**Atomic Structure – Periodic Table**

**2002 Question 5**

(a) Define first ionisation energy. **(8)**

(b) Account fully for the trends in first ionisation energies of elements across the second period of the periodic table (i.e. Li to Ne). **(15)**

(c) Account for the trend in first ionisation energies of the elements going down Group II of the periodic table, i.e. the alkaline-earth metals. **(6)**

The approximate values for the first eight ionisation energies of magnesium are given in the following table.



(d) Explain why there is an increase in these ionisation energy values. **(9)**

(e) Account for the dramatic increase in ionisation energy going from the second to the third ionisation. Between which two ionisations would you expect the next dramatic increase to occur if the data for further ionisation energies of magnesium were examined? Give a reason for your answer. **(12)**

**2002 Question10. (b)**

(i) What is the colour of the light associated with the line emission spectrum of sodium? **(4)**

(ii) Explain how line emission spectra occur. **(12)**

(iii) What evidence do line emission spectra provide for the existence of energy levels in atoms? **(6)**

(iv) Why is it possible for line emission spectra to be used to distinguish between different elements?**(3)**

**2002 Question 11. (b)**

What are alpha-particles (α-particles)? **(7)**

Describe the experiment carried out by Rutherford and his co-workers that led to the discovery of the nucleus. Explain how Rutherford interpreted the results of this experiment to conclude that the atom has a nucleus. **(18)**

**2003 Question 5.**

(*a*) Define

(*i*) *energy level*

(*ii*) *atomic orbital*.

(*iii*) Write the electronic configuration (*s*, *p*, etc.) of nitrogen. **(8)**

(*iv*) Describe how the electrons are arranged in the orbitals of the highest occupied sub-level of a nitrogen atom in its ground state. **(6)**

(*b*) Define *electronegativity*. **(6)**

(*i*) Describe using dot and cross diagrams the bonding in the water molecule**. (9)**

(*ii*) What is the shape of the water molecule? Which of the following angles, 104°, 107°, 109°, 120° or 180° would you expect to be closest to the bond angle in the water molecule? Explain your answer. **(12)**

(*c*) The diagram on the right shows a thin stream of water flowing from a burette. What would you observe if a charged rod was brought close to the thin stream of water? Explain your answer. **(9**)

**2003 Question 10.(*c*)**

A student was given samples of the following salts:

**sodium sulfate (Na2SO4) sodium sulfite (Na2SO3) potassium sulfate (K2SO4)**

(*i*) What test could be carried out to distinguish between the sodium salts and the potassium salt? **(4)**

What observation would you make in this test? **(6)**

(*ii*) Describe the test which could be carried out to distinguish between the sulfate salts and the sulfite salt. **(15)**

 **2003 Question 11. (*b*)**

 The diagram shows a sketch of the trend in the first ionisation energies for the elements 3 to 10 in the periodic table.

(*i*) Account for the general increase in ionisation energies across these elements. **(7)**

(*ii*) Explain why the ionisation energies of element number 4 and 7 are exceptionally high relative to the general trend. **(12)**

(*iii*) How does the definition of second ionisation energy differ from that of first ionisation energy? **(6)**

 **2004 Question 5**

(*a*) Write the electron configuration (*s, p,* etc.) of the nitrogen atom. **(5)**

Show, using dot and cross diagrams, the bond formation in a nitrogen molecule.

Describe the bonding in the nitrogen molecule in terms of sigma (σ) and pi (π) bonding. **(9)**

What type of intermolecular forces would you expect to find in nitrogen gas? Explain your answer. **(6)**

(*b*) Define *first ionisation energy*. **(9)**

There is a general increase in first ionisation energy across a period of the periodic table. State the **two** principal reasons for this trend. **(6)**

The table shows the first and second ionisation energies of nitrogen, oxygen, neon and sodium. Account for the decrease in first ionisation energy between nitrogen and oxygen.

Explain why the second ionisation energy of sodium is significantly (about nine times) higher than the first while the increase in the second ionisation energy of neon compared to its first is relatively small (less than twice the first). **(15)**

**2004 Question 10. (*b*)**

Describe how Bohr used line emission spectra to explain the existence of energy levels in atoms. **(13)**

(*i*) Why does each element have a unique line emission spectrum? **(6)**

(*ii*) The fact that each element has a unique line spectrum forms the basis for an instrumental technique which can be used to detect heavy metals and to measure their concentrations in a soil or a water sample. Name the instrumental technique.**(3)**

(*iii*) Bohr’s atomic theory was later modified. Give **one** reason why this theory was updated. **(3)**

**2005 Question 10**

(*b*) *The minimum energy required to completely remove the most loosely bound electron from a mole of gaseous atoms in their ground state* defines an important property of every element.

(*i*) Identify the energy quantity defined above. State the unit used to measure this quantity. **(7)**

(*ii*) Using **X** to represent an element, express the definition above in the form of a balanced chemical equation. **(6)**

(*iii*) Would it take more or less energy to remove the most loosely bound electron from an atom if that electron were not in its ground state? Explain. **(6)**

(*iv*) An element has a low first ionisation energy value and a low electronegativity value. What does this information tell you about how reactive the element is likely to be, and what is likely to happen to the atoms of the element when they react? **(6)**

**2005 Question 5.**

(*b*) Define *atomic radius* (*covalent radius*). **(6)**

Describe and account for the trend in atomic radii (covalent radii) of the elements

(*i*) across the second period, (*ii*) down any group, of the periodic table. **(15)**

(*c*) Define *covalent bond*. (6) Distinguish between a sigma (σ) and a pi (π) covalent bond. **(6)**

**2006 Question 5.**

**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)**

**2006 Question10 (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 6Li and 92.6 % of 7Li. **(6)**

**2007 Question 5.**

(*a*) Define *energy level*. **(5)**

Write the electron configuration (*s*, *p*) for the sulfur atom in its ground state, showing the arrangement in atomic orbitals of the highest energy electrons. **(6)**

State how many (*i*) energy levels, (*ii*) orbitals, are occupied in a sulfur atom in its ground state. **(6)**

(*b*) Use electronegativity values (Mathematical Tables p 46) to predict the type of bond expected between hydrogen and sulfur. Write the chemical formula for hydrogen sulfide. Use clear dot and cross diagrams to show the bonding in hydrogen sulfide. **(15)**

Would you expect the hydrogen sulfide molecule to be *linear* or *non-linear* in shape? Justify your answer. **(6)**

(*c*) Hydrogen sulfide has a boiling point of 212.3 K and water has a boiling point of 373 K. Account for the difference in the boiling points of these substances. **(6)**

Would you expect hydrogen sulfide to be soluble in water? Explain your answer. **(6)**

**2007 Question 11. (a)**



In 1910 Rutherford (pictured right) and his co-workers carried out an experiment in which thin sheets of gold foil were bombarded with alpha particles. The observations made during the experiment led to the discovery of the atomic nucleus.

(*i*) Describe the model of atomic structure which existed immediately *prior* to this experiment.  **(7)**

(*ii*) In this experiment it was observed that most of the alpha particles went straight through the gold foil. Two other observations were made. State these other observations and explain how each helped Rutherford deduce that the atom has a nucleus. **(12)**

**2008 Question 5**



(*a*) Define *electronegativity*. **(5)**

(*b*) State and explain the trend in electronegativity values down the first group in the periodic table of the elements. **(9)**

(*c*) Use electronegativity values to predict the types of bonding (*i*) in water, (*ii*) in methane,
 (*iii*) in magnesium chloride. **(9)**

(*d*) Use dot and cross diagrams to show the formation of bonds in magnesium chloride. **(6)**

(*e*) Explain the term *intermolecular forces*. **(6)**

(*f*) Use your knowledge of intermolecular forces to explain why methane has a very low boiling point (b.p. = –164 ºC).

The relative molecular mass of methane is only slightly lower than that of water but the boiling point of water is much higher (b.p. = 100 ºC). Suggest a reason for this. **(6)**

(*g*) The diagram shows a thin stream of liquid flowing from a burette.

A stream of water is deflected towards a positively charged rod whereas a stream of cyclohexane is undeflected. Account for these observations.

Explain what would happen in the case of the stream of water if the positively charged rod were replaced by a negatively charged rod. **(9)**

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**2008 Question 10**

1. (*i*) Define *energy level*. **(4)**

(*ii*) Distinguish between *ground state* and *excited state* for the electron in a hydrogen atom. **(6)**

The diagram shows how Bohr related the lines in the hydrogen emission spectrum to the existence of atomic energy levels.

 (*iii*) Name the series of lines in the visible part of the line emission spectrum of hydrogen. **(3)**

(*iv*) Explain how the expression *E*2 – *E*1 = *hf* links the occurrence of the visible lines in the hydrogen spectrum to energy levels in a hydrogen atom. **(12)**

**2009 Question 5**

**5.** (*a*) Define *first ionisation energy* of an element. **(8)**

(*b*) Use the values on page 45 of the Mathematics Tables to plot a graph on graph paper of first ionisation energy *versus* atomic number for the elements with atomic numbers from 10 to 20 inclusive. **(12)**

(*c*) Account fully for

(*i*) the general increase in ionisation energy values across the third period of the Periodic Table,

(*ii*) the peaks which occur in your graph at elements 12 and 15,

(*iii*) the sharp decrease in ionisation energy value between elements 18 and 19. **(18)**

(*d*) Write the *s*, *p* electron configuration for the potassium atom.

Hence state how many (*i*) energy sub-levels, (*ii*) individual orbitals, are occupied by electrons in a potassium atom.

Explain why there are electrons in the fourth main energy level of potassium although the third main energy level is incomplete. **(12)**



**2009 Question 11. (c)**

In 1922, Francis Aston, pictured right, was awarded the Nobel Prize in chemistry for detecting the existence of isotopes using the first mass spectrometer.

(*i*) What are isotopes? **(7)**

(*ii*) What is the principle of the mass spectrometer? **(9)**

(*iii*) Calculate, to two decimal places, the relative atomic mass of a sample of neon shown by mass spectrometer to be composed of 90.50% of neon–20 and 9.50% of neon–22. **(9)**

**Rates of Reaction**

**2003 Question 7.**

(*a*) Define *rate of a chemical reaction*. **(5)**

Calcium carbonate (marble chips) reacts with hydrochloric acid according to the following equation.

**CaCO3 + 2HCl = CaCl2 + CO2 + H2O**

Using simple experiments involving marble chips, **CaCO3**, and hydrochloric acid, **HCl**, describe how you could demonstrate the effects of (*i*) *particle size*, (*ii*) *concentration* on the rate of a chemical reaction. **(18)**

(*b*) What is a *catalyst*? **(6)**

Catalytic converters are used in cars.

(*i*) Identify **one** reaction which is catalysed in the catalytic converter in a car.

 State **one** of the environmental benefits of this process. **(12)**

(*ii*) Name **one** element used as a catalyst in a catalytic converter.
 What type of catalysis is involved in a catalytic converter? **(9)**

**2004 Question 8.**

(*a*) Define the *rate of a chemical reaction*. **(5)**

Explain why increasing the temperature has a significant effect on
the rate of a reaction. **(6)**

(*b*) The diagram shows a reaction profile diagram for an endothermic reaction. Name the quantities of energy marked **A** and **B**.

Copy this diagram into your answer book and indicate clearly on your diagram the likely effect of adding a catalyst on the energy profile for the reaction. **(12)**

(*c*) Catalytic converters are fitted to all modern cars with petrol engines.

Name **two** elements used as catalysts in a catalytic converter.

Name **one** substance which poisons the catalysts in a catalytic converter. **(9)**

(*d*) The oxidation of potassium sodium tartrate by hydrogen peroxide catalysed by cobalt (II) ions provides evidence for the intermediate formation theory of catalysis. State the observations you would make when carrying out this experiment. Explain how these observations provide evidence for the intermediate formation theory. **(18)**

**2005 Question 3.**

Hydrogen peroxide decomposes rapidly in the presence of a manganese (IV) oxide (**MnO2**) catalyst.
(*a*) Write a balanced equation for the decomposition of hydrogen peroxide. **(5)**

(*b*) Draw a labelled diagram of an apparatus a student could assemble to measure the rate of decomposition of hydrogen peroxide in the presence of a manganese (IV) oxide (**MnO2**) catalyst. Indicate clearly how the reaction could be started at a time known exactly, and how the gas produced is collected and its volume measured. **(12)**

(*c*) A student has a choice of using the same mass of finely powdered manganese (IV) oxide or coarsely

powdered (granulated) manganese(IV) oxide. Which of these would you expect to have a greater average rate of reaction over the first minute of the reaction? Give a reason for your answer. **(6)**

A set of results obtained in an experiment to measure the rate of decomposition of hydrogen peroxide, in a solution of known volume and concentration, is given in the table.



(*d*) Plot a graph to illustrate the volume of oxygen produced *versus* time. **(12)**

(*e*) Use the graph to determine (*i*) the volume of oxygen produced during the first 2.5 minutes and (*ii*) the instantaneous rate of the reaction at 2.5 minutes. **(9)**

(*f*) What changes would you expect in the graph if the experiment were repeated using a solution of the same volume but exactly half the concentration of the original hydrogen peroxide solution? **(6)**

**2006 Question 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.

**Na2S2O3 + 2HCl → 2NaCl + SO2 + S + H2O**

(*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)**

**2007 Question 9**

(*a*) Define the *rate of a chemical reaction*.

 Why does the rate of chemical reactions generally decrease with time? **(8)**

(*b*) The rate of reaction between an excess of marble chips (**CaCO3**) (diameter 11 – 15 mm) and 50 cm3 of 20 M
 hydrochloric acid was monitored by measuring the mass of carbon dioxide produced.



The table shows the total mass of carbon dioxide gas produced at stated intervals over 9 minutes.
Plot a graph of the mass of carbon dioxide produced *versus* time. **(12)**

Use the graph to determine (*i*) the instantaneous rate of reaction in grams per minute at 4.0 minutes,

(*ii*) the instantaneous rate of reaction at this time in moles per minute. **(9)**

(*c*) Describe and explain the effect on the rate of reaction of repeating the experiment using 50 cm3 of 1.0 M hydrochloric acid and the same mass of the same size marble chips. **(6)**

(*d*) Particle size has a critical effect on the rate of a chemical reaction.

(*i*) Mark clearly on your graph the approximate curve you would expect to plot if the experiment were repeated using 50 cm3 of 2.0 M HCl and using the same mass of marble chips but this time with a diameter range of

1 – 5 mm. **(6)**

(*ii*) Dust explosions present a risk in industry. Give **three** conditions necessary for a dust explosion to occur. **(9)

2008 Question 3.**

(*a*) Hydrogen peroxide solution is an oxidising reagent. Draw *or* describe the warning symbol put on a container of hydrogen peroxide solution to indicate this hazard. (5)

(*b*) Write a balanced equation for the decomposition of hydrogen peroxide. (6)

(*c*) Solid manganese(IV) oxide catalyst was added to a hydrogen peroxide solution at a time known exactly and the rate of production of gas was monitored as the hydrogen peroxide decomposed. Draw a labelled diagram of an apparatus that could be used to carry out this experiment. (12)

(*d*) The table shows the volumes of gas (at room temperature and pressure) produced at intervals over

****12 minutes.

Plot a graph of the volume of gas produced *versus* time. Explain why the graph is steepest at the beginning. (15)

(*e*) Use your graph to

(*i*) determine the instantaneous rate of gas production at 5 minutes,

(*ii*) calculate the total mass of gas produced in this experiment. (12)

**2009 Question 9**

(*a*) Explain (*i*) activation energy, (*ii*) effective collision. **(8)**

The effect of temperature on the rate of a chemical reaction was investigated using dilute solutions of hydrochloric acid and sodium thiosulfate. Suitable volumes and concentrations of the solutions were used.

The reaction is represented by the following balanced equation.

**2HCl + Na2S2O3** → **2NaCl + H2O + S + SO2**

Describe how the time for the reaction between the solutions of hydrochloric acid and sodium thiosulfate was obtained at room temperature. **(6)**

In a reaction mixture what effect, if any, does an increase in temperature of 10 K have on each of the following:

(*i*) the number of collisions, (*ii*) the effectiveness of the collisions, (*iii*) the activation energy. **(9)**



(*b*) The catalytic oxidation of methanol using platinum wire is illustrated in the diagram.

State **one** observation made during the experiment.

Name any **two** products of the oxidation reaction.

What type of catalysis is involved in this reaction? **(12)**

Explain **one** way in which the presence of the platinum catalyst speeds up the oxidation of the hot methanol.

Explain how a catalyst poison interferes with this type of catalysis. **(9)**

Give another example of a reaction which involves the same type of catalysis, indicating clearly the reactant(s) and the catalyst. **(6)**

**Volumetric Analysis**

**2002 Question 1**

Vinegar is a solution of ethanoic acid (acetic acid). Some bottles of vinegar are labelled “White Wine Vinegar”.

(a) What compound in white wine is converted to ethanoic acid in vinegar? What type of chemical process converts this compound to ethanoic acid? **(8)**

The concentration of ethanoic acid in vinegar was measured as follows: A 50 cm3 sample of vinegar was diluted to 500 cm3 using deionised water. The diluted solution was titrated against 25 cm3 portions of a standard 0.12 M sodium hydroxide solution, using a suitable indicator.

(b) Describe the procedure for accurately measuring the 50 cm3 sample of vinegar and diluting it to 500 cm3. **(12)**

(c) Name the piece of equipment that should be used to measure the ethanoic acid solution during the titration. State the procedure for washing and filling this piece of equipment in preparation for the titration. Name a suitable indicator for this titration. **(15)**

The titration reaction is

**CH3COOH + NaOH** → **CH3COONa + H2O**

After carrying out a number of accurate titrations of the diluted solution of ethanoic acid against the 25 cm3 portions of the standard 0.12 M sodium hydroxide solution, the mean titration figure was found to be 20.5 cm3.

(d) Calculate the concentration of ethanoic acid in the diluted vinegar solution in moles per litre and hence calculate the concentration of ethanoic acid in the original sample of vinegar.

Express this concentration in terms of % (w/v). **(15)**

**2003 Question 1**

Iron tablets may be used in the treatment of anaemia. To analyse the iron (II) content of commercially available iron tablets a student used four tablets, each of mass 0.360 g, to make up 250 cm3 of solution in a volumetric flask using dilute sulfuric acid and deionised water.

About 15 cm3 of dilute sulfuric acid was added to 25 cm3 portions of this iron (II) solution and the mixture then titrated with a 0.010 M solution of potassium manganate(VII), **KMnO4**.

(*a*) Why was it important to use dilute sulfuric acid as well as deionised water in making up the solution from the tablets?**(5)**

(*b*) Describe in detail the procedure for making up the 250 cm3 solution from the tablets. **(18)**

(*c*) Why was more dilute sulfuric acid added before the titrations were commenced? **(6)**

(*d*) How was the end-point detected? **(3)**

The titration reaction is described by the equation

**MnO4 + 5Fe2+ + 8H+ = Mn2+ + 5Fe3+ + 4H2O**

(*e*) In the titrations the 25 cm3 portions of the iron(II) solution made from the tablets required 13.9 cm3 of the
 0.010 M **KMnO4** solution. Calculate

 (*i*) the concentration of the iron(II) solution in moles per litre

 (*ii*) the mass of iron(II) in one tablet

(*iii*) the percentage by mass of iron(II) in each tablet. **(18)**

**2006 Question 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 (**Na2CO3.xH2O**). An 8.20 g sample of the crystals was weighed accurately on a clock glass and then made up to 500 cm3 of solution in a volumetric flask. A pipette was used to transfer 25.0 cm3 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 cm3.

The titration reaction is described by the equation:

 **Na2CO3 + 2HCl = 2NaCl + CO2 + H2O**

(*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 (**Na2CO3**) 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)**

**2007 Question 1**

A solution of sodium thiosulfate was prepared by weighing out a certain mass of crystalline sodium thiosulfate (**Na2S2O3.5H2O**) on a clock glass, dissolving it in deionised water and making the solution up carefully to 500 cm3 in a volumetric flask. A burette was filled with this solution and it was then titrated against 25.0 cm3 portions of previously standardised 0.05 M iodine solution in a conical flask. The average titre was 20.0 cm3. The equation for the titration reaction is

**2S2O32- + I2** → **2I¯ + S4O62-**

(*a*) Sodium thiosulfate is not a primary standard. Explain fully the underlined term. **(8)**

(*b*) Describe how the crystalline thiosulfate was dissolved and how the solution was transferred to the

 Volumetric flask and made up to exactly 500 cm3 **(15)**

(*c*) Pure iodine is almost completely insoluble in water.

 What must be added to bring iodine into aqueous solution? **(3)**

(*d*) A few drops of freshly prepared starch solution were added near the end point as the indicator for this titration. What sequence of colours was observed in the conical flask from the start of the titration until the end point was reached? **(12)**

(*e*) Calculate the molarity of the sodium thiosulfate solution and its concentration in grams of crystalline sodium thiosulfate (**Na2S2O3.5H2O**) per litre. **(12)**

**2008 Question 1**

To determine the concentration of ethanoic acid, **CH3COOH**, in a sample of vinegar, the vinegar was first diluted and then titrated against 25.0 cm3 portions of a previously standardised 0.10 M solution of sodium hydroxide, **NaOH**. One rough and two accurate titrations were carried out. The three titration figures recorded were 22.9, 22.6 and 22.7 cm3, respectively.

(*a*) Why was the vinegar diluted? **(5)**

(*b*) Describe the correct procedures for measuring exactly 25.0 cm3 of vinegar and diluting it to exactly
 250 cm3 using deionised water. **(15)**

(*c*) The equation for the titration reaction is:

**CH3COOH + NaOH →CH3COONa + H2O**

Name an indicator suitable for this titration. Justify your choice of indicator.

State the colour change at the end point.

Calculate the concentration of the diluted solution of ethanoic acid in (*i*) moles per litre,

 (*ii*) grams per litre. **(12)**

(*d*) State the concentration of ethanoic acid in the original vinegar sample in grams per litre.

 Express this concentration in terms of % (w/v). **(15)**

(*e*) Ethanoic acid is a carboxylic acid.

 Identify the carboxylic acid which occurs in nettles and stinging ants.  **(3)**

**2009 Question 1**

The Fe2+ content of iron tablets was determined by titration with a freshly standardised solution of potassium manganate(VII), **KMnO4.** The equation for the titration reaction is

**5Fe2+ + MnO4– + 8H+** → **5Fe3+ + Mn2+ + 4H2O**

(*a*) Why are iron tablets sometimes medically prescribed? **(5)**

(*b*) Why must potassium manganate (VII) solutions be standardised? Why was it necessary to standardise the potassium manganate (VII) solution *immediately* before use in the titration? What reagent is used for this purpose? **(9)**

(*c*) Describe how exactly 250 cm3 of Fe2+ solution was prepared from five iron tablets, each of mass 0.325 g.

 Why was some dilute sulfuric acid used in making up this solution? **(12)**

(*d*) Explain why additional dilute sulfuric acid must be added to the titration flask before each titration is carried
out. **(6)**

(*e*) On average, 18.75 cm3 of 0.01 M potassium manganate (VII) was required to react with 25.0 cm3 portions of the iron solution prepared from the five tablets.

Calculate

(*i*) the molarity of the Fe2+ solution,

(*ii*) the total mass of iron in the 250 cm3 of solution,

(*iii*) the percentage by mass of iron in the tablets. **(18)**