



**Coimisiún na Scrúduithe Stáit  
State Examinations Commission**

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**LEAVING CERTIFICATE EXAMINATION, 2006**

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**CHEMISTRY - HIGHER LEVEL**

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TUESDAY, 20 JUNE - AFTERNOON 2.00 to 5.00

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**400 MARKS**

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Answer **eight** questions in all

These **must** include at least **two** questions from **Section A**

All questions carry equal marks (50)

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**Information**

Relative atomic masses: H = 1, C = 12, O = 16, Na = 23, Ca = 40.

Universal gas constant,  $R = 8.3 \text{ J K}^{-1} \text{ mol}^{-1}$

Molar volume at s.t.p. = 22.4 litres

Avogadro constant =  $6 \times 10^{23} \text{ mol}^{-1}$

## Section A

Answer at least two questions from this section [see page 1 for full instructions].

1. An experiment was carried out to determine the percentage water of crystallisation and the degree of water of crystallisation,  $x$ , in a sample of hydrated sodium carbonate crystals ( $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ ). An 8.20 g sample of the crystals was weighed accurately on a clock glass and then made up to 500 cm<sup>3</sup> of solution in a volumetric flask. A pipette was used to transfer 25.0 cm<sup>3</sup> portions of this solution to a conical flask. A previously standardised 0.11 M hydrochloric acid (HCl) solution was used to titrate each sample. A number of accurate titrations were carried out. The average volume of hydrochloric acid solution required in these titrations was 26.05 cm<sup>3</sup>.

The titration reaction is described by the equation:

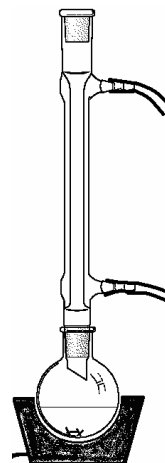
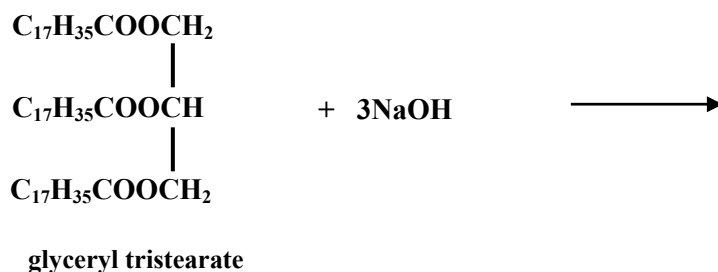


- (a) Identify a primary standard reagent which could have been used to standardise the hydrochloric acid solution. (5)
- (b) Name a suitable indicator for the titration and state the colour change observed in the conical flask at the end point. Explain why not more than 1 – 2 drops of indicator should be used. (12)
- (c) (i) Describe the correct procedure for rinsing the burette before filling it with the solution it is to deliver.  
(ii) Why is it important to fill the part below the tap of the burette? (12)
- (d) From the titration figures, calculate the concentration of sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) in the solution in  
(i) moles per litre,  
(ii) grams per litre. (9)
- (e) Calculate the percentage water of crystallisation present in the crystals and the value of  $x$ , the degree of hydration of the crystals. (12)
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2. A sample of soap was prepared in the laboratory by refluxing a mixture of approximately 5 g of animal fat, 2 g of sodium hydroxide pellets (an excess) and 25 cm<sup>3</sup> of ethanol in an apparatus like that drawn on the right.

(a) Why was the reaction mixture refluxed? Name the *type* of reaction which occurs during the reflux stage of the preparation. (8)

(b) Complete and balance the equation below for the reaction between glyceryl tristearate, an animal fat, and sodium hydroxide. (9)



(c) What is the purpose of the ethanol? Why is it desirable to remove the ethanol after reflux? Describe with the aid of a labelled diagram how you would remove the ethanol after the reflux stage of the experiment. (12)

(d) Describe how a pure sample of soap was obtained from the reaction mixture. (9)

(e) At the end of the experiment, what is the location  
 (i) of the second product of the reaction,  
 (ii) of the excess sodium hydroxide? (6)

(f) What would you observe, upon shaking, if a little of the soap prepared in this experiment is added to  
 (i) a test tube containing deionised water,  
 (ii) a test tube containing mineral water from a limestone region? (6)

3. A number of tests were carried out on a sample of swimming pool water to test its quality.

(a) A colorimetric experiment was used to estimate the concentration of free chlorine in the sample. What is the general principle of all colorimetric experiments? (8)

(b) Identify a suitable reagent to test for free chlorine in swimming pool water and state the colour which develops when this reagent reacts with free chlorine. (6)

(c) Describe briefly how you would estimate the concentration of free chlorine in a sample using either a comparator or a colorimeter. (12)

(d) Give the name or formula of a *free chlorine* species in the swimming pool water. Give a reason why the concentration of free chlorine in treated drinking water is usually between 0.2 - 0.5 p.p.m. whereas in swimming pool water it should be between 1 - 5 p.p.m. (9)

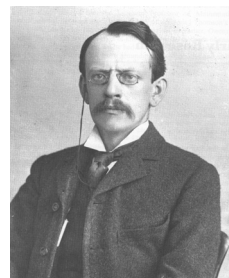
(e) When 1200 cm<sup>3</sup> of swimming pool water was filtered, the mass of the filter paper, upon drying, had increased by 0.78 g. When 250 cm<sup>3</sup> of the filtered water was evaporated to dryness the mass of the residue obtained was 0.32 g. Calculate the concentration in p.p.m.  
 (i) of suspended solids,  
 (ii) of dissolved solids. (15)

## Section B

[See page 1 for instructions regarding the number of questions to be answered.]

4. Answer **eight** of the following items (a), (b), (c), etc. (50)

- (a) Write the electron configuration (*s, p, etc.*) of a chromium atom in its ground state.
- (b) Name the scientist, shown in the photograph, who identified cathode rays as subatomic particles.
- (c) Give **two** differences between a nuclear reaction and a chemical reaction.
- (d) Calculate the percentage carbon, by mass, in methylbenzene.
- (e) What is (i) the conjugate acid and (ii) the conjugate base of  $\text{H}_2\text{O}$ ?
- (f) What contribution did Newlands make to the systematic arrangement of the elements known to him?
- (g) What observation is made when a sample of ethanal is heated with Fehling's reagent?
- (h) The concentration of an aqueous solution of sodium hydroxide ( $\text{NaOH}$ ) is 0.2 g per litre. Calculate its pH.
- (i) Under what circumstances can ionic compounds conduct electricity?
- (j) Which class of organic compound is responsible for the odour associated with fruits such as apples, oranges, pears, bananas and strawberries?
- (k) Answer part A or part B.



**A** State **two** uses of nitrogen gas based on its chemical stability.

*or*

**B** Name **two** metals, one a main group metal, the other a transition element, both of which are protected from further corrosion by the oxide layer which forms on their surfaces.

5. (a) (i) Describe how you would carry out a flame test on a sample of potassium chloride. (8)
- (ii) Why do different elements have unique atomic spectra? (6)
- (iii) What instrumental technique is based on the fact that each element has unique atomic spectra? (3)

Bohr's model of the atom explained the existence of energy levels on the basis of atomic spectra. Bohr's theory was later modified to incorporate the idea of *orbitals* in recognition of the wave nature of the electron and Heisenberg's uncertainty principle.

- (iv) Define *atomic orbital*. (6)
- (v) What does Heisenberg's uncertainty principle say about an electron in an atom? (6)
- (b) (i) Define *electronegativity*. (6)
- (ii) Explain why there is a general increase in electronegativity values across the periods in the periodic table of the elements. (6)
- (iii) Explain, in terms of the structures of the atoms, the trend in reactivity down Group I (the alkali metal group) of the periodic table. (9)

6. (a) The table shows the octane numbers of four hydrocarbons.

Name	Formula	Octane No.
hexane	C <sub>6</sub> H <sub>14</sub>	25
cyclohexane	C <sub>6</sub> H <sub>12</sub>	83
benzene	C <sub>6</sub> H <sub>6</sub>	100
2,2,4-trimethylpentane	C <sub>8</sub> H <sub>18</sub>	100

(i) What is meant by the octane number of a fuel? (8)

(ii) Hexane has the lowest octane number of the four compounds listed. What structural feature of the molecule contributes to this? (3)

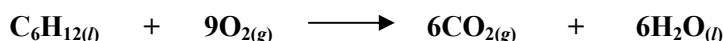
(iii) In the case of each of the other three compounds, identify the structural feature of its molecules which contributes to it having a high octane number. (9)

(iv) Name the process carried out in an oil refinery that converts hexane to compounds such as cyclohexane and benzene. Why is the use of benzene in petrol strictly controlled? (6)

(b) (i) Give **two** reasons why oxygenates such as MTBE are added to petrol.

(ii) Give **two** reasons why the addition of lead to petrol has been discontinued. (12)

(c) The combustion of cyclohexane may be described by the following balanced equation:



Given that the heats of formation of cyclohexane, carbon dioxide and water are  $-156$ ,  $-394$  and  $-286 \text{ kJ mol}^{-1}$ , respectively, calculate the heat of combustion of cyclohexane. (12)

7. (a) Define the *activation energy* of a chemical reaction. (5)

(b) Give **two** reasons why the rate of a chemical reaction increases as the temperature rises. Which of these is the more significant? Why? (12)

(c) Describe how you could investigate the effect of temperature on the rate of the reaction between a 0.1 M sodium thiosulfate solution and a 2 M hydrochloric acid solution. (12)  
The reaction is described by the following balanced equation.



(d) When silver nitrate and sodium chloride solutions are mixed a precipitate appears immediately. Explain the speed of this reaction compared to the slower reaction when solutions of sodium thiosulfate and hydrochloric acid are mixed. (6)

(e) What type of catalysis occurs in the catalytic converter of a modern car? Give the names *or* formulas of **two** substances entering a car's catalytic converter and the names *or* formulas of the substances to which they are converted in the interior of the catalytic converter. (15)

8. (a) (i) What is *hard water*? (5)

(ii) A supply of hard water is treated for domestic use by ion-exchange. You may assume that all the hardness is due to **Ca(HCO<sub>3</sub>)<sub>2</sub>**. Explain in words or using a balanced equation how a cation exchange resin, represented by **RNa**, softens this water supply. (6)

(iii) In the treatment of water for drinking, what is meant by the term *flocculation*? Name a flocculating agent. (9)

(iv) What substance is added to water to adjust the pH if the water is too acidic? Why is it undesirable to have the pH of drinking water below 6? (6)

(b) (i) Explain how an acid-base indicator, which is itself a weak acid, and may be represented by **HX**, functions. (9)

(ii) Draw a clearly labelled diagram of the titration curve you would expect to obtain when 50 cm<sup>3</sup> of a 0.1 M sodium hydroxide (**NaOH**) solution is added slowly to 25 cm<sup>3</sup> of a 0.1 M ethanoic acid (**CH<sub>3</sub>COOH**) solution. (9)

(iii) Explain with reference to your diagram why phenolphthalein is a suitable indicator for a titration of sodium hydroxide with ethanoic acid. (6)

9. The alkenes are a homologous series. Ethene ( $C_2H_4$ ) is the first member of the series.

(a) What is meant by a *homologous series*? (5)

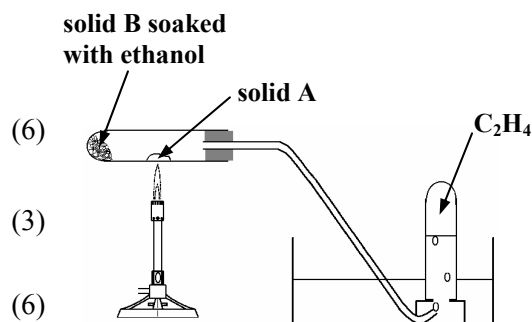
(b) Ethene may be made in a school laboratory using the arrangement of apparatus drawn on the right.

(i) Give the name and formula of the solid **A** which is heated using the Bunsen burner.

(ii) Identify the solid **B** which is used to keep the ethanol at the end of the test tube.

(iii) What precaution should be observed when heating is stopped? Why is this necessary?

(iv) Give **one** major use of ethene gas.



(c) Describe the mechanism for the bromination of ethene. (9)

State and explain **one** piece of experimental evidence to support this mechanism. (6)

(d) Draw the structures and give the systematic (IUPAC) names for **two** alkene isomers of molecular formula  $C_4H_8$ . (12)

10. Answer any **two** of the parts (a), (b) and (c) (2 × 25)

(a) (i) What are *isotopes*? (4)

(ii) Define *relative atomic mass*,  $A_r$ . (6)

(iii) What is the principle on which the mass spectrometer is based? (9)

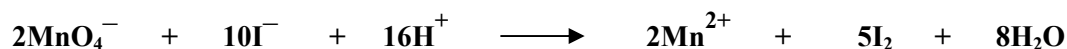
(iv) Calculate the relative atomic mass of a sample of lithium, given that a mass spectrometer shows that it consists of 7.4 % of  ${}^6\text{Li}$  and 92.6 % of  ${}^7\text{Li}$ . (6)

(b) Define *oxidation* in terms of change in oxidation number. (4)

What is the oxidation number of (i) chlorine in  $\text{NaClO}$  and (ii) nitrogen in  $\text{NO}_3^-$ ? (6)

State and explain the oxidation number of oxygen in the compound  $\text{OF}_2$ . (6)

Using oxidation numbers or otherwise, identify the reducing agent in the reaction between acidified potassium manganate(VII) and potassium iodide solutions represented by the balanced equation below. Use your knowledge of the colours of the reactants and products to predict the colour change you would expect to see if you carried out this reaction. (9)

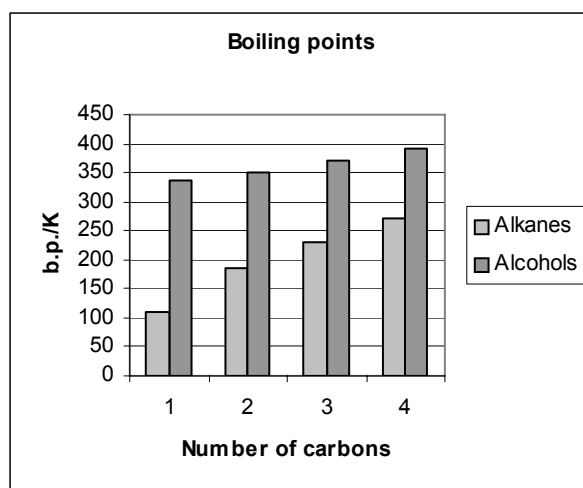


(c) The chart compares the boiling points of alkanes and primary alcohols containing from one to four carbon atoms.

(i) Give **two** reasons why each of these alcohols has a higher boiling point than the corresponding alkane. (7)

(ii) Explain why the difference in boiling points between methane and methanol is 226.5 K while the difference in boiling points between butane and butanol is only 119 K. (6)

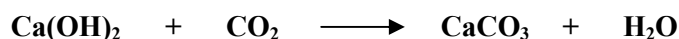
(iii) Describe, in general terms, the solubilities of methane, methanol, butane and butanol in water. (12)



11. Answer any **two** of the parts (a), (b) and (c)

(2 × 25)

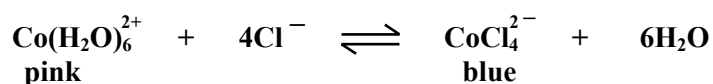
- (a) (i) What is an *ideal gas*? (4)
- (ii) Give **one** reason why a real gas like carbon dioxide deviates from ideal behaviour. (3)
- (iii) Assuming ideal behaviour, how many moles of carbon dioxide are present in 720 cm<sup>3</sup> of the gas at 10 °C and a pressure of 1 × 10<sup>5</sup> Pa? Give your answer correct to one significant figure. (9)
- (iv) How many molecules of carbon dioxide are present in this quantity of carbon dioxide? (3)
- (v) The reaction between carbon dioxide and limewater is represented by the following balanced equation.



What mass of calcium hydroxide is required to react completely with the quantity of carbon dioxide gas given in (iii) above? (6)

(b) State *Le Châtelier's principle*. (7)

The following equilibrium is set up in solution by dissolving cobalt(II) chloride crystals in water to form the pink species  $\text{Co(H}_2\text{O)}_6^{2+}$  and then adding concentrated hydrochloric acid until the solution becomes blue.



- (i) When the solution becomes blue, has reaction ceased? Explain. (6)
- (ii) The forward reaction is endothermic. State and explain the colour change observed on cooling the reaction mixture. (6)
- (iii) Other than heating, mention **one** way of reversing the change caused by cooling the reaction mixture. (6)

(c) Answer part **A** or part **B**

**A**

Select **one** of the manufacturing processes below and answer the questions which follow:

**ammonia manufacture**

**nitric acid manufacture**

**magnesium oxide manufacture**

- (i) What are the raw materials for the manufacturing process you have chosen? Describe how the raw materials are treated before they become the feedstock for the manufacturing process. (12)
- (ii) Name **one** product of the process you have chosen, which, if discharged, could cause pollution. (3)
- (iii) State the most important use of the *main* product of the process you have chosen. What makes this product particularly suitable for this use? (10)

**or**

**B**

A blast furnace may be used in the extraction of iron from iron ore.

- (i) What materials must be added to a blast furnace in operation? (12)
- (ii) Name the principal reducing agent in the blast furnace and write a balanced equation for its reaction with haematite ( $\text{Fe}_2\text{O}_3$ ). (9)
- (iii) Why is the pig iron produced in a blast furnace further processed into steel? (4)

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