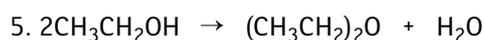
# Practical 6.1

## Preparation of methyl 3-nitrobenzoate (spec ref 4.10.5)

1. The yield will probably be no more than 50%. Reasons include: reaction does not go to completion, side reactions, loss in transfer.

2. The melting point of the pure product is 78–80 °C. A pure sample will have a sharp melting point within this range. An impure sample will begin melting lower than this temperature and melt over a broader range.

3. $HNO_3 \ + \ 2H_2SO_4 \ \rightarrow \ NO_2^+ \ + \ 2HSO_4^- \ + \ H_3O^+$

4.
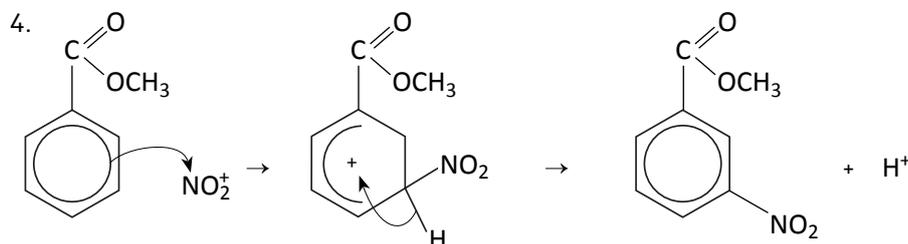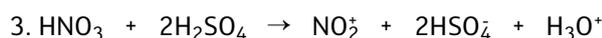

5. The temperature is controlled to prevent further nitration of the benzene ring.