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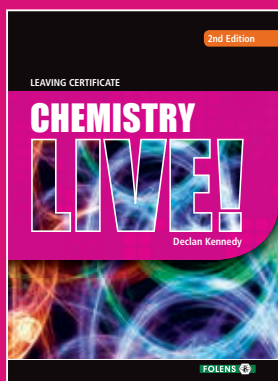


EXAM

**Exam Questions and
Solutions by Topic**

2002–2006

**COMPLEMENTS
USE OF**



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LEAVING CERTIFICATE

CHEMISTRY

HIGHER LEVEL

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2 The Atom

The Atom chapter gives an introduction to the development of the idea of the atom and the discovery of sub-atomic particles. Questions on this chapter have appeared in question 4 of the paper and as a part of question 11. It is important to be able to identify the pioneering scientists in this area of research and to be aware of their contributions, particularly the work of Ernest Rutherford and his team in the discovery of the nucleus.

2002 Question 11

- (b) 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)

Answer

- (b) Rutherford bombarded a gold foil (leaf) with alpha particles. (3)
He used a detector (fluorescent screen) to determine what happened to the alpha particles after they had struck the gold foil. (3)
- ⊗ Most of the alpha particles passed through the gold foil without being deflected.
 - ⊗ Some of the particles were deflected at large angles.
 - ⊗ Some of the alpha particles bounced back along their own path.

Any two of these three points: (2 × 3)

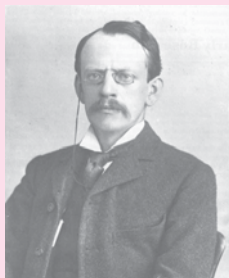
From this information Rutherford concluded that:

- ⊗ most of the mass of the atom is concentrated in a dense core (i.e. the nucleus).
- ⊗ the nucleus is positively charged.
- ⊗ most of the atom consists of empty space.

Any one of these three points: (6)

2006 Question 4

- (b) Name the scientist, shown in the photograph, who identified cathode rays as subatomic particles.



Answer

- (b) J.J. Thomson (6)

TIP: Full marks are awarded for the name Thomson. In general, this is the case when naming famous scientists.

3 Arrangements of Electrons in the Atom

The *Arrangements of Electrons in the Atom* chapter covers the nature of the electron and how electrons are distributed in the atom. It includes Mandatory Student Experiment No. 1, the flame tests (as given in the list of Mandatory Experiments at the start of the textbook). Details of this experiment are frequently asked as part of one of the questions in Section A of the exam paper or as part of question 10. It is important to learn the various definitions and principles from this chapter as they are often asked in questions 4, 5 and 10 of the exam. In particular, it is important to understand and to be able to clearly explain how Niels Bohr linked the idea of energy levels with emission spectra.

2002 Question 4

- (c) Define *atomic orbital*. (6)

Answer

- (c) An atomic orbital is the region in space within which there is a high probability of finding an electron. (6)

2002 Question 10

- (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)

Answer

- (i) Yellow [or orange] (4)
(ii) The main points are as follows:
⊗ Atoms exist in the ground state, where the electrons have the lowest amount of energy possible. (3)
⊗ If energy is provided to an atom in its ground state, a specific amount of this energy is absorbed and the electrons jump to higher energy levels, i.e. an excited state. (3)
⊗ This excited state is unstable and the electrons fall back down to lower energy levels. (3)
⊗ As electrons fall back, the excess energy is released in the form of light of a definite amount of energy. The frequency of the light depends on the difference in energy between the levels and is given by the equation $E_2 - E_1 = hf$. Therefore a line emission spectrum is obtained. (3)
(iii) Since only definite amounts of energy are emitted (3), this implies that electrons can occupy only definite energy levels (3).
(iv) Each element has its own unique emission spectrum (3) because each element has its own number of electrons distributed among the energy levels.

2003 Question 5

- (a) Define (i) *energy level* (ii) *atomic orbital*. (8)

Answer

- (i) An energy level is the fixed energy value that an electron in an atom may have. (4)
 (ii) An orbital is a region in space within which there is a high probability of finding an electron. (4)

2003 Question 10

- (c) A student was given samples of the following salts:

sodium sulfate (Na_2SO_4)

sodium sulfite (Na_2SO_3)

potassium sulfate (K_2SO_4)

- (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)

Answer

- (i) Flame test (4)
 The sodium salts would give a yellow [or orange] colour to the flame (3), and the potassium salt would give a lilac colour to the flame (3).

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)

Answer

- (b) The main points in Bohr's explanation of the fact that each element has a unique emission line spectrum are as follows:
- ⊗ Electrons in the ground state have energy of fixed values, i.e. the energy of the electrons in a particular orbit or energy level is quantised. (4)
 - ⊗ If energy is provided to an atom in its ground state, a specific amount of this energy is absorbed and the electrons jump to higher levels, i.e. an excited state. (3)
 - ⊗ This excited state is unstable and the electrons fall back down to lower energy levels. (3)
 - ⊗ As electrons fall back, the excess energy is released in the form of light of a definite amount of energy. The frequency of the light depends on the difference in energy between the levels and is given by the equation $E_2 - E_1 = hf$. Therefore a line emission spectrum is obtained. (3)
- (i) Each element has a unique emission line spectrum because each element has a different number of electrons, and each element has its own arrangement of these electrons in different energy levels [i.e. different electron configurations] (3). As a result, the different electron transitions (3) give rise to different emission spectra.
- (ii) atomic absorption spectrometry (3)

Exam Edge

- (iii) Bohr's theory only worked to explain the emission spectrum of hydrogen. When his theory was applied to atoms with more than one electron, it failed to account for many of the lines in the emission spectra of these atoms. (3)

TIP: Reference to the Heisenberg Uncertainty Principle in (iii) above would also merit full marks.

2005 Question 4

- (c) Name the series of coloured lines in the line emission spectrum of hydrogen corresponding to transitions of electrons from higher energy levels to the second ($n = 2$) energy level.
- (e) Distinguish between an *atomic orbital* and a *sub-level*.

Answer

- (c) the Balmer series (6)
- (e) An atomic orbital is the region in space within which there is a high probability of finding an electron. (3)
- A sub-level is a sub-division of a main energy level and consists of one or more orbitals of the same energy. (3)

TIP: When you are asked to distinguish between items such as the above, it is recommended that you simply state the definitions.

2006 Question 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)

Answer

- (i) Place a small sample of the potassium chloride on the end of a Nichrome wire (4). Use the wire to hold the salt in the flame of a Bunsen burner (4).

TIP: Note that the method of using a damp wooden splint was not awarded any marks (until 2009) even though it gives excellent results, is commonly used, and avoids cross contamination of the sample.

- (iii) Different elements have unique emission spectra because each element has a different number of electrons. Therefore, each element has its own arrangement of electrons in energy levels [i.e. different electron configurations] (3). Therefore, the different electron transitions (3) give rise to different emission spectra.

3 Arrangements of Electrons in the Atom

- (iii) atomic absorption spectrometry (3)
- (iv) An atomic orbital is the region in space within which there is a high probability of finding an electron. (6)
- (v) Heisenberg's Uncertainty Principle states that it is impossible to measure at the same time both the velocity (3) and the position (3) of an electron.

TIP: In the definition of Heisenberg's Uncertainty Principle it is important that you include the phrase "at the same time" when stating the definition. If you are unsure of this Principle, it is recommended that you study page 20 in your textbook.

4 The Periodic Table

Questions from the *Periodic Table* chapter appear regularly in questions 4 and 5 of the exam paper. It is important to know the definitions from this chapter of the textbook. It is essential to be able to accurately write out the electronic configuration of elements and ions. Be alert to those configurations that don't conform to what might be expected, such as those of chromium and copper. The technique of mass spectrometry features frequently on the exam paper and it is important to be able to explain the principle on which this separation technique is based. Aspects of the historical development of the Periodic Table appear regularly in exam questions and it is important to know the names of the relevant scientists, and to know the details of their involvement in this area of research. Calculation of relative atomic mass is straightforward, and you should be familiar with how to carry out this operation.

2002 Question 4

- (a) What are *isotopes*? (6)
(b) Write the electronic configuration of a neutral copper atom. (6)

Answer

- (a) Isotopes are atoms of the same element **(3)** (i.e. they have the same atomic number) that have different mass numbers due to the different number of neutrons **(3)** in the nucleus.
(b) Copper has 29 electrons: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^{10}$ (6)

2002 Question 7

- (c) What are the fundamental processes that occur in a mass spectrometer? (15)

Answer

- (c) The fundamental processes that occur in a mass spectrometer are (i) vaporisation, (ii) ionisation, (iii) acceleration, (iv) separation in a magnetic field and (v) detection. (5 × 3)

2003 Question 4

- (a) How many (i) electrons and (ii) neutrons has $^{37}_{17}\text{Cl}^-$? (6)
(d) On what principle is the analytical technique *mass spectrometry* based? (6)
(i) State *two* ways in which Mendeleev's periodic table of elements differs from that of Moseley. (6)

Answer

- (a) (i) 18 electrons **(3)**; (ii) 20 neutrons **(3)**
(d) The principle of mass spectrometry is that charged particles moving in a magnetic field **(3)** are separated according to the masses **(3)** of the particles.
(i) Mendeleev's table was arranged in order of increasing relative atomic mass (atomic weight), but the modern Periodic Table is arranged in order of increasing atomic number. (3)
Mendeleev's table had gaps in it while the modern Periodic Table has no gaps. (3)

Note: Further examples of differences are given on pages 29–30 of the textbook.

2003 Question 5

- (a) (iii) Write the electronic configuration (s, p, etc.) of nitrogen. (6)
- (iv) Describe how the electrons are arranged in the orbitals of the highest occupied sub-level of a nitrogen atom in its ground state. (3)

Answer

- (a) (iii) Nitrogen has seven electrons: $1s^2, 2s^2, 2p^3$ (3)
- (iv) Each p orbital contains one electron, i.e. $2p_x^1, 2p_y^1, 2p_z^1$ (3)

TIP: Full marks in part (iv) can be obtained by drawing a simple diagram such as the one shown in Fig. 4.1.

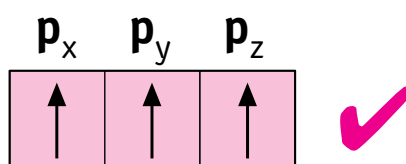


Fig. 4.1

2004 Question 4

- (a) Define relative atomic mass. (6)

Answer

- (a) Relative atomic mass is the average of the mass numbers of the isotopes of the element (3) as they occur naturally, taking their abundances into account and expressed on a scale in which the atoms of the carbon-12 isotope (3) have a mass of exactly 12 units. (3)

2004 Question 5

- (a) Write the electronic configuration (s, p, etc.) of the nitrogen atom. (5)

Answer

Nitrogen = $1s^2, 2s^2, 2p^3$ (5)

2005 Question 4

- (d) What contribution did Dobereiner make to the systematic arrangement of the elements? (6)

Answer

Dobereiner suggested that elements of similar properties (3) can be classified in groups of three [triads]. (3)

2005 Question 5

- (a) What are *isotopes*? (5)

Answer

- (a) Isotopes are atoms of the same element (i.e. they have the same atomic number) that have different mass numbers due to the different number of neutrons in the nucleus. (5)

2006 Question 4

- (a) Write the electron configuration (*s, p, etc.*) of a chromium atom in its ground state. (6)
- (f) What contribution did Newlands make to the systematic arrangement of the elements known to him? (6)

Answer

- (a) Chromium = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1, 3d^5$ (6)
- (f) Newlands arranged the known elements in order of increasing atomic weight (relative atomic mass) and noticed that the properties of the elements repeated with every eighth element [Law of Octaves]. (6)

2006 Question 10

- (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)

Answer

- (i) Isotopes are atoms of the same element (i.e. they have the same atomic number) that have different mass numbers due to the different number of neutrons in the nucleus. (4)
- (ii) Relative atomic mass is the average of the mass numbers of the isotopes of the element (3) as they occur naturally, taking their abundances into account (3) and expressed on a scale in which the atoms of the carbon-12 isotope (3) have a mass of exactly 12 units.
- (iii) The principle of mass spectrometry is that charged particles moving in a magnetic field (3) are separated (3) according to the masses (3) of the particles.
- (iv) In 100 atoms of lithium there are:
- | | | | | | |
|-------------------------|---|----------|---|-------|-----------|
| 7.4 atoms of mass 6 | = | 7.4 × 6 | = | 44.4 | |
| 92.6 atoms of mass 7 | = | 92.6 × 7 | = | 648.2 | |
| Total mass of 100 atoms | | | | = | 692.6 (3) |
| Average mass of 1 atom | | | | = | 6.926 (3) |

5 Chemical Bonding: Chemical Formulas

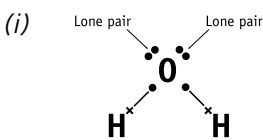
The *Chemical Bonding: Chemical Formulas* chapter is important in that it teaches you several fundamental principles that apply to many aspects of chemistry. Questions on this chapter appear frequently in questions 4, 5, 10 and 11 of the exam paper. You should be able to draw diagrams to show the electron arrangements in covalently-bonded molecules (showing only the electrons in the outside shells of the atoms). You should be familiar with how overlap of atomic orbitals is used to explain the formation of covalent bonds and be able to distinguish between sigma and pi bonding. You should also know how the number of bonding electron pairs and lone electron pairs influences the shape of covalent molecules. It is essential that you understand the concept of electronegativity, as it has a major bearing on chemical bonding and on the properties of substances. It is also important to understand what is meant by polarity and how this relates to intermolecular forces. You should understand the relationship between intermolecular forces and the solubility and boiling point of a substance. Be sure you can describe the demonstration experiment used to decide if a liquid is polar or non-polar.

2003 Question 5

- (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)



Answer

- (b) Electronegativity is the relative attraction (3) that an atom in a molecule has for the shared pair of electrons in a covalent bond (3).
- (i)  (9)
- Fig. 5.1
- (ii) The water molecule is V-shaped. (3)
The closest bond angle is 104° . (3)
The two lone pairs of electrons exert a greater force of repulsion than that between the two bond pairs (3) and so the bond angle is reduced from 109.5° to 104.5° (3).
- (c) The stream of water would be attracted to the rod. (6)
The water is attracted to the rod because water is polar. (3)

2004 Question 4

- (b) Account for the difference in the shapes of the ammonia (NH_3) and boron trifluoride (BF_3) molecules.
- (c) The boiling points of hydrogen and oxygen are 20.0 K and 90.2 K respectively. Account for the higher boiling point of oxygen.

Answer

- (b) The ammonia molecule is pyramidal because the central nitrogen atom is surrounded by three bond pairs and one lone pair of electrons (3). Boron trifluoride is triangular because the central boron atom is surrounded by three bond pairs (3).

Note: Full marks can also be obtained from the diagrams as shown in Fig. 5.3.



Fig. 5.3

- (c) The strength of van der Waals forces increases as the molecules get bigger. Oxygen is a bigger molecule than hydrogen, therefore the van der Waals forces are stronger in oxygen than they are in hydrogen. (6)

TIP: Another reason that oxygen has a higher boiling point is that it has a higher relative molecular mass than hydrogen. However, this explanation was only awarded (3) marks.

2004 Question 5

- (a) 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)

Answer

- (a)
- (3)

Fig. 5.4

The bonding in the nitrogen molecule consists of one sigma (3) and two pi bonds (3).

The type of intermolecular force that exists between nitrogen molecules is van der Waals forces. (3)

You would expect to find van der Waals forces in nitrogen gas because the nitrogen molecule is non-polar. (3)

2005 Question 4

- (a) Define *electronegativity*. (6)
- (b) What are the possible shapes for molecules of general formula AB_2 ? (6)

Answer

- (a) Electronegativity is the relative attraction (3) that an atom in a molecule has for the shared pair of electrons in a covalent bond (3).
- (b) The two possible shapes are linear (3) and V-shaped (3).

2005 Question 5

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

Answer

- (c) A covalent bond is the chemical bond formed by sharing (3) a pair of electrons (3).
 A sigma bond is formed by the head-on overlap of atomic orbitals. (3)
 A pi bond is formed by the sideways overlap of p orbitals. (3)

TIP: Diagrams showing sigma bond formation and pi bond formation can get full marks here provided they are properly labelled.

2006 Question 4

- (i) Under what circumstances can ionic compounds conduct electricity? (6)

Answer

- (i) Ionic compounds conduct electricity either when they are dissolved in water (3), or when they are molten (3).

2006 Question 5

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

Answer

- (b) (i) Electronegativity is the relative attraction (3) that an atom in a molecule has for the shared pair of electrons in a covalent bond (3).

6 Chemical Equations: Tests for Anions

The ability to correctly write balanced chemical equations is an essential skill for you to develop. While much of the *Chemical Equations: Tests for Anions* chapter is devoted to teaching you this skill, questions based on balancing equations are to be found throughout the exam paper. Note that in most cases, writing down the correct formula for each substance involved in a chemical reaction is awarded half of the marks, while the remaining marks are given for correct balancing. In general, no marks are awarded if any formula is incorrect, regardless of the balancing. Mandatory Student Experiment No. 2 (the anion tests) is included in this chapter and questions on parts of this experiment have appeared in Section A of the exam paper and also in question 4 and question 10 of Section B. In preparation for these questions, study Table 6.1: *Summary of tests for anions* on page 79 of your *Chemistry Live!* textbook.

2002 Question 4

(f) How could you test for the presence of nitrate ions in aqueous solution?

(6)

Answer

(f) Add the freshly prepared iron(II) sulfate solution to the solution of nitrate ions (3). Pour concentrated sulfuric acid down the inside of the test-tube. A brown ring (3) is formed at the junction of the two layers.

2003 Question 10

(c) A student was given samples of the following salts:

sodium sulfate (Na_2SO_4)

sodium sulfite (Na_2SO_3)

potassium sulfate (K_2SO_4)

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

(15)

Answer

(ii) To distinguish between the sulfate and sulfite salts, do the following:

- ⊗ Add barium chloride (3) solution to each test-tube; a white precipitate is formed (3).
- ⊗ Add dilute hydrochloric acid (3) to each test-tube. The sulfate precipitate is insoluble in dilute hydrochloric acid (3). The sulfite precipitate dissolves in dilute hydrochloric acid (3).

2004 Question 4

(j) How could the presence of sulfite ions in aqueous solution be detected?

(6)

Answer

(j) Add barium chloride solution, and the appearance of a white precipitate (3) that dissolves in HCl solution confirms presence of sulfite ions (3).

2005 Question 4

(f) How could you confirm the presence of nitrate ions in an aqueous solution?

(6)

Answer

- (f) Add the freshly prepared iron(II) sulfate solution to the solution of nitrate ions **(3)**. Pour concentrated sulfuric acid down the inside of the test-tube and a brown ring **(3)** is formed at the junction of the two layers.

Note: Marks were not awarded for the second point if you did not mention pouring the concentrated sulfuric acid down the inside of the test-tube.

7 Trends in the Periodic Table

Questions on the *Trends in the Periodic Table* chapter appear most frequently in questions 5, 10 and 11 of the exam paper. You need to know the definitions in this chapter well. You must also have a good understanding of the concept of first ionisation energy. You must be able to describe the trends in atomic radii, first ionisation energies and electronegativities of elements when moving across a period of the Periodic Table and also down a group. Equally importantly, you must be able to explain why these trends arise, with two reasons generally being required for full marks. You must also be able to describe and account for any deviations from these general trends, in particular with reference to first ionisation energy. In addition, trends in the properties of the elements of groups I, VII and 0 are included in this chapter.

2002 Question 5

Refer to the data in the Mathematics Tables, pages 44–46, in answering this question.

- Define *first ionisation energy*. (8)
- 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)
- Account fully for the trend in first ionisation energies of 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.

Ionisation	1 st	2 nd	3 rd	4 th	5 th	6 th	7 th	8 th
ionisation energy (kJ mol ⁻¹)	730	1450	7750	10500	13600	18000	21500	25600

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

Answer

- The first ionisation energy of an element is the energy required to completely remove the most loosely bound electron (3) from a neutral gaseous atom (3) in its ground state (2).

TIP: This was the first time that the phrase “in its ground state” was required. For many of the previous years, full marks for this definition were always awarded for just the first two points. The final point “in its ground state” had never been required before and is not found in any of the chemistry dictionaries! In the former syllabus, each question was marked out of 66 marks. An extra two marks had to be found to make the total marks equal 50 in the new syllabus, so those two marks are awarded for mentioning the ground state.

- There is a general increase in first ionisation energy across the period because of the following:
 - increasing nuclear charge (3)
 - decreasing atomic radius (3)

Exceptions to the general trend include the following:

- ☹ Beryllium is more stable (3) than boron, and nitrogen is more stable than oxygen (3).

7 Trends in the Periodic Table

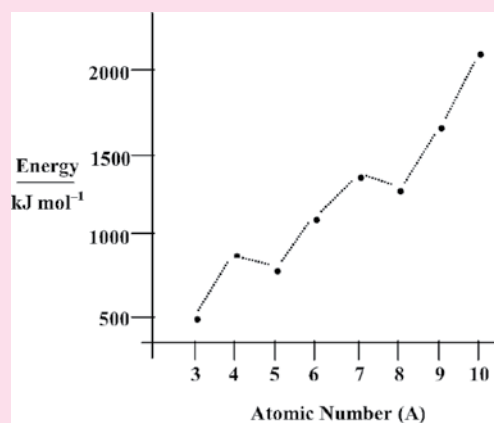
- ⊗ These exceptions in ionisation trends are explained by the fact that any sublevel that is completely filled (3), as is the case with beryllium and magnesium ($\text{Be} = 1s^2, 2s^2$ and $\text{Mg} = 1s^2, 2s^2, 2p^6, 3s^2$), or exactly half filled (3), as is the case with nitrogen and phosphorous ($\text{N} = 1s^2, 2s^2, 2p^3$ and $\text{P} = 1s^2, 2s^2, 2p^6, 3s^2, 3p^3$) has extra stability.

TIP: Be careful to use the term sublevel here, not level or orbital.

- (c) Down the group the first ionisation energy decreases for two reasons:
- ⊗ **Increasing atomic radius (3).** [The outermost electrons are becoming further away from the attractive force of the nucleus. Therefore, it becomes easier to remove an electron from the outer shell, i.e. the ionisation energy values decrease (3)].
 - ⊗ **The screening effect of inner electrons (3).** [Even though the nuclear charge increases down a group, this increase in nuclear charge is essentially cancelled out by the screening effect of the intervening shells of electrons. This means that the outermost electrons are somewhat shielded from the attractive force of the positively charged nucleus (3)].
- (d) The increase in ionisation energy values is explained by the following:
- ⊗ For each electron that is removed, the size of the positive charge on the ion left behind increases by one.
 - ⊗ The ion becomes progressively smaller.
- (Either one of these points gets 6 marks; the other point then gets 3 marks)
- (e) The dramatic increase in ionisation energy going from the second to the third ionisation is due to the ion having a filled outer shell (6) of electrons. [The filled outer shell ($2s^2 2p^6$) has extra stability and therefore will require more energy to remove electrons from it.]
- The next dramatic rise will be between the 10th and 11th electrons (3), as the 11th electron is also removed from a full shell (3).

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)



Answer

- (i) The ionisation energies increase across these elements because of the following:

- ⊗ increasing nuclear charge.
- ⊗ decreasing atomic radius.

(One correct point gets 4 marks; two correct points get 7 marks)

- (iii) The ionisation energy of element number 4 (beryllium) is exceptionally high because the sublevels are completely filled (3), and therefore has extra stability (3). [Alternatively, writing $\text{Be} = 1s^2, 2s^2$ merits 3 marks].

The ionisation energy of element number 7 (nitrogen) is exceptionally high because the sublevels are exactly half filled (3) and has extra stability (3). [Alternatively, writing $\text{N} = 1s^2, 2s^2, 2p^3$ merits 3 marks.]

- (iii) The second ionisation energy of an element refers to the removal of a second electron from the positive ion formed when the first electron is removed **(6)**. [Alternatively, writing the equation $X_{(g)}^+ \rightarrow X_{(g)}^{2+} + e^-$ merits full marks.]

2004 Question 5

- (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)**

Element	First ionisation energy (kJ mol ⁻¹)	Second ionisation energy (kJ mol ⁻¹)
Nitrogen	1400	2860
Oxygen	1310	3390
Neon	2080	3950
Sodium	494	4560

Answer

- (b) The first ionisation energy of an element is the energy required to completely remove the most loosely bound electron **(3)** from a neutral gaseous atom **(3)** in its ground state **(3)**.

The two principal reasons for the increase in ionisation energy across a period are as follows:

- ⊗ Increasing nuclear charge. **(3)**
- ⊗ Decreasing atomic radius. **(3)**

Nitrogen is more stable than oxygen **(3)** because the outermost sublevel of the nitrogen atom is half-full, i.e. $N = 1s^2, 2s^2, 2p^3$ **(3)**.

The second ionisation energies of sodium and neon differ because of the following:

- ⊗ When the single electron in the outer shell (3s) of the sodium atom is removed, the remaining ion will have the stable electronic configuration of neon ($2s^2 2p^6$) **(3)**. This means that relatively little energy is needed to remove this electron.
- ⊗ The second electron removed from sodium is therefore taken from a full shell, which is very stable and so far more energy is required to remove this electron. **(3)**
- ⊗ The second electron taken from neon comes from the same shell as the first, so the increase in ionisation energy is not as great as the first. **(3)**

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)**

Answer

- (b) The atomic radius (covalent radius) of an atom is defined as half the distance between the nuclei **(3)** of two atoms of the same element that are joined together by a single covalent bond **(3)**.

7 Trends in the Periodic Table

- (i) Atomic radii decrease across the second period (3). This is because there is an increase in nuclear charge (3), without a corresponding increase in the screening effect of the inner shells (3).
- (ii) When moving down a group, atomic radii increase (3) because there is an increase in the number of shells (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 was 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)

Answer

- (i) This is a definition of the first ionisation energy of an element (4). Ionisation energy is measured in kilojoules per mole (3).
- (ii) $X_{(g)} \rightarrow X_{(g)}^{+} + e^{-}$ (6)

TIP: Note that full marks requires the inclusion of the state symbols (i.e. "g" for gaseous).

- (iii) It would require less energy (3) to remove an electron from an atom in an excited state than if it were in its ground state. This is because if the atom is excited then it has already absorbed a certain amount of energy and the electron will have been promoted to a shell further out from the nucleus (3). Less energy will therefore be needed to remove the electron entirely from the atom.
- (iv) The element will be very reactive (3). The atom is likely to become a positively charged ion when it reacts because it will have a tendency to lose electrons (3).

Note: If the element had a high ionisation energy value then it would come from the right-hand side of the Periodic Table and would react by gaining electron(s). In that case, the low electronegativity would mean that the element would be relatively unreactive.

2006 Question 5

- (b) (ii) Explain why there is a general increase in electronegativity values across the periods in the periodic table of 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)

Answer

- (iii) Electronegativity increases across a period because there is an increase in nuclear charge (3) and a decrease in atomic radius (3).

Exam Edge

- (iii) The alkali metals react by losing the single electron in their outermost shell.
- ⊗ Down the group this electron becomes less strongly attracted to the nucleus.
 - ⊗ This happens because (i) the number of shells increases
 - ⊗ and (ii) the screening effect cancels out the increase in nuclear charge.
 - ⊗ This means that elements become more reactive down the group.

(Any one of these points gets 6 marks; any other point gets 3 marks)

8 Radioactivity

Questions on the *Radioactivity* chapter have appeared as part of questions 4, 5, 10 and 11 of the exam paper. You must know the historical background to the discovery of radioactivity and the contributions of Henri Becquerel and Pierre and Marie Curie. You must know what constitutes each of the three types of radioactive radiation and, for each type, be able to identify an isotope that emits this type of radiation. Knowing a use for each of these isotopes is important. You must be able to compare many of the properties of the three types of radioactive radiation, such as their penetrating ability. You will need to be able to write out nuclear equations for each of the types of radioactive decay, and be able to distinguish between nuclear and chemical reactions. You must also be able to describe the use of the carbon-14 isotope in radiocarbon dating.

2002 Question 11

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

Answer

- (b) An alpha particle is a group of two protons and two neutrons. (7)
[Alternatively: An alpha particle is a helium (4) nucleus (3) (not an atom)].

2003 Question 4

- (f) List the following three types of radiation in order of increasing penetrating power (6)
alpha- (α -) **beta-** (β -) **gamma-** (γ -)

Answer

- (f) The correct sequence is alpha; beta; gamma radiation. (6)

2004 Question 11

- (a) Define radioactivity. (6)
(i) State **two** properties of beta (β) particles. (6)
(ii) Write an equation for the nuclear reaction involved in the beta decay of ^{14}C (carbon-14). (6)
(iii) Explain how the carbon-14 isotope allows certain archaeological discoveries to be dated. (7)

Answer

- (a) Radioactivity is the spontaneous breaking up of unstable nuclei (3), with the emission of one or more types of radiation (3).
(i) Beta particles are negatively charged (3) and are stopped by a 5 mm thickness of aluminium (3).
[Alternatively: Electrons (3) that are emitted at high speed from the nucleus; they have negligible mass (3).]
(ii) $^{14}_6\text{C} \rightarrow ^{14}_7\text{N} + ^0_{-1}\text{e}$ (6)

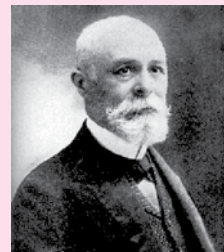
TIP: Don't mix the positioning of the atomic and mass numbers; the total charges and total masses must be equal on each side of the equation.

Exam Edge

- (iii) An organism takes in radioactive carbon-14 while it is alive (during photosynthesis or consumption), and some of this carbon-14 becomes incorporated into the tissues of the organism. The ratio of carbon-12 to carbon-14 remains constant until the organism dies **(3)**. After death, no more carbon-14 is taken in, and the carbon-14 in the organism's tissues begins to break down due to radioactive decay. As a result, the proportion of carbon-14 in the tissues decreases. The age of an archaeological artefact that contains plant or animal material can be estimated by measuring the ratio of carbon-12 to carbon-14 in the artefact, and by knowing the half-life of carbon-14 **(4)**.

2005 Question 5

- (a) Name the scientist pictured on the right who is credited with the discovery in 1896 that uranium salts emit radiation. **(3)**
Give an example of a radioactive isotope and state **one** common use made of this isotope. **(9)**



Answer

- (a) The scientist is Henri Becquerel. **(3)**
Americium-241 **(6)** is an example of a radioactive isotope and is used in smoke alarms **(3)**. [Other examples are carbon-14 **(6)** dating archaeological remains **(3)**, and cobalt-60 **(6)** used in cancer treatment **(3)**.]

TIP: Note that stating just the name of the isotope is insufficient. To obtain full marks, the mass number must be included in your answer.

2006 Question 4

- (c) Give **two** differences between a nuclear reaction and a chemical reaction. **(6)**

Answer

- (c) In a chemical reaction bonds are broken and new bonds are made. Chemical changes only involve electrons – the nuclei of the atoms stay the same **(3)**. However, in a nuclear reaction, the change in the nucleus of an element forms a new element **(3)**.

9, 10 & 11 The Mole Concept; Properties of Gases; Stoichiometry I

These three chapters, *The Mole Concept*, *Properties of Gases* and *Stoichiometry I*, are combined because related questions on the exam paper tend to be drawn from all three. These questions have appeared in questions 3, 4, 10 and 11 of the exam paper. Mandatory Student Experiment No. 3 (determining the relative molecular mass of a volatile liquid) is included in these chapters and has been asked in Section A. Make sure that you know all of the definitions and laws in these chapters. The ability to carry out several simple calculations is required in most of the questions based on these chapters. These include calculations of the number of moles and mass of a substance, the volume of a fixed amount of a gas under various conditions, and the number of particles present in a given amount of a substance. You also need to be able to calculate the percentage (by mass) of an element in a compound, and the empirical formula of a compound. An understanding of the kinetic theory of gases and the limitations of the concept of an ideal gas are needed. The use of the Equation of State for an Ideal Gas has appeared on several occasions – make sure you know the correct units to use in this formula! Finally, you must be able to relate the number of moles and masses of all substances involved in a given chemical reaction, and be able to base your answer on a balanced chemical equation.

2002 Question 4

(i) State Charles's law.

(6)

Answer

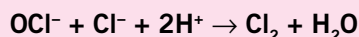
Charles' law states that, at constant pressure, the volume of a fixed mass of gas (3) is directly proportional to its temperature measured on the Kelvin scale (3).

2002 Question 11

(a) (i) In what type of household product would you expect to find sodium hypochlorite? (4)

(ii) A solution of sodium hypochlorite, NaOCl, is labelled as having a concentration of 5% (w/v). Express the concentration of the sodium hypochlorite solution in grams per litre. (6)

100 cm³ of this 5% (w/v) solution were reacted with excess chloride ion and acid according to the equation.



(iii) How many molecules of chlorine gas were liberated? (9)

(iv) What volume would this quantity of chlorine gas occupy at s.t.p.? (6)

Answer

(i) Sodium hypochlorite is found in bleach. (4)

(ii) 5% (w/v) = 5 grams in 100 cm³ (3)

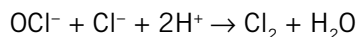
= 50 g/L (3)

Exam Edge

(iii) Relative molecular mass (M_r) of $\text{NaOCl} = 23 + 16 + 35.5 = 74.5$

$$\begin{aligned}\text{Number of moles in } 100 \text{ cm}^3 &= \frac{\text{mass}}{\text{relative molecular mass}} \\ &= \frac{5}{74.5} \\ &= 0.067 \text{ mol}\end{aligned}\quad (3)$$

From balanced equation:



1 mole $\text{OCl}^- \rightarrow 1$ mole of Cl_2

0.067 mole $\text{OCl}^- \rightarrow 0.067$ mole Cl_2 (3)

\Rightarrow Number of molecules of chlorine released $= 0.067 \times 6 \times 10^{23} = 4 \times 10^{22}$ (3)

(iv) Volume occupied by 0.067 moles at s.t.p. $= 0.067 \times 22.4 = 1.5$ litres (6)

2003 Question 3

In an experiment to determine the relative molecular mass of a volatile liquid a sample of the liquid is vaporised at a given temperature and pressure and its volume is measured. The mass of the sample is also measured. The number of moles of liquid is then calculated using the formula $PV = nRT$ and from this the relative molecular mass of the liquid is calculated.

(a) What is meant by a *volatile liquid*? (5)

(b) Describe with the aid of a labelled diagram how you would carry out this experiment to determine the relative molecular mass of a volatile liquid. From your description, it should be clear how the *mass*, *volume*, and *temperature*, of the sample are measured. (21)

(c) How may the *pressure* be measured? (6)

In an experiment to measure the relative molecular mass of a volatile liquid 0.275 g of the liquid was vaporised at 97°C . The volume occupied was found to be 95 cm^3 . The pressure was $1 \times 10^5 \text{ Pa}$.

(d) Calculate the number of moles of the volatile liquid vaporised. (12)

(e) Calculate the relative molecular mass of the volatile liquid. (6)

Answer

(a) A volatile liquid is a liquid that has a low boiling point (or vaporises easily). (5)

(b)

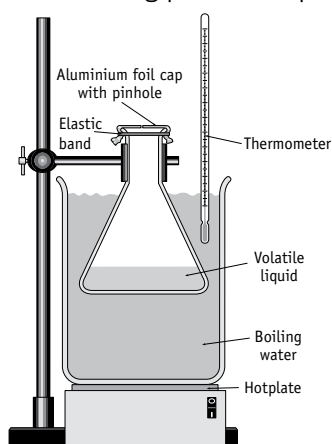


Fig. 10.1

TIP: In the diagram you must have at least one of the following labelled: flask, covered with foil with a pinhole; flask immersed so that at least half is under water.

(3 marks for any one correct label)

9, 10 & 11 The Mole Concept; Properties of Gases; Stoichiometry I

The key points in the description of the experiment are as follows:

Find the mass of the liquid:

- ⊗ Get mass of foil and flask, add liquid and arrange as in diagram. (3)
- ⊗ Heat until the liquid is gone [heat until flask appears empty]. (3)
- ⊗ Cool, dry and reweigh foil and flask. (3)
- ⊗ Get mass of the sample by subtraction. (3)

Find the volume of the liquid:

- ⊗ Fill flask and empty into graduated cylinder. (3)

Find the temperature of the vapour:

- ⊗ Use thermometer to read temperature of water. (3)

TIP: This last point can be obtained from your diagram if you have a thermometer in it.

An alternative experiment:

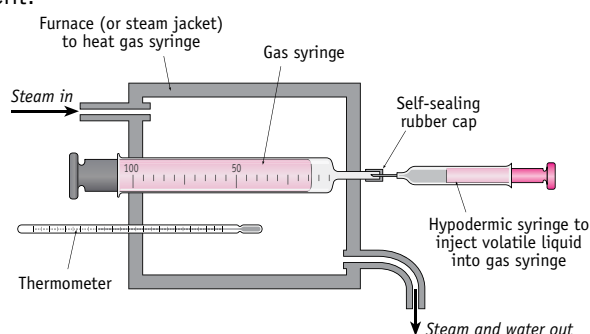


Fig. 10.2

- ⊗ Draw a small volume of air into the gas syringe and place it into the furnace. (3)
- ⊗ Pass steam through the furnace and make a note of the reading on the thermometer when the temperature has become steady. (3)
- ⊗ Record the volume of air in the gas syringe (V_1). (3)
- ⊗ Draw a small volume of the volatile liquid into the hypodermic syringe and find their combined mass (M_1). (3)
- ⊗ Inject some of the volatile liquid into the gas syringe. (3)
- ⊗ Reweigh the hypodermic syringe and its contents (M_2). (3)
- ⊗ By subtraction, find the mass of the volatile liquid that was injected ($M_2 - M_1$). (3)
- ⊗ Record the total volume of air plus vapour in the gas syringe (V_2), and by subtraction find the volume of the vaporised volatile liquid ($V_2 - V_1$). (3)

(c) The pressure is measured using a barometer. (6)

(d) Pressure = 1×10^5 Pa

Volume = 95×10^{-6} m³ (3)

$R = 8.3 \text{ J K}^{-1}\text{mol}^{-1}$

$T = (97 + 273) = 370 \text{ K}$ (3)

$$pV = nRT$$

$$\Rightarrow n = \frac{pV}{RT}$$

$$= \frac{(1 \times 10^5)(95 \times 10^{-6})}{(8.3)(370)} \quad (3)$$

$$= 0.0031 \text{ mol} \quad (3)$$

(e) 0.0031 moles of liquid = 0.275 g

$$\text{Therefore, 1 mole of liquid} = \frac{0.275}{0.0031} \quad (3)$$

$$= 88.7 \quad (3)$$

2003 Question 4

- (b) How many electrons are there in 2.3 g of sodium metal, **Na**? (6)
- (c) The famous Irish scientist shown on the right, was born in 1627. He was a son of the Earl of Cork. Give a statement on the gas law that bears his name. (6)
- (g) What is the percentage by mass of nitrogen in ammonium nitrate, **NH₄NO₃**? (6)



Answer

- (b) 1 mole of sodium metal contains 6×10^{23} atoms
 23 g of sodium metal contains 6×10^{23} atoms
 2.3 g of sodium metal contains $0.1 \times 6 \times 10^{23}$ atoms (3)
 $= 6 \times 10^{22}$ atoms
 Each atom of Na contains 11 electrons
 \Rightarrow Total number of electrons $= 11 \times 6 \times 10^{22}$
 $= 6.6 \times 10^{23}$ (3)
- (c) Boyle's law states that, at constant temperature, the volume of a fixed mass of gas (3) is inversely proportional to its pressure (3).
- (g) M_r of **NH₄NO₃** $= (2 \times 14) + (3 \times 16) + (4 \times 1) = 80$
 Percentage of nitrogen $= \frac{28 \times 100}{80}$ (3)
 $= 35\%$ (3)

2004 Question 4

- (d) State *Charles' law*. (6)
- (h) What is the percentage by mass of iron in iron(III) oxide (**Fe₂O₃**)? (6)

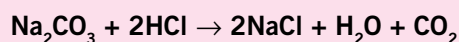
Answer

- (d) Charles' law states that at constant pressure, the volume of a fixed mass of gas (3) is directly proportional to its temperature when measured on the Kelvin scale (3).
- (h) M_r of **Fe₂O₃** $= (2 \times 56) + (3 \times 16) = 160$
 Percentage of Fe $= \frac{112 \times 100}{160}$ (3)
 $= 70\%$ (3)

2004 Question 10

- (a) Hydrochloric acid is severely corrosive to skin and eyes and toxic by inhalation or ingestion. It should be handled carefully and stored safely.

The entire contents of a bottle containing 2.5 litres of concentrated hydrochloric acid were accidentally spilled in a laboratory. The spilled acid was neutralised by adding solid powdered sodium carbonate. The neutralisation reaction is described by the following equation.



The spilled acid was a 36% (w/v) solution of hydrogen chloride in water.

- (i) Calculate the number of moles of hydrochloric acid spilled. (10)
 (ii) What was the minimum mass of anhydrous sodium carbonate required to completely neutralise all of the spilled hydrochloric acid? (9)
 (iii) What volume of carbon dioxide in litres, at room temperature and pressure, was produced in this neutralisation reaction? (6)

Answer

- (i) 36% (w/v) = 36 grams in 100 cm³ (4)

i.e. in 100 cm³ of HCl there are 36 g

in 1 cm³ of HCl there are $\frac{36}{100}$ grams

In 2.5 L (2500 cm³) there are $\frac{36 \times 2500}{100}$
 = 900 grams (3)

M_r of HCl = 1 + 35.5 = 36.5

\Rightarrow Number of moles of HCl = $\frac{\text{mass}}{\text{relative molecular mass}}$
 $= \frac{900}{36.5}$
 = 24.66 moles (3)

- (ii) From the balanced equation: $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$

From the equation we know that 2 moles HCl reacts with 1 mole of Na_2CO_3

\Rightarrow 24.66 mol of HCl will completely neutralise 12.33 mol of Na_2CO_3 (3)

M_r of Na_2CO_3 = $(2 \times 23) + 12 + (3 \times 16) = 106$

\Rightarrow Mass of 12.33 moles of Na_2CO_3 = 12.33×106
 = 1306.98 grams (6)

- (iii) From the balanced equation $\text{Na}_2\text{CO}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{H}_2\text{O} + \text{CO}_2$

1 mole of $\text{Na}_2\text{CO}_3 \rightarrow$ 1 mole CO_2

12.33 mole of $\text{Na}_2\text{CO}_3 \rightarrow$ 12.33 mole CO_2 (3)

Volume of 12.33 mol of CO_2 at room temperature and pressure

= 12.33×24
 = 295.92 litres (3)

Note: If you use 22.4 litres in the calculation (i.e. the value for molar volume of a gas at s.t.p.) then you will lose 3 marks.

2004 Question 10

- (c) State *Avogadro's law*. (5)
- (i) What is an ideal gas? (5)
- (ii) State **one** reason why ammonia gas deviates from ideal gas behaviour. (3)
- (iii) A small quantity of the volatile organic solvent propanone (**C₃H₆O**) evaporates at room temperature and pressure. Use the equation of state for an ideal gas to calculate the volume, in litres, of propanone vapour formed when 0.29 g of liquid propanone evaporates taking room temperature as 20 °C and room pressure as 101 kPa. (12)

Answer

- (c) Avogadro's law states that equal volumes of gases contain equal numbers of molecules (3), under the same conditions of temperature and pressure (2).
- (i) An ideal gas is one that obeys all the assumptions of the kinetic theory of gases (3) under all conditions of temperature and pressure (2).
- (ii) Ammonia is polar (3) and so there are strong forces of attraction [intermolecular forces or hydrogen bonds] between the molecules (3).
- (iii) M_r of C₃H₆O = (3 × 12) + (6 × 1) + 16 = 58
- $$\Rightarrow \text{Number of moles (n) of C}_3\text{H}_6\text{O} = \frac{0.29}{58} = 0.005$$

Note: When using the Equation of State for an ideal gas, the pressure must be given in pascals, the volume in cubic metres and the temperature in kelvin. Also note that there are one thousand litres in a cubic metre.

$$p = 101 \times 10^3 \text{ Pa}$$

$$n = 0.005$$

$$R = 8.3 \text{ JK}^{-1}\text{mol}^{-1}$$

$$T = (273 + 20) = 293 \text{ K}$$

$$pV = nRT$$

$$\Rightarrow V = \frac{nRT}{p}$$

$$= \frac{0.005 \times 8.3 \times 293}{101 \times 10^3} \quad (3)$$

$$= 1.2 \times 10^{-4} \text{ m}^3 \quad (3)$$

$$= 0.12 \text{ litres} \quad (3)$$

2005 Question 4

- (h) When 3.175 g of copper reacts with chlorine gas 6.725 g of copper chloride is formed. Find by calculation the empirical formula of the chloride.

Answer

- (h) Mass of chlorine = 6.725 g – 3.175 g = 3.55 g
- Number of moles of Cu = $\frac{3.175}{63.5}$ = 0.05
- Number of moles of Cl atoms = $\frac{3.55}{35.5}$ = 0.1 (3)

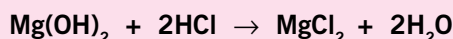
$$\text{Cu} : \text{Cl} = 0.05 : 0.1 = 1 : 2$$

⇒ Empirical formula = CuCl_2

(3)

2005 Question 10

- (a) An indigestion tablet contains a mass of 0.30 g of magnesium hydroxide [Mg(OH)_2] as its only basic ingredient. The balanced chemical equation for the reaction between magnesium hydroxide and hydrochloric acid ($\text{HCl}_{(\text{aq})}$), the acid produced in the stomach, is as follows:



- (i) Calculate the volume of 1.0 M HCl neutralised by two of these indigestion tablets. Give your answer correct to the nearest cm^3 . (8)
- (ii) What mass of salt is formed in this neutralisation? (5)
- (iii) How many magnesium ions are present in this amount of the salt? (6)
- (iv) Another indigestion remedy consists of a suspension of magnesium hydroxide [Mg(OH)_2] in water and is marked 6% (w/v). What volume of this second indigestion remedy would have the same neutralising effect on stomach acid as two of the indigestion tablets mentioned earlier? (6)

Answer

- (i) Mass of Mg(OH)_2 in two tablets = $2 \times 0.3 = 0.6 \text{ g}$

$$M_r \text{ of } \text{Mg(OH)}_2 = 24 + (2 \times 16) + (2 \times 1) = 58$$

$$\Rightarrow \text{Number of moles of } \text{Mg(OH)}_2 = \frac{0.6}{58} = 0.01 \quad (3)$$

From the balanced equation $\text{Mg(OH)}_2 + 2\text{HCl} \rightarrow \text{MgCl}_2 + 2\text{H}_2\text{O}$

From equation 1 mole of Mg(OH)_2 reacts with 2 moles of HCl

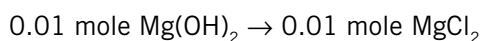
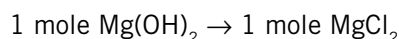
$$\Rightarrow \text{Number of moles of } \text{HCl} \text{ required} = 0.01 \times 2 = 0.02 \quad (3)$$

1 mole of HCl is contained in 1 litre of solution

⇒ 0.02 mole would be contained in 0.02 litre of solution

$$= 20 \text{ cm}^3 \text{ of } \text{HCl} \quad (2)$$

- (ii) From the balanced equation $\text{Mg(OH)}_2 + 2\text{HCl} \rightarrow \text{MgCl}_2 + 2\text{H}_2\text{O}$



$$\Rightarrow \text{Number of moles of } \text{MgCl}_2 \text{ formed} = 0.01 \quad (3)$$

$$M_r \text{ of } \text{MgCl}_2 = 24 + (35.5 \times 2) = 95$$

$$\Rightarrow \text{Mass of } \text{MgCl}_2 \text{ formed} = 0.01 \times 95 = 0.95 \text{ g} \quad (2)$$

- (iii) 1 mole of MgCl_2 contains 6×10^{23} magnesium ions

$$0.01 \text{ mole of } \text{MgCl}_2 \text{ contains } 0.01 \times 6 \times 10^{23} \text{ magnesium ions} \quad (3)$$

$$= 6 \times 10^{21} \text{ Mg ions} \quad (3)$$

- (iv) Mass of Mg(OH)_2 required (as in original sample) = 0.6 g

$$6\% \text{ (w/v)} = 6 \text{ grams in } 100 \text{ cm}^3 \quad (3)$$

$$\Rightarrow \text{There would be } 0.6 \text{ g in } 10 \text{ cm}^3 \quad (3)$$

2005 Question 11

- (b) (i) Define a *mole of a substance*. (7)
- (ii) State *Avogadro's law*. (6)
- (iii) A foil balloon has a capacity of 10 litres. How many atoms of helium occupy this balloon when it is filled with a 10% (v/v) mixture of helium in air at room temperature and pressure? (12)

Answer

- (i) One mole of a substance is the amount of that substance which contains 6×10^{23} particles of that substance. (7)
- (ii) Avogadro's law states that equal volumes of gases contain equal numbers of molecules (3), under the same conditions of temperature and pressure (3).
- (iii) 10% (v/v) means that 100 litres of the mixture contain 10 litres of helium. (3)
- \Rightarrow In 10 litres of the mixture there will be 1 litre of helium. (3)
- 1 mole occupies 24 litres at room temperature and pressure.
- \Rightarrow 1 litre is occupied by $\frac{1}{24}$ mole of helium
- $= 0.0417 \text{ mol}$ (3)
- 1 mole of helium contains 6×10^{23} atoms
- Therefore, 0.0417 mole of helium contains $0.0417 \times 6 \times 10^{23}$ atoms
- $= 2.5 \times 10^{22}$ atoms (3)

2006 Question 4

- (d) Calculate the percentage carbon, by mass, in methylbenzene. (3)

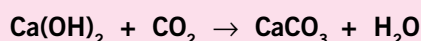
Answer

$$M_r \text{ of } C_7H_8 = (7 \times 12) + (8 \times 1) = 92$$

$$\begin{aligned} \text{Percentage of carbon} &= \frac{84}{92} \times 100 \\ &= 91.3\% \end{aligned} \quad (3)$$

2006 Question 11

- (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 cm³ of the gas at 10 °C and a pressure of 1×10^5 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)

Answer

- (i) An ideal gas is one that obeys all of the assumptions of the kinetic theory of gases under all conditions of temperature and pressure. (4)
- (ii) One reason that carbon dioxide deviates from ideal behaviour is because there are strong forces of attraction between the molecules [or because the molecules occupy a volume]. (3)
- (iii) $p = 1 \times 10^5 \text{ Pa}$
 $V = 720 \times 10^{-6} \text{ m}^3$ (3)
 $R = 8.3 \text{ JK}^{-1}\text{mol}^{-1}$
 $T = (10 + 273) = 283 \text{ K}$ (3)
 $pV = nRT$
 $\Rightarrow n = \frac{pV}{RT}$
 $= \frac{(1 \times 10^5)(720 \times 10^{-6})}{(8.3)(283)}$
 $= 0.03 \text{ mole (correct to one significant figure)}$ (3)

Alternatively, this problem could also be solved using the Combined Gas Law.

Given	s.t.p.
$V_1 = 720 \text{ cm}^3$	$V_2 = ?$
$p_1 = 1 \times 10^5 \text{ kPa}$	$p_2 = 1 \times 10^5 \text{ kPa}$
$T_1 = 10^\circ\text{C} + 273 = 283 \text{ K}$	$T_2 = 273 \text{ K}$

Substitute into

$$\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$$

$$\frac{1 \times 10^5 \times 720}{283} = \frac{1 \times 10^5 \times V_2}{273} \quad (3)$$

$$V_2 = \frac{1 \times 10^5 \times 720 \times 273}{1 \times 10^5 \times 283} \quad (3)$$

$$= 694.56 \text{ cm}^3$$

At s.t.p. 22.4 L is occupied by one mole

22,400 cm³ is occupied by 1 mole

1 cm³ is occupied by $\frac{1}{22400}$ mole

Therefore, 694.56 cm³ is occupied by $\frac{694.56}{22400}$ mole
 $= 0.03 \text{ mole (correct to one significant figure)}$ (3)

- (iv) 1 mole of CO₂ contains 6×10^{23} molecules
 0.03 mole of CO₂ contains $0.03 \times 6 \times 10^{23}$
 $= 1.8 \times 10^{22}$ molecules (3)
- (v) From the balanced equation: $\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$
 1 mole of Ca(OH)₂ reacts with 1 mole of CO₂
 Therefore, 0.03 mole Ca(OH)₂ reacts with 0.03 mole CO₂ (3)
 M_r of Ca(OH)₂ = $40 + (2 \times 16) + (2 \times 1) = 74$
 1 mole of Ca(OH)₂ = 74 g
 0.03 mole Ca(OH)₂ = 0.03×74
 $= 2.22 \text{ g}$ (3)

12 Acids and Bases

The *Acids and Bases* chapter is usually examined on the Leaving Certificate exam paper by means of short questions given as part of full questions. The questions are usually based on the theory of acids and bases as well as an understanding of conjugate acid-base pairs. Pay particular attention to the definitions highlighted in this chapter of *Chemistry Live!* as these definitions are the key to getting good marks on the exam paper.

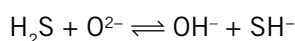
2003 Question 8

- (a) Define (i) *an acid*, (ii) *a base* according to the Brønsted-Lowry theory. (8)
Identify the *acid*, and *conjugate acid* in the following system. (6)



Answer

- (i) An acid is a proton donor. (4)
(ii) A base is a proton acceptor. (4)



H_2S is the acid (since it donates a proton). (3)

OH^- is the conjugate acid. (3)

Note: OH^- is an acid because it donates a proton to SH^- . The base O^{2-} changes into this conjugate acid by accepting a proton.

2004 Question 4

- (e) Write (i) the conjugate acid and (ii) the conjugate base of HPO_4^{2-} .

Answer

- (i) The conjugate acid is H_2PO_4^- . (Recall that a base changes into a conjugate acid when it accepts a proton.) (3)
(ii) The conjugate base is PO_4^{3-} . (Recall that an acid changes into a conjugate base when it donates a proton.) (3)

2006 Question 4

- (e) What is (i) the conjugate acid and (ii) the conjugate base of H_2O ?

Answer

- (i) H_3O^+ (Recall that a base changes into a conjugate acid when it accepts a proton.) (3)
(ii) OH^- (Recall that an acid changes into a conjugate base when it donates a proton.) (3)

13 Volumetric Analysis: Acid-Base

The questions on the *Volumetric Analysis: Acid-Base* chapter usually appear in question 1 of the Leaving Certificate Higher Level Chemistry examination paper. A full question (question 1) is always given on volumetric analysis. This question is based on either *Volumetric Analysis: Acid-Base* (Chapter 13) or *Volumetric Analysis: Oxidation-Reduction* (Chapter 15) or on some of the titrations related to water analysis (Chapter 19). In addition, some short questions given as part of question 1 may be based on material covered in other chapters of the textbook, e.g. acid-base indicators (Chapter 18), or the naming of carboxylic acids (Chapter 22).

2002 Question 1

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

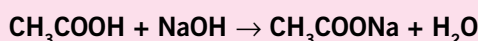
- (a) What compound in white wine is converted to ethanoic acid in vinegar?
What type of chemical process converts this compound into ethanoic acid? (8)

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



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 the ethanoic acid in the original sample of vinegar.
Express this concentration in terms of % (w/v). (15)

Answer

- (a) Ethanol. [The formula C₂H₅OH is also acceptable for full marks.] (4)
The process which converts this compound to ethanoic acid is known as oxidation. (4)
- (b) The main steps for measuring out the 50 cm³ of vinegar and diluting to 500 cm³ are as follows:
- ⊗ Pipette two 25 cm³ samples of vinegar (3) into a clean volumetric flask (3).
 - ⊗ Add deionised water. (3)
 - ⊗ Make sure the bottom of the meniscus is at the mark on the volumetric flask (3).

TIP: Marks for the first point above (6 marks in total) can also be obtained using a labelled diagram as shown in Fig. 13.1.

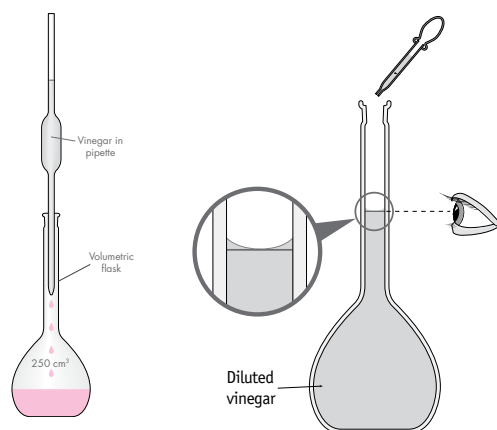


Fig. 13.1

- (c) The piece of equipment used to measure the ethanoic acid solution is called a burette. (3)

The procedure for washing and filling the burette is as follows:

- ⊗ Rinse out the burette with deionised water and with some of the diluted vinegar.
- ⊗ Using a small funnel, fill the burette with the diluted vinegar. Remove the funnel.
- ⊗ Ensure that the space below the tap is filled.
- ⊗ Adjust the bottom of the meniscus to the zero graduation mark. **Any three of these points: (3 × 3)**

A suitable indicator for this titration is phenolphthalein. (3)

- (d) Given:

$$V_a = 20.5$$

$$M_a = ?$$

$$n_a = 1$$

$$V_b = 25$$

$$M_b = 0.12$$

$$n_b = 1$$

$$\frac{V_a \times M_a}{n_a} = \frac{V_b \times M_b}{n_b}$$

$$\frac{20.5 \times M_a}{1} = \frac{25 \times 0.12}{1} \quad (3)$$

$$M_a = \frac{25 \times 0.12}{20.5}$$

$$= 0.15 \text{ moles per litre.} \quad (3)$$

TIP: Expressing your answer to two decimal places is perfectly acceptable since the maximum number of decimal places used in the question is two.

Since the solution was diluted by a factor of ten, the concentration of ethanoic acid in original vinegar is as follows:

$$= 0.15 \times 10$$

$$= 1.5 \text{ mol/L} \quad (3)$$

$$= 1.5 \times 60 \text{ g/L}$$

$$= 90 \text{ g/L} \quad (3)$$

Since there are 90 g of ethanoic acid in a litre of vinegar, there are 9 g of ethanoic acid in 100 cm³ of vinegar. Therefore, % w/v = 9% (3)

Answers:

Concentration of ethanoic acid in diluted vinegar = 0.15 mol/L

Concentration of ethanoic acid in the original sample of vinegar = 1.5 mol/L

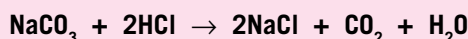
Concentration of ethanoic acid in original sample of vinegar = 9 % w/v.

TIP: It is important that you state clearly the answers to the questions asked on the examination paper. Using the above layout will make the answers very clear to the examiner.

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 ($\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$). An 8.20 g sample of the crystals was weighed accurately on a clock glass and then made up to 500 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 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. (12)
- (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 (Na_2CO_3) in the solution in
(i) moles per litre, (9)
(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)

Answer

- (a) Anhydrous sodium carbonate. (5)

TIP: If you left out the word “anhydrous” you lost (2) marks. If you are unsure of the concept of primary standards, study pages 168–169 of the textbook.

- (b) A suitable indicator is methyl orange. (3)
The colour change observed at the end point is from yellow (3) to pink (3). [The colours “orange” and “red” are also acceptable.]
No more than 1 or 2 drops of indicator should be used because an indicator is a weak acid or a weak base and too much indicator could affect the accuracy of the titration. (3)
- (c) (i) 1. Rinse out the burette with deionised water. (3)
2. Rinse out the burette with the hydrochloric acid solution. (3)
- (ii) It is important to fill the part below the tap of the burette because if this is not done it will cause an inaccurate result. This is because some of the measured volume would be used to fill the part below the tap, i.e. all the measured amount of solution has not been delivered into the conical flask. (6)

(d) (i) Given:

$$V_a = 26.05$$

$$M_a = 0.11$$

$$n_a = 2$$

$$V_b = 25$$

$$M_b = ?$$

$$n_b = 1$$

$$\frac{V_a \times M_a}{n_a} = \frac{V_b \times M_b}{n_b}$$

$$\frac{26.05 \times 0.11}{2} = \frac{25 \times M_b}{1}$$

$$M_b = \frac{26.05 \times 0.11}{2 \times 25} \quad (3)$$

$$= 0.05731 \text{ moles per litre}$$

$$= 0.0573 \times 106$$

$$= 6.075 \text{ grams per litre} \quad (3)$$

Answers: (i) 0.05731 mol/L

(ii) 6.075 g/L

(e) Since the concentration of the sodium carbonate solution is 0.05731 moles per litre, there must be

$$\frac{0.05731}{2} = 0.0286 \text{ moles of sodium carbonate in } 500 \text{ cm}^3 \text{ solution.}$$

$$1 \text{ mole Na}_2\text{CO}_3 = 106 \text{ g}$$

$$\therefore 0.0286 \text{ moles Na}_2\text{CO}_3 = 0.0286 \times 106 = 3.0316 \text{ g}$$

In the 8.20 g of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ crystals, there are only 3.0316 g of Na_2CO_3 .

$$\therefore \text{mass of water of crystallisation} = 8.20 - 3.0316$$

$$= 5.1684 \text{ g}$$

$$\% \text{ water of crystallisation} = \frac{5.1684}{8.20} \times 100$$

$$= 63.029 \% \quad (3)$$

TIP: To allow for rounding off of figures, any answer in the region 62.9 % to 63.2% was awarded full marks.

We now calculate the value of x in the formula $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$. We note from the above that by dissolving 8.20 g of washing soda in 500 cm^3 of solution, we obtain a solution which contains 0.0286 moles per 500 cm^3 .

$$\text{Therefore, } 0.0286 \text{ moles} = 8.20 \text{ g}$$

$$1 \text{ mole} = \frac{8.20}{0.0286}$$

$$= 286.713 \quad (3)$$

$$\text{Relative molecular mass of } \text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = 287$$

$$106 + 18x = 287 \quad (3)$$

$$18x = 287 - 106$$

$$= 181$$

$$x = 10 \text{ (x must be a whole number)} \quad (3)$$

Percentage of water of crystallisation = 63.029%

Value of $x = 10$

14 Oxidation and Reduction

The *Oxidation and Reduction* chapter is usually examined on the Leaving Certificate examination paper by means of short questions given as parts of full questions. The questions are usually based on the theory of oxidation and reduction as well as testing your understanding of oxidation numbers. In addition, since this chapter includes Mandatory Student Experiment No. 8 (to study some oxidation-reduction reactions), some questions on this experiment have also been asked on examination papers. Pay particular attention to the definitions highlighted in this chapter of *Chemistry Live!* as these definitions are the key to getting good marks on the examination paper. In addition, make sure you get plenty of practice in balancing redox equations as this task is frequently asked on exam papers.

2002 Question 4

- (e) What is the oxidation number of sulfur in $\text{Na}_2\text{S}_2\text{O}_3$?

Answer

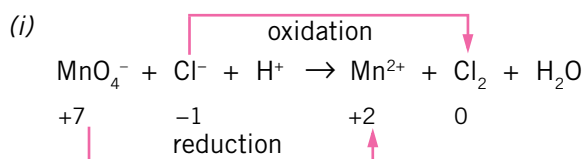
- (e) The oxidation number of sulfur is +2. (6)

2002 Question 10

- (a) Define *oxidation number*. (4)
- (i) Using oxidation numbers, identify which species is being oxidised and which is being reduced in the following reaction. (12)
- $$\text{MnO}_4^- + \text{Cl}^- + \text{H}^+ \rightarrow \text{Mn}^{2+} + \text{Cl}_2 + \text{H}_2\text{O}$$
- (ii) Hence, or otherwise, balance the equation. (9)

Answer

- (a) Oxidation number is the charge that an atom has or appears to have when electrons are distributed according to certain rules. (4)



Cl^- is being oxidised (6), MnO_4^- is being reduced (6).

TIP: Since the question specifies that you must use oxidation numbers, you must ensure that you clearly indicate the change that occurs in oxidation numbers as shown in the above equation. Simply stating that " Cl^- is oxidised and MnO_4^- is reduced" will get you only half marks.

- (ii) $2\text{MnO}_4^- + 10\text{Cl}^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 5\text{Cl}_2 + 8\text{H}_2\text{O}$ (9)

TIP: If the complete equation is correctly balanced, 9 marks will be awarded. However, you can obtain 3 or 6 marks if you correctly balance parts of the equation. Therefore, it is important to lay out your work as shown in Examples 14.9 and 14.10 on page 208 of your textbook. Showing these steps will help you to obtain the 3 or 6 marks in the marking scheme.

2004 Question 4

- (g) What is the oxidation number of (i) oxygen in H_2O_2 and (ii) bromine in KBrO_3 ? (6)

Answer

- (g) (i) The oxidation number of oxygen is -1 . (3)
 (ii) The oxidation number of bromine is $+5$. (3)

TIP: If you are unsure of the above answers, study the rules for assigning oxidation numbers on pages 197–198 of your textbook.

2005 Question 11

- (a) (i) Define *oxidation* in terms of change in oxidation number. (4)
 (ii) What is observed when chlorine gas is bubbled into an aqueous solution of sodium bromide? Explain your answer in terms of oxidation and reduction. (9)

Answer

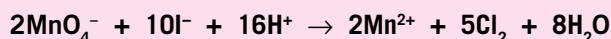
- (i) Oxidation is an increase in oxidation number. (4)
 (ii) The solution turns orange. [The colours red-brown, red or yellow are also acceptable in describing the colour of the solution.] (3)
 1. The bromide ions are oxidised to bromine. (3)
 2. The chlorine is reduced to chloride ions. (3)

TIP: You can also get full marks if you write the equation for the reaction:



2006 Question 10

- (b) Define oxidation in terms of change in oxidation number. (4)
 What is the oxidation number of (i) chlorine in NaClO and (ii) nitrogen in NO_3^- ? (6)
 State and explain the oxidation number of oxygen in the compound 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)



Answer

- (b) Oxidation is an increase in oxidation number. (4)
 (i) The oxidation number of chlorine is $+1$. (3)
 (ii) The oxidation number of nitrogen is $+5$. (3)
 The oxidation number of oxygen in OF_2 is $+2$. (3)
 The reason for this is because fluorine is more electronegative than oxygen. (3)
 $2\text{MnO}_4^- + 10\text{I}^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 5\text{I}_2 + 8\text{H}_2\text{O}$
 $\begin{array}{ccccccc} +7 & & -1 & & +2 & & 0 \end{array}$

14 Oxidation and Reduction

The reducing agent is I^- (since it reduces the Mn from an oxidation number of +7 to +2). (3)
The colour would change from purple (3) to red (3).

TIP: *Some flexibility was allowed in describing the colour of the iodine solution (because the colour depends on the concentration of the solution). Other acceptable colours are brown, orange or yellow.*

15 Volumetric Analysis: Oxidation-Reduction

The questions on the *Volumetric Analysis: Oxidation-Reduction* chapter usually appear on question 1 of the Leaving Certificate Higher Level Chemistry examination paper. There is always a full question (question 1) given on volumetric analysis. This question is based on either *Volumetric Analysis: Acid-Base* (Chapter 13) or *Volumetric Analysis: Oxidation-Reduction* (Chapter 15) or on some of the titrations related to water analysis (Chapter 19). In addition, some short questions given as part of question 1 may be based on material covered in other chapters of the textbook, such as acid-base indicators (Chapter 18) or the naming of carboxylic acids (Chapter 22).

2003 Question 1

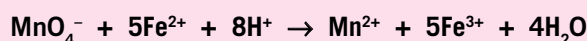
Iron tablets may be used in the treatment of anaemia.

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 is 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



- (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)

Answer

- (a) The sulfuric acid prevents the Fe²⁺ ions being oxidised to Fe³⁺ ions by the oxygen in the air or oxygen dissolved in the water. (5)
- (b) The procedure for making up the 250 cm³ of solution from the tablets may be summarised as follows:
 - ⊗ Grind up the tablets using a pestle and mortar.
 - ⊗ Add dilute sulfuric acid to the mortar and wash into a beaker of dilute sulfuric acid.
 - ⊗ Stir the solution to dissolve the tablets.
 - ⊗ Using a funnel and glass rod, transfer the solution to a volumetric flask.
 - ⊗ Transfer all the washings from the beaker to the volumetric flask.
 - ⊗ Make sure that the flask is on a level surface and that your eye is level with the mark on the volumetric flask.

15 Volumetric Analysis: Oxidation–Reduction

- ⊗ Bring the bottom of the meniscus up to the mark with deionised water.
- ⊗ When bringing the meniscus up to the mark, add the water drop by drop.
- ⊗ Invert the stoppered flask about twenty times.

Any six of these nine points: (6 × 3)

- (c) More sulfuric acid is added to ensure that there are enough H^+ ions to help completely reduce the MnO_4^- ions to Mn^{2+} ions. If the sulfuric acid is not added, a brown colour is observed due to the formation of manganese dioxide. (6)
- (d) The end point is detected by the observation of a faint permanent pink colour. (3)

(e) Given:

$$\begin{array}{l} V_o = 13.9 \\ M_o = 0.010 \\ n_o = 1 \\ V_r = 25 \\ M_{\text{red}} = ? \\ n_r = 5 \end{array} \quad \begin{array}{c} \text{MnO}_4^- \quad \text{Fe}^{2+} \\ \frac{V_o \times M_o}{n_o} = \frac{V_r \times M_{\text{red}}}{n_r} \\ \frac{13.9 \times 0.010}{1} = \frac{25 \times M_{\text{red}}}{5} \quad (3) \\ M_{\text{red}} = \frac{13.9 \times 0.010 \times 5}{25} \\ = 0.0278 \text{ moles per litre (3)} \end{array}$$

(Rel molecular mass of $\text{FeSO}_4 = 152$)

$$\begin{aligned} &= 0.0278 \times 152 \text{ grams per litre} \\ &= 4.22 \text{ grams per litre} \end{aligned}$$

Therefore, in the 250 cm^3 volumetric flask there are $\frac{4.22}{4} = 1.0564$ grams of FeSO_4

Thus, since we used four tablets, a single tablet contains $\frac{1.0564}{4}$ grams of $\text{FeSO}_4 = 0.2641$ grams

The percentage of iron in FeSO_4 is $\frac{56}{152} \times 100 = 36.84 \%$

Therefore, mass of iron in each tablet is 36.84% of $0.2641 \text{ g} = 0.0973 \text{ g}$ (9)

Mass of one tablet = 0.36 g

% Fe in each tablet = $\frac{0.0973}{0.36} \times 100 = 27.027\%$ (3)

Answers:

- (i) Concentration of iron(II) solution = 0.0278 moles per litre
- (ii) Mass of iron(II) in one tablet = 0.0973 g
- (iii) Percentage by mass of iron(II) in each tablet = 27%

16 Rates of Reactions

The questions on the *Rates of Reactions* chapter often appear as full questions on the Leaving Certificate Higher Level Chemistry examination paper. Recall that there are two mandatory experiments in this chapter and therefore a full question can appear in either Section A or in Section B of the examination paper. In some cases, this chapter is examined in the form of short questions given as parts of full questions.

2002 Question 3

To investigate the effect of concentration on a reaction rate, a student measured 100 cm³ of a 0.10 M solution of sodium thiosulfate into a conical flask, added 10 cm³ of 1.0 M hydrochloric acid, and then placed the flask on top of a cross on a sheet of white paper as shown in the diagram. The student noted the time (in minutes) taken for the cross to become obscured by the pale yellow precipitate formed in the solution. The reciprocal of the time (1/time) was used as a measure of the initial rate of the reaction.

Samples of the 0.10 M solution of sodium thiosulfate were diluted to make 100 cm³ portions of 0.08, 0.06, 0.04 and 0.02 M sodium thiosulfate. Each of these was, in turn, reacted with 10 cm³ of 1.0 M hydrochloric acid as described above. The results obtained are shown in the following table.

Concentration of sodium thiosulfate solution (M)	Time taken for the cross to become obscured (minutes)	1/time (min ⁻¹) i.e. Rate
0.10	1.25	0.80
0.08	1.56	0.64
0.06	2.08	0.48
0.04	3.13	0.32
0.02	6.25	0.16



- Identify the pale yellow precipitate that obscured the cross on the sheet of paper. (5)
- Describe the procedure for preparing the 0.08 M solution of sodium thiosulfate from the 0.10 M solution. (12)
- Plot a graph to show the relationship between the initial rate of this reaction (1/time) and the concentration of the sodium thiosulfate solution. What conclusion can be drawn from the graph about the relationship between the rate of reaction and the concentration of the sodium thiosulfate? (18)
- Use the graph to determine how long it would have taken for the cross on the sheet of paper to become obscured if the student had used a 0.05 M sodium thiosulfate solution. (9)
- Explain why the reciprocal of the time (1/time) may be used as a measure of the initial rate of the reaction. (6)

Answer

- Sulfur (5)
- The procedure for preparing the 0.08 M solution of sodium thiosulfate from the 0.1 M solution may be summarised in the following steps:
 - ⊗ Use a graduated cylinder.
 - ⊗ Place 80 cm³ of the 0.1 M sodium thiosulfate solution in the conical flask.
 - ⊗ Add 20 cm³
 - ⊗ Of deionised water
 (4 × 3)

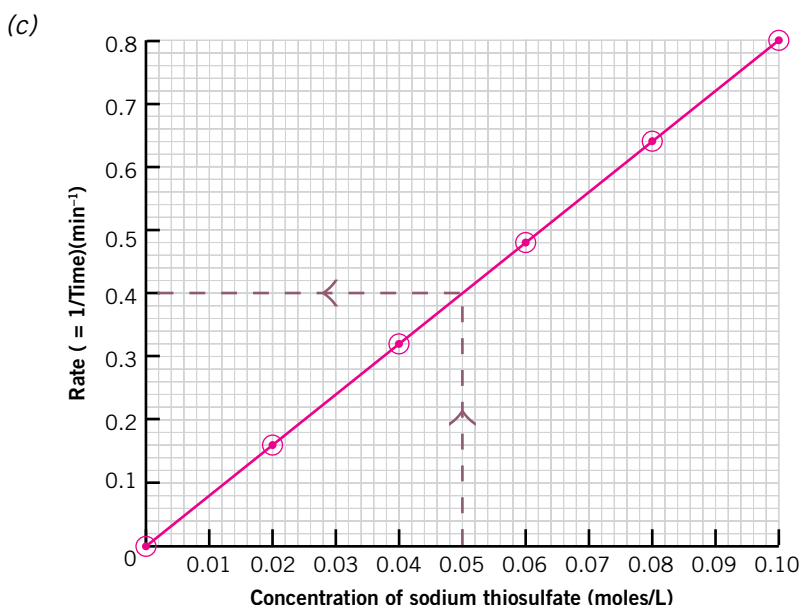


Fig. 16.1

The main points from the above graph are as follows:

Axes correct (3), correctly labelled (3), points correctly plotted (3), straight line through the origin (3).

A straight line through the origin is obtained showing that the rate of reaction is directly proportional to the concentration of the sodium thiosulfate solution. (6)

(d) From the graph, the rate corresponding to a 0.05 M sodium thiosulfate solution is 0.4 min⁻¹. (3)

$$\Rightarrow \frac{1}{t} = 0.4 \quad (3)$$

$$\Rightarrow t = 2.5 \text{ minutes} \quad (3)$$

(e) Time and rate are inversely proportional (e.g. when the rate is doubled, the time is halved).

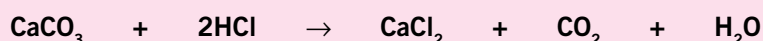
$$\text{Rate} \propto \frac{1}{\text{time}}$$

The shorter the length of time taken for the cross to be obscured, the faster the rate of reaction. The longer the time it takes the cross to be obscured, the slower the rate of reaction. (6)

2003 Question 7

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

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



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)

Answer

(a) The rate of reaction is defined as the change in concentration per unit time of any one reactant or product. (5)

(i) Effect of particle size on rate:

- Place a conical flask containing about 20 g of **large** marble chips and a graduated cylinder containing about 50 cm³ of dilute hydrochloric acid on the top of the electronic balance. Add the HCl to the marble chips. (3)

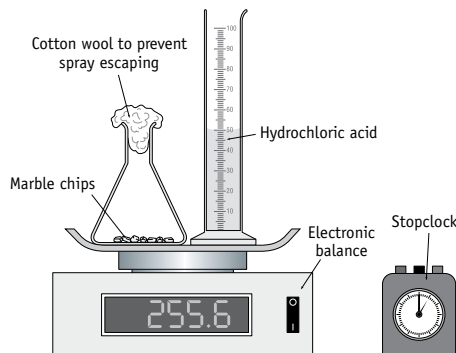


Fig. 16.2

- Note the time it takes for the reaction to finish. (3)
- Repeat the experiment, but this time use the same mass of **small** marble chips and add the same volume of HCl of the same concentration. It is observed that the smaller particles react faster. (3)

TIP: Full marks for the results of this experiment can be obtained from a graph as shown in Fig. 16.3.

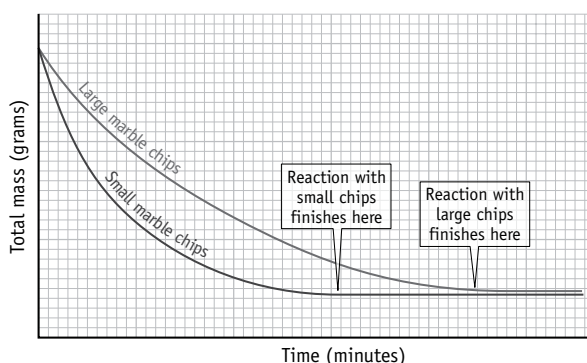


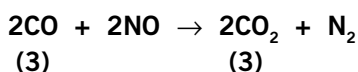
Fig. 16.3

(ii) Effect of particle size on concentration:

- Place a conical flask containing about 20 g of marble chips and a graduated cylinder containing about 50 cm³ of dilute hydrochloric acid on the top of the electronic balance. Add the HCl to the marble chips. (3)
- Note the time it takes for the reaction to finish. (3)
- Repeat the experiment, but this time add the same volume of HCl of different concentrations to the same mass of equal-sized marble chips. It is observed that the reaction is faster with the more concentrated hydrochloric acid. (3)

(b) A catalyst is a substance that alters the rate of a chemical reaction (3), but is not consumed in the reaction (3).

(i) The main reaction which happens in the catalytic converter is the reaction between the polluting gases themselves (carbon monoxide and nitrogen monoxide) to form carbon dioxide and nitrogen:



16 Rates of Reactions

The environmental benefit is that CO which is poisonous, and NO which leads to acid rain, are being removed by the catalytic converter. (6)

TIP: You did not need to reproduce the above equation in answering this question for the examination. Full marks were given if you could explain in words what is happening.

(ii) Platinum (3)

[Palladium and rhodium are also acceptable answers.]

The type of catalysis involved is heterogeneous catalysis. (6)

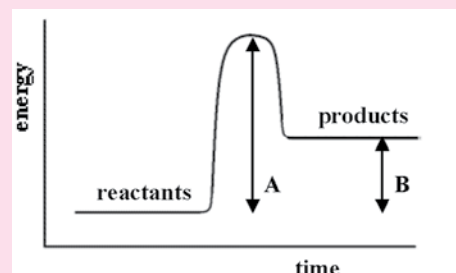
2004 Question 8

(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)

Answer

(a) The rate of reaction is defined as the change in concentration per unit time of any one reactant or product. (5)

Reactions go faster at higher temperatures because more of the colliding molecules have the minimum activation energy needed to react. (6)

(b) A = Activation Energy (3)

B = Heat of reaction (ΔH) (3)

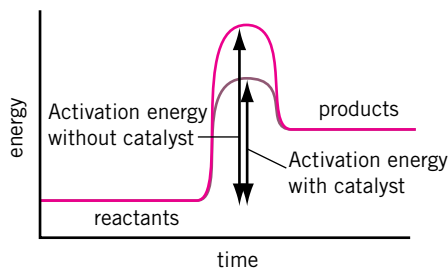


Fig. 16.4

The key points in the above diagram are as follows:

⊗ The diagram shows a reduction in the activation energy. (3)

⊗ The products line is shown above the reactants line. (3)

(c) Platinum and Palladium (2 × 3)

[Rhodium is also an acceptable answer.]

Lead poisons the catalysts in a catalytic converter. (3)

(d) The observations made are as follows:

- ⊗ The solution is pink at the start of the reaction.
 - ⊗ The solution changes to a green colour.
 - ⊗ Fizzing is observed.
 - ⊗ The solution turns pink again.
- (4 × 3)

The colour change from pink to green indicates the formation of a new substance and the color changing back to its original pink colour suggests that the new substance is an intermediate. (6)

2005 Question 3

Hydrogen peroxide decomposes rapidly in the presence of a manganese(IV) oxide (MnO_2) 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 (MnO_2) 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.

Time/minute	0	1	2	3	4	5	6	7	8
Volume of O_2/cm^3	0.0	13.5	23.4	30.5	35.4	38.3	39.6	40.0	40.0

(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)

Answer

(a) $2\text{H}_2\text{O}_2 \longrightarrow 2\text{H}_2\text{O} + \text{O}_2$ (5)

(b) The key points from the diagram are as follows:

- ⊗ Conical flask with hydrogen peroxide and manganese dioxide. (3)
- ⊗ Delivery tube connected to gas collection system over water (or gas syringe). (3)
- ⊗ Inverted graduated cylinder with graduations marked. (3)

To start the reaction at an exact time, the manganese dioxide is added to the hydrogen peroxide solution, the flask is immediately stoppered and the clock started at the same time (3).

Alternatively, the catalyst could be shown in a small test-tube inside the conical flask.

Note: A common mistake made by students in drawing the diagram to illustrate this reaction is to include a dropping funnel. This is not correct as the start time is not exact and the volume collected is incorrect due to displacement of air.

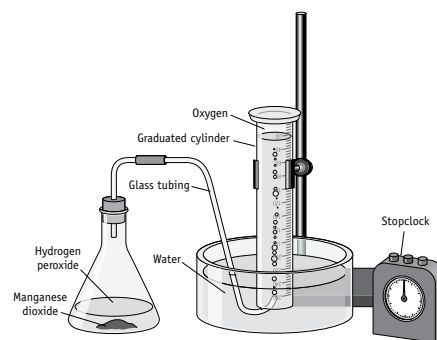


Fig. 16.5(a)

OR

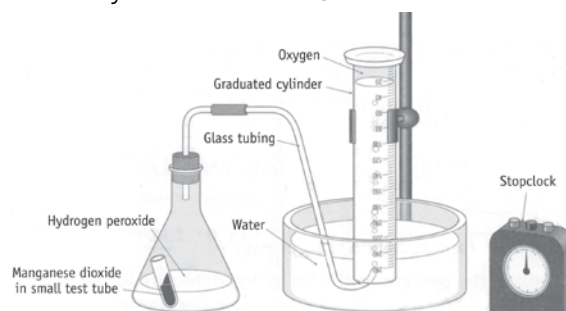


Fig. 16.5(b)

16 Rates of Reactions

- (c) The finely powdered manganese dioxide would give a faster rate of reaction (3), as there is a greater surface area available (3).
- (d) The main points to be included in the graph Fig. 16.6 are as follows:
- ⊗ The axes must be labelled, and the scales shown. (3)
 - ⊗ The points must be correctly plotted using graph paper. (6)
 - ⊗ The curve must be drawn to (0,0). (3)
- (e) (i) From graph Fig. 16.6, the volume of oxygen produced during first 2.5 mins = 27 cm³ (3)
[Any answer in the range 26.5 cm³–28.5 cm³ is acceptable.]

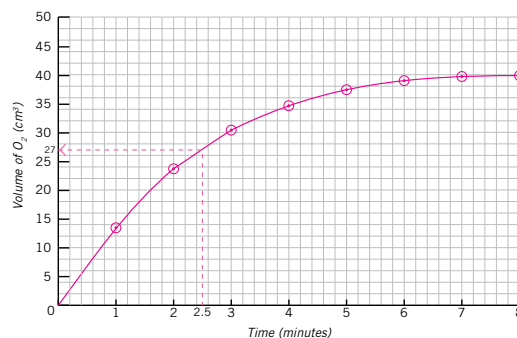


Fig. 16.6

Note: For clarity, two graphs are used to show the required answers for this question. However, in the Leaving Certificate examination, everything can be shown in one graph.

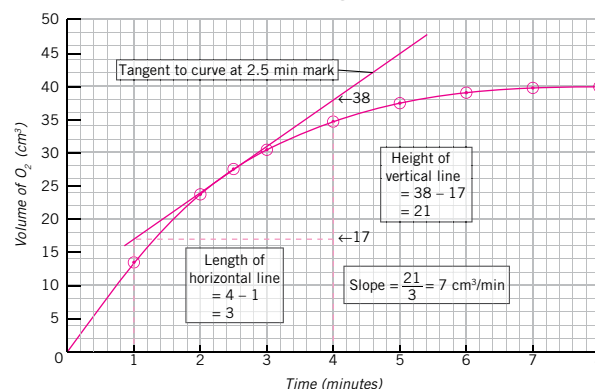


Fig. 16.7

- (ii) From Fig. 16.7 we see that the instantaneous rate of reaction at 2.5 minutes = $\frac{21}{3} = 7$ cm³/min (6)
(Any answer in the range 6.0 – 8.0 cm³/min is acceptable.)
- (f) The rise in the graph would be less steep, i.e. it would take longer to level off. (3)
There would be less oxygen produced. We would expect only half the final volume of oxygen to be produced (20 cm³). (3)

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.
- $$\text{Na}_2\text{S}_2\text{O}_3 + 2\text{HCl} \rightarrow 2\text{NaCl} + \text{SO}_2 + \text{S} + \text{H}_2\text{O}$$
- (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)

Answer

- (a) The Activation Energy is the minimum energy which colliding particles must have for a reaction to occur. (5)

- (b) The rate of a chemical reaction increases as the temperature rises because of the following:
- ⊗ There is an increased number of collisions since the moving molecules have more energy. (3)
 - ⊗ More of the collisions are effective, i.e. the colliding molecules have the minimum Activation Energy needed to react. (3)

The more significant reason is the second one above. (3)

The reason why this is more significant is because if molecules with low energies collide, they simply bounce apart without any reaction taking place. Thus, an increase in the number of collisions will not necessarily lead to the reaction taking place. The important thing is that the rise in temperature must be great enough to ensure that the colliding molecules have enough energy to overcome the Activation Energy. (3)

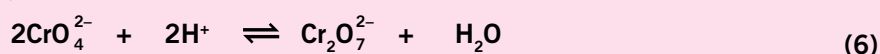
- (c) The main points in this experiment are as follows:
- ⊗ Pour 100 cm³ of the sodium thiosulfate solution into a conical flask. Place the conical flask on a sheet of white paper on which a mark has been made. Pour 10 cm³ of the hydrochloric acid solution into the conical flask, swirl the flask and start the stopclock.
 - ⊗ Look at the cross through the solution and when the degree of “miliness” reaches such a stage that the cross is no longer visible, record the time taken.
 - ⊗ Heat the same volumes of the solutions to a different temperature.
 - ⊗ Repeat the experiment at different temperatures of solutions. The hotter solution takes less time to obscure the cross, i.e. increased temperature causes an increased reaction rate. (4 × 3)
- (d) (i) When the solutions of silver nitrate and sodium chloride are mixed, the precipitate appears immediately because the ions in solution simply collide with each other. No bond breaking is needed. However, the reaction between sodium thiosulfate and hydrochloric acid is much slower because it takes time for the covalent bonds to be broken. (6)
- (e) The type of catalysis is heterogeneous catalysis. (3)
Carbon monoxide and nitrogen monoxide enter the catalytic converter of the car. (2 × 3)
The carbon monoxide is converted to carbon dioxide. (3)
The nitrogen monoxide is converted to nitrogen. (3)

17 Chemical Equilibrium

The questions on the *Chemical Equilibrium* chapter often appear as full questions on the Leaving Certificate Higher Level Chemistry examination paper. In some examination papers, only parts of questions (usually either question 10 or question 11) have examined the topic of equilibrium. Recall that there is one mandatory experiment in this chapter (Mandatory Student Experiment No. 15) and therefore a full question can be given in either Section A or Section B of the examination paper. On rare occasions, a short question on equilibrium has been given as part of question 4.

2002 Question 4

- (g) What colour change will occur if concentrated sulfuric acid is added to the following equilibrium mixture? Give a reason for your answer.



Answer

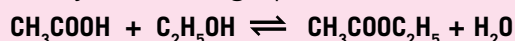
- (g) The colour change will be from yellow to orange. (3)
In order to relieve the stress of the added H^+ ions, the equilibrium shifts from left to right. (3)

2002 Question 10

- (c) State *Le Chatelier's principle*. (7)

When 30 g of ethanoic acid and 23 g of ethanol were placed in a conical flask and a few drops of concentrated sulfuric acid added, an equilibrium was set up with the formation of ethylethanoate and water.

The equilibrium is represented by the following equation.

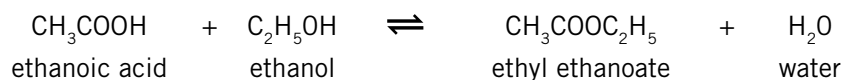


When the equilibrium mixture was analysed it was found to contain 10 g of ethanoic acid.

- (i) Write the equilibrium constant expression, K_c , for this reaction. (6)
(ii) Calculate the value of the equilibrium constant K_c . (12)

Answer

- (c) Le Chatelier's Principle states that if a stress is applied to a system at equilibrium (3), the system readjusts to relieve the stress applied (4).
(i)



$$K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]}$$

(6)

TIP: You must ensure that you use square brackets for the K_c expression or else you will lose all of the marks.

(ii) First convert grams to moles.

Initially,

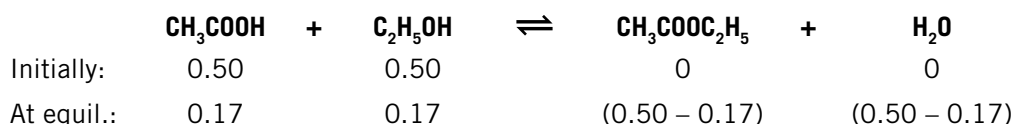
$$\text{no. of moles ethanoic acid} = \frac{\text{Mass}}{\text{Rel Mol mass}} = \frac{30}{60} = 0.50$$

$$\text{no. of moles ethanol} = \frac{\text{Mass}}{\text{Rel Mol mass}} = \frac{23}{46} = 0.50$$

At equilibrium,

$$\text{no. of moles ethanoic acid} = \frac{\text{Mass}}{\text{Rel Mol mass}} = \frac{10}{60} = 0.166 = 0.17$$

Therefore, we can write the equation:



Therefore, the equilibrium concentrations are as follows:



V is the volume of the container. We are not given it in the question as the V terms cancel out in the K_c expression.

$$K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]} = \frac{(0.33) \times (0.33)}{(0.17) \times (0.17)} = 3.77 \quad (3)$$

Answer: $K_c = 3.77$

2003 Question 11

(a) State *Le Chatelier's principle*. (7)

A gaseous mixture of hydrogen, iodine and hydrogen iodide form an equilibrium according to the following equation.

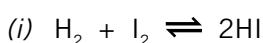


(i) Write an expression for the equilibrium constant, K_c , for this system. (6)

(ii) The value of the equilibrium constant, K_c , for this reaction is 50 at 721 K. If 2 moles of hydrogen iodide gas were introduced into a sealed vessel at this temperature calculate the amount of hydrogen iodide gas present when equilibrium was reached. (12)

Answer

(a) Le Chatelier's Principle states that if a stress is applied to a system at equilibrium (3), the system readjusts to relieve the stress applied (4).



$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \quad (6)$$

17 Chemical Equilibrium

(ii) Let x = number of moles of H_2 and I_2 formed.

	H_2	+	I_2	\rightleftharpoons	$2HI$	
Initially:	[0]		[0]		[2]	
At equil.:	x		x		$2 - 2x$	
Conc. at equil.:	$[x]$		$[x]$		$[2 - 2x]$	(3)

$$K_c = \frac{[HI]^2}{[H_2][I_2]}$$

$$\frac{(2 - 2x)^2}{x^2} = 50 \quad (3)$$

$$\frac{2 - 2x}{x} = 7.07 \quad (3)$$

Solving $x = 0.22$

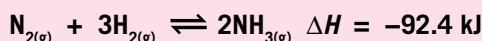
Therefore, equilibrium concentration of $[HI] = 2 - 2x = 2 - 0.44 = 1.56$ moles (3)

Answer: 1.56 moles.

2004 Question 9

(a) What is meant by *chemical equilibrium*? Why is it described as a *dynamic state*? (8)

Consider the following reversible chemical reaction:



(b) Use Le Chatelier's principle to predict the levels (high or low) of temperature and pressure needed to maximise the yield of ammonia when equilibrium is established. Give a reason (i) for the temperature level you have predicted, (ii) for the level of pressure you have predicted. (12)

(c) Are the temperature levels predicted using Le Chatelier's principle actually used to maximize ammonia yield in industry? Explain your answer. (6)

(d) What is the effect of a catalyst on a reversible reaction? (6)

(e) In an experiment 6.0 moles of nitrogen and 18.0 moles of hydrogen were mixed and allowed to come to equilibrium in a sealed 5.0 litre vessel at a certain temperature. It was found that there were 6.0 moles of ammonia in the equilibrium mixture.

Write the equilibrium constant expression for the reaction and calculate the value of the equilibrium constant (K_c) at this temperature. (18)

Answer

(a) Chemical equilibrium is a state of dynamic balance where the rate of the forward reaction equals the rate of the reverse reaction. (5)

It is called a dynamic (moving) state because the reactants are continually forming products and the products are continually forming reactants. (3)

(b) $N_2 + 3H_2 \rightleftharpoons 2NH_3$

Low temperature is needed to maximise the yield of ammonia. (3)

High pressure is needed to maximise the yield of ammonia. (3)

Exam Edge

- (i) Low temperature is needed since the forward reaction is exothermic so the low temperature would absorb this heat. (3)
- (ii) High pressure is needed since there are fewer molecules on the right hand side of the equation than there are on the left, so the high temperature would push the equilibrium to the right hand side. (3)
- (c) No, the temperature levels predicted are not used to maximise ammonia yields in industry. (3)
The reason for this is because if the temperature is too low, the rate of reaction is too slow. (3)
- (d) The effect of a catalyst on a reversible reaction is that it brings the system to equilibrium more quickly. This increases the productivity of the plant and makes the process more economical. (6)
- (e)

	N_2	+	3H_2	\rightleftharpoons	2NH_3	
Initially:	6 moles		18 moles		0 moles	
	$6 - x$		$18 - 3x$		$2x = 6 \Rightarrow x = 3$	
At equil.:	3 moles		9 moles		6 moles	(3)
Conc. at equil. ($V = 5$):	[0.6]		[1.8]		[1.2]	(3)

$$K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

$$= \frac{(1.2)^2}{(0.6)(1.8)^3}$$

$$= 0.41$$

Answer: $K_c = 0.41$

Note: In the Leaving Certificate examination, quite a number of students made an error in working out the number of moles at equilibrium in this problem. Recall that we are told that 6 moles of ammonia are formed at equilibrium. From the balanced equation, this must mean that 3 moles of nitrogen (half the number of moles of ammonia) are present at equilibrium. Since the balanced equation tells us that there are three times as many moles of hydrogen at equilibrium as nitrogen, this must mean that there are 9 moles of hydrogen present at equilibrium.

2005 Question 9

- (a) State *Le Chatelier's principle*. (5)
- (b) A student is provided with glassware and other laboratory apparatus as well as the following chemicals: potassium dichromate(VI) ($\text{K}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$), hydrochloric acid ($\text{HCl}_{(\text{aq})}$), sodium hydroxide (NaOH), cobalt(II) chloride crystals ($\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$) and deionised water (H_2O).
- (i) Describe clearly how the student could use a selection of the chemicals listed above to establish a chemical equilibrium. Write a balanced equation for the equilibrium. (12)
- (ii) Describe how the student could then demonstrate the effect of concentration on that chemical equilibrium. State the observations made during the demonstration. (9)
- (c) The value of K_c for the following equilibrium reaction is 4.0 at a temperature of 373 K.
- $$\text{CH}_3\text{COOH} + \text{C}_2\text{H}_5\text{OH} \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O}$$
- (i) Write the equilibrium constant (K_c) expression for this reaction. (6)
- (ii) What mass of ethyl ethanoate ($\text{CH}_3\text{COOC}_2\text{H}_5$) would be present in the equilibrium mixture if 15 g of ethanoic acid and 11.5 g of ethanol were mixed and equilibrium was established at this temperature? (18)

Answer

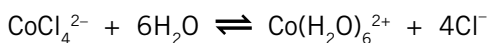
(a) Le Chatelier's Principle states that if a stress is applied to a system at equilibrium (2), the system readjusts to relieve the stress applied (3).

(b) (i) Either Example 1 or Example 2 below is acceptable for full marks (4 × 3):

Example 1

Dissolve the cobalt chloride (3) in water (3).

The following equilibrium is set up:

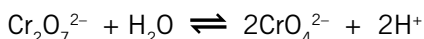


Correct formulas (3); Equation balanced (3)

Example 2

Dissolve the potassium dichromate (3) in water (3).

The following equilibrium is set up:



Correct formulas (3); Equation balanced (3)

(ii) **Example 1**

Add hydrochloric acid (3). The colour of the solution changes from red (3) to blue (3).

Example 2

Add sodium hydroxide (3). The colour of the solution changes from orange (3) to yellow (3).

$$(c) \quad (i) \quad K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]} \quad (6)$$

(ii) Convert grams to moles.

$$\text{no. of moles ethanoic acid} = \frac{\text{Mass}}{\text{Rel Mol mass}} = \frac{15}{60} = 0.25$$

$$\text{no. of moles ethanol} = \frac{\text{Mass}}{\text{Rel Mol mass}} = \frac{11.5}{46} = 0.25$$

Let x = number of moles of ethyl ethanoate formed

	CH_3COOH	+	$\text{C}_2\text{H}_5\text{OH}$	\rightleftharpoons	$\text{CH}_3\text{COOC}_2\text{H}_5$	+	H_2O	
Initially	0.25		0.25		0		0	(3)
At equil.:	$0.25 - x$		$0.25 - x$		x		x	(3)
Conc. at equil.:	$[0.25 - x]$		$[0.25 - x]$		[x]		[x]	

$$K_c = \frac{[\text{CH}_3\text{COOC}_2\text{H}_5][\text{H}_2\text{O}]}{[\text{CH}_3\text{COOH}][\text{C}_2\text{H}_5\text{OH}]} = 4$$

$$\text{i.e.} \quad \frac{(x)^2}{(0.25 - x)^2} = 4 \quad (3)$$

Taking the square root of both sides:

$$\frac{x}{0.25 - x} = 2 \quad (3)$$

$$x = 0.5 - 2x$$

$$3x = 0.5$$

$$x = 0.166 \quad (3)$$

Exam Edge

i.e. moles of ethyl ethanoate at equilibrium = 0.17 (to two decimal places)

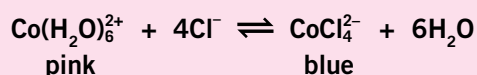
$$= 0.17 \times 88$$
$$= 14.96 \text{ g} \quad (3)$$

TIP: Full marks were awarded for any answer between 14.08 and 14.96 as the actual answer depended on the amount of rounding off, i.e. 0.16 or 0.166 or 0.17.

2006 Question 11

- (b) State *Le Chatelier's principle*. (7)

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



- (i) When the solution becomes blue, has the 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)

Answer

- (b) If a stress is applied to a system at equilibrium (4), the system readjusts to relieve the stress applied (3).
- (i) No, the reaction has not ceased when the solution becomes blue (3). The reason for this is because the forward and reverse reactions are still taking place, i.e. the reactants are changing to products and the products are changing to reactants (3).
- (ii) When the reaction mixture is cooled, the colour changes from blue to pink (3). This happens because the equilibrium is shifted to the left since this direction is exothermic, and hence the stress is relieved on cooling the mixture (3).
- (iii) One way of reversing the change caused by cooling would be to add a source of chloride ions, e.g. add HCl or NaCl (6). [Another way would be to remove water.]

18 pH and Indicators

Questions on the *pH and Indicators* chapter have appeared as part of question 4 and often as part of the question on water (Chapter 19). It is advisable to revise Chapter 12 (Acids and Bases) in conjunction with your revision of this chapter as it is important to understand the concepts of strong and weak acids and bases as well as the concept of conjugate acids and bases. The definition of pH, which is best stated in the form of the mathematical formula, is essential – this formula is central to all of the calculations in this chapter. You are required to demonstrate how the pH of neutral water at 25 °C is calculated as 7. The calculation of the pH of strong acids or bases is straightforward while, for weak acids and bases, the calculation is slightly more challenging. The formulas (given on page 291 of the textbook) for calculating $[H^+]$ for a weak acid and $[OH^-]$ for a weak base should be memorised. You will need to be able to use the idea of chemical equilibrium, referring to Le Chatelier's Principle, to explain how an acid-base indicator functions. You must understand how to choose the correct indicator for a particular titration and be able to explain your choice with reference to the data from titration curves.

2002 Question 4

- (d) The value of the dissociation constant for ethanoic acid is $1.8 \times 10^{-5} \text{ l mol}^{-1}$.
Calculate the pH of a 0.01 M solution of ethanoic acid.

(6)

Answer

$$\begin{aligned} (d) [H^+] &= \sqrt{K_a \times M_a} \\ &= \sqrt{(1.8 \times 10^{-5})(0.01)} \\ &= 0.000424 \\ \text{pH} &= -\log_{10}[H^+] \\ &= -\log(0.000424) \\ &= 3.37 \end{aligned}$$

(3)

(3)

2003 Question 8

- (b) Define pH. (6)
A bottle of vinegar is labelled 6% (w/v) acetic acid (ethanoic acid). The dissociation constant, K_a , for ethanoic acid is 1.8×10^{-5} . Calculate the approximate pH of the vinegar solution. (12)

Answer

(b) $\text{pH} = -\log_{10}[H^+]$ (6)

Note: The square brackets indicate concentration in moles per litre.

$$\begin{aligned} 6\% \text{ (w/v)} &= 6 \text{ g of ethanoic acid in } 100 \text{ cm}^3 \\ &= 60 \text{ g per litre} \\ M_r \text{ of } \text{CH}_3\text{COOH} &= (2 \times 12) + (2 \times 16) + (4 \times 1) = 60 \\ \Rightarrow \text{Concentration of ethanoic acid (} M_a \text{)} &= 1 \text{ mole per litre} \\ [H^+] &= \sqrt{K_a \times M_a} \\ &= \sqrt{(1.8 \times 10^{-5})(1)} \\ &= 0.00424 \end{aligned}$$

(3)

(3)

(3)

$$\begin{aligned}
 \text{pH} &= -\log_{10}[\text{H}^+] \\
 &= -\log(0.00424) \\
 &= 2.37
 \end{aligned}
 \tag{3}$$

2004 Question 4

(e) Write (i) the conjugate acid and (ii) the conjugate base of HPO_2^{4-} . (6)

Answer

(i) The conjugate acid is H_2PO_4^- . (3)

(ii) The conjugate base is PO_4^{3-} . (3)

2004 Question 11

(b) Define pH. (7)

(i) What are the limitations of the pH scale? (6)

(ii) Calculate the approximate pH of a vinegar solution that contains 4.5 g of ethanoic acid per 100 cm^3 . The value of K_a for ethanoic acid is 1.8×10^{-5} . (12)

Answer

(b) $\text{pH} = -\log_{10}[\text{H}^+]$ (7)

[Alternatively: The pH of a solution is the negative logarithm to the base 10 of the hydrogen ion concentration measured in moles per litre].

(i) The pH scale does not work with concentrated solutions **(3)** as it is found that once the concentration goes above 1M, complete dissociation does not always occur. Therefore, calculations of pH for solutions of concentrations greater than 1M are not accurate. In addition, pH is limited to the 0 – 14 range **(3)** [and only to solutions in water].

(ii) 4.5 grams per 100 $\text{cm}^3 = 45$ grams per litre (3)

$$M_r \text{ of } \text{CH}_3\text{COOH} = (2 \times 12) + (2 \times 16) + (4 \times 1) = 60$$

$$\Rightarrow \text{Number of moles of ethanoic acid per litre} = \frac{45}{60}$$

$$= 0.75 \text{ mol/litre} \tag{3}$$

$$\text{i.e. } M_a = 0.75 \text{ mol/L}$$

$$[\text{H}^+] = \sqrt{K_a \times M_a}$$

$$= \sqrt{(1.8 \times 10^{-5})(0.75)} \tag{3}$$

$$= 0.00367$$

$$\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.00367)$$

$$= 2.43 \tag{3}$$

Answer: $\text{pH} = 2.43$

2005 Question 8

- (b) Identify **one** species acting as an acid, and also identify its conjugate base, in the following system.



- (c) Calculate the pH of a 0.002 M solution of methanoic acid (**HCOOH**).

The value of K_a for methanoic acid is 1.8×10^{-4} . (12)

Answer

- (b) One species acting as an acid is H_2F^+ . (3)

[Alternative is HCl]

The conjugate base of H_2F^+ is HF.

(3)

[Alternatively, the conjugate base of HCl is Cl^- .]

- (c) $[\text{H}^+] = \sqrt{K_a \times M_a}$ (3)

$$= \sqrt{(1.8 \times 10^{-4})(0.002)}$$

$$= 0.0006$$

(3)

$$\text{pH} = -\log_{10} [\text{H}^+]$$

$$= -\log_{10} (0.0006)$$

$$= 3.22$$

(6)

2006 Question 4

- (e) What is (i) the conjugate acid and (ii) the conjugate base of **H_2O** ? (6)

- (h) The concentration of an aqueous solution of sodium hydroxide (**NaOH**) is 0.2 g per litre.

Calculate its pH. (6)

Answer

- (e) (i) The conjugate acid is H_3O^+ . (3)

(ii) The conjugate base is OH^- .

(3)

- (h) M_r of NaOH = 23 + 16 + 1 = 40

$$\Rightarrow [\text{NaOH}] = \frac{0.2}{40}$$

$$= 0.005 \text{ mol/L}$$

(3)

NaOH is a strong base $\Rightarrow [\text{NaOH}] = [\text{OH}^-] = 0.005$

$$\text{pOH} = -\log_{10} [\text{OH}^-]$$

$$= -\log_{10} 0.005$$

$$= 2.3$$

$$\text{pH} + \text{pOH} = 14$$

$$\Rightarrow \text{pH} = 14 - 2.3 = 11.7$$

(3)

2006 Question 8

- (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 cm³ of a 0.01 M sodium hydroxide (**NaOH**) solution is added slowly to 25 cm³ of a 0.1 M ethanoic acid (**CH₃COOH**) 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)

Answer

- (i) The indicator dissociates in water as follows:



If H^+ ions from an acid are added to the indicator in accordance with Le Chatelier's Principle, the above equilibrium shifts from right to left (3) in order to absorb the added H^+ ions. Therefore, the colour of the HX predominates. (3)

If a base containing OH^- ions is added to the indicator the OH^- ions react with the H^+ ions in the above equilibrium reaction to remove them to form water. In accordance with Le Chatelier's Principle, more H^+ ions must be formed and therefore the equilibrium is shifted to the right (3). Thus, the colour of the X^- predominates.

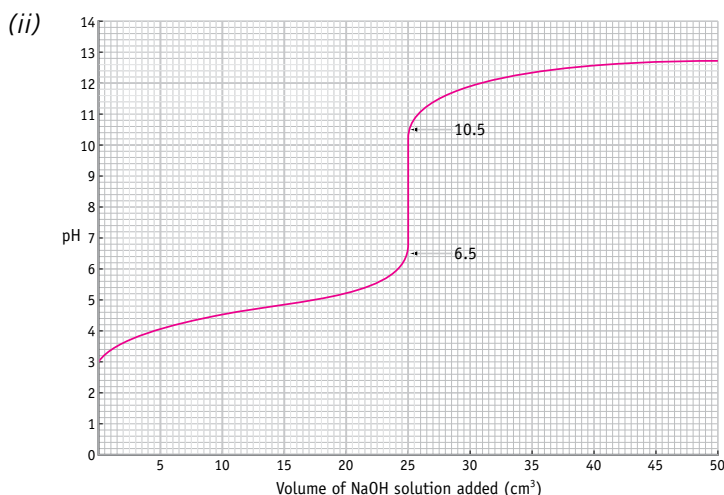


Fig. 18.1

Note: The main points to note in the diagram are as follows:

- ⊗ This is a titration between a weak acid and a strong base, so the end point will occur between approx. pH 6.5 to 10.5 (shown on the graph, 3 marks)
- ⊗ The concentrations of the acid and base are identical, so the end point will occur when 25 cm³ of the base have been added. (shown on the graph, 3 marks)
- ⊗ The other 3 marks are obtained for placing relevant pH values on the vertical axis of the graph.

- (iii) Phenolphthalein is colourless at a pH of 7 and pink at a pH of 10. Thus, the colour change of phenolphthalein takes place between pH 8.3 and pH 10.0. This range lies within the steep part of the graph. (6)

19 Environmental Chemistry – Water

The questions on the *Environmental Chemistry – Water* chapter often appear as full questions on the Leaving Certificate Higher Level Chemistry examination paper. Since the contents of this chapter are also related to pH (Chapter 18), parts of this chapter have also been given as sections of full questions on pH. Recall that there are four mandatory experiments in this chapter and therefore a full question can appear in either Section A or Section B of the examination paper. In some cases, this chapter is examined as one of the three parts in question 10 on the examination paper.

2002 Question 9

- (a) What property of water makes it very useful in the human body as a medium in which chemical reactions occur, and allows it to become polluted or contaminated in other situations? (5)
- The treatment of a water supply for domestic use may involve several stages.
- (i) These stages may involve *sedimentation*, *flocculation*, and *filtration*.
Describe what happens at each of these three stages. (18)
- (ii) Various chemicals are often added in other stages of water treatment.
Identify one other stage in water treatment which involves the addition of a chemical to the water.
Name one chemical added during this stage and state why this chemical is added. (9)
- (b) (i) Distinguish between the primary and secondary stages of sewage treatment. (12)
- (ii) What is the purpose of tertiary treatment? (6)

Answer

- (a) This important property of water is that it is a good solvent. (5)
- (i) **Sedimentation** or **settlement**: In this stage water is stored in tanks and the suspended (3) material in the water settles (3) at the bottom of the tanks.
- Flocculation**: The word “flocculation” means coagulation (3) or precipitation. Small suspended solids in the water can be made to form larger particles (called “flocs”) by adding certain chemicals to the water. Such chemicals are called flocculating agents (3).
- Tip: Note that in the above question, marks were also awarded for giving the name of a particular flocculating agent, e.g. aluminium sulfate (3).**
- Filtration**: The water is passed through large beds of sand (3). These filter beds remove any remaining suspended solids (3).
- (ii) **Chlorination** is another stage in the treatment of water. (3)
The chemical added in this stage is chlorine. (3)
Chlorine is added to the water in order to sterilise it, i.e. to get rid of any harmful micro-organisms in the water. (3)
- (b) (i) **Primary Treatment of Sewage**: Primary treatment is a mechanical process involving screening and settlement (3). Solids that are floating in the sewage are removed (3) by passing the sewage through steel bars. This process is called **screening**. Many materials in sewage are present as suspended particles that are carried along in flowing liquids. However, these suspended particles will settle out when the flow of liquid stops. Therefore, the sewage is allowed to flow into large settling tanks where it remains for a few hours. The solids settle to the bottom of these tanks and form a sludge. This process is called **settlement** or **sedimentation**.

Secondary Treatment of Sewage: Secondary treatment is a biological **(3)** process in which the levels of suspended and dissolved organic materials are reduced. The sewage is pumped into a large aeration tank where the sewage provides nutrients for a large number of growing micro-organisms. A rotor in the aeration tank continually churns the liquid and sludge with the air **(3)**

- (ii) **Tertiary Treatment:** Tertiary treatment involves the removal of phosphorus compounds [phosphates] **(3)** and nitrogen compounds [nitrates] **(3)** from effluents.

Note:

1. Either the term phosphate or phosphorus compounds are awarded full marks. Similarly, either the term nitrate or nitrogen compounds are awarded full marks.
2. Do not get confused between water treatment (for domestic use) and sewage treatment (before discharging into local waterway) when answering examination questions!

2003 Question 4

- (e) What happens during *secondary* sewage treatment?

(6)

Answer

- (e) Secondary treatment is a biological **(3)** process in which the levels of suspended and dissolved organic materials are reduced. The sewage is pumped into a large aeration tank where the sewage provides nutrients for a large number of growing micro-organisms. A rotor in the aeration tank continually churns the liquid and sludge with the air **(3)**.

2003 Question 8

- (c) The *free chlorine* present in swimming-pool water can be measured using a colorimeter or comparator.
- (i) What is the principle on which the technique that uses each of these methods is based? **(12)**
 - (ii) What is meant by *free chlorine*? **(6)**

Answer

- (c) (i) The principle on which each of these techniques is based is that absorbance [or intensity of colour] **(6)** is proportional **(3)** to concentration **(3)**.

Note: This question was badly answered by many students in the Leaving Certificate examination. If you are unsure of the above answer, study pages 323–327 in your textbook. The relationship is summarised in Fig. 19.25 on page 326 of the textbook.

- (ii) Free chlorine is chlorine that exists in water as hypochlorous acid **(3)** and hypochlorite ion **(3)**.

Tip: As well as the above names, full marks were also given for the correct formulas of hypochlorous acid (HOCl) and hypochlorite ion (OCl^-).

2004 Question 1

In an experiment to determine the total hardness of a water sample containing both calcium and magnesium ions, a solution of the reagent **edta** (ethylenediaminetetraacetic acid) in the form of its disodium salt (represented by $\text{Na}_2\text{H}_2\text{Y}$) was titrated against a sample of the water using a suitable indicator. The reaction between the ions (represented by M^{2+}) in the hard water and the **edta** reagent may be represented as



- (a) Name a suitable indicator or this titration.
What colour change is observed at the end point of the titration using this indicator? (8)
- (b) Describe the correct procedure for rinsing the burette and filling it with **edta** reagent. (15)
- (c) The addition of a small quantity of another solution to the water in the conical flask is essential before commencing the titration. What solution must be added and what is its purpose? (6)
- (d) In the experiment it was found that 100 cm^3 portions of the water required an average titre of 8.10 cm^3 of 0.010 M **edta** solution. Calculate the total hardness in
(i) moles per litre,
(ii) grams per litre expressed in terms of CaCO_3 and
(iii) p.p.m. expressed in terms of CaCO_3 . (15)
- (e) A whitish deposit is often found on the insides of kettles in hard water districts. If some of this deposit is scraped into a test tube and dilute hydrochloric acid is added a reaction is observed. Write a balanced equation for this reaction. (6)

Answer

- (a) The name of the indicator is Eriochrome Black T (4) or Solochrome Black.
At the end point, the indicator changes from red to blue.
- (b) The procedure may be summarised as follows:
- ⊗ Wash out the burette with deionised water.
 - ⊗ Wash out the burette with edta.
 - ⊗ Clamp the burette vertically on the retort stand.
 - ⊗ Using a funnel, fill the burette almost completely with the edta solution. Remove the funnel after filling.
 - ⊗ Open the tap and ensure that the space below the tap is filled. Check that there are no air bubbles in the nozzle.
 - ⊗ Adjust the bottom of the meniscus of the liquid to the zero graduation mark.
- Any four points: (6 + 3 × 3)
- (c) A buffer solution is added. (3)
The buffer solution is added because it is necessary to keep the pH around 10 for the indicator to work satisfactorily. (3)

(d) GIVEN:

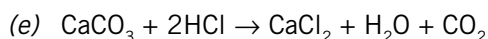
	Ca^{2+}	edta	
$V_{\text{Ca}} = 100$	$\frac{V_{\text{Ca}} \times M_{\text{Ca}}}{n_{\text{Ca}}}$	$=$	$\frac{V_{\text{ed}} \times M_{\text{ed}}}{n_{\text{ed}}}$
$M_{\text{Ca}} = ?$			
$n_{\text{Ca}} = 1$	$\frac{100 \times M_{\text{Ca}}}{1}$	$=$	$\frac{8.10 \times 0.01}{1}$
$V_{\text{ed}} = 8.10$			
$M_{\text{ed}} = 0.01$	M_{Ca}	$=$	$\frac{8.10 \times 0.01}{100}$
$n_{\text{ed}} = 1$			
			$= 0.00081 \text{ moles/L CaCO}_3$
			$= 0.00081 \times 100 \text{ g/L CaCO}_3$ (Rel molecular mass of $\text{CaCO}_3 = 100$)
			$= 0.081 \text{ g/L CaCO}_3$
			$= 0.081 \times 1000 \text{ mg/L CaCO}_3$
			$= 81 \text{ mg/L CaCO}_3$
			$= 81 \text{ p.p.m. CaCO}_3$

Answers:

(i) 0.00081 moles/L

(ii) 0.081 g/L

(iii) Total hardness of water = 81 p.p.m.



Correct formulas (3); Balancing (3)

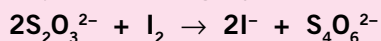
2005 Question 1

In an experiment to measure the concentration of dissolved oxygen in a river water sample, a bottle of water was filled from the river and was analysed immediately. The experiment was carried out as follows:

A few cm^3 each of concentrated manganese(II) sulfate (MnSO_4) solution and alkaline potassium iodide (KOH/KI) solution was added to the water in the bottle. The stopper was carefully replaced in the bottle and the bottle was shaken to ensure mixing of the reagents with the water. A brownish precipitate was produced. The stopper was removed from the bottle and a few cm^3 of concentrated sulfuric acid (H_2SO_4) were added carefully down the inside of the neck of the bottle using a dropper. The precipitate dissolved and a golden-brown solution was produced. The concentration of iodine (I_2) in the solution was found by titrating it in 50 cm^3 portions against a standard (0.01 M) sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$) solution.

- (a) Why was it necessary to analyse the sample of river water *immediately*? (5)
- (b) In making additions to the sample, why should the solutions be *concentrated*? (6)
- (c) Describe how the additions of the concentrated solution of manganese(II) sulfate (MnSO_4) and alkaline potassium iodide (KOH/KI) to the bottle of river water should be carried out. What essential precaution should be taken when replacing the stopper of the bottle after these additions are made? (9)
- (d) Describe clearly the procedure for using a pipette to measure exactly 50 cm^3 portions of the iodine (I_2) solution into the titration flask. (9)
- (e) What indicator is used in this titration? State when the indicator should be added to the titration flask and describe the colour change observed at the end point. (9)

- (f) The titration reaction is described by the following equation.



Calculate the concentration of the iodine solution in moles per litre given that 6.0 cm³ of the 0.01 M sodium thiosulfate (**Na₂S₂O₃**) solution were required in the titration for complete reaction with 50 cm³ portions of the iodine solution. **(6)**

- (g) For every one mole of oxygen gas (**O₂**) in the water sample 2 moles of iodine (**I₂**) are liberated in this experiment. Hence calculate the concentration of dissolved oxygen in the water sample in p.p.m. **(6)**

Answer

- (a) It was necessary to analyse the sample of water immediately so that the oxygen content did not change due to photosynthesis. **(5)**

Note: Full marks were also awarded to students who commented that the oxygen content could change as a result of respiration. (Recall that photosynthesis would increase the oxygen content and respiration would decrease the oxygen content).

- (b) The solutions used are concentrated to minimise the amount of the water sample that is displaced. [An alternative answer is that the solutions are concentrated to make sure that a small volume supplies an excess of reagents]. **(6)**

- (c) When adding the concentrated solutions use a dropper **(3)** and insert the end of the dropper under the surface of the water **(3)** in the bottle.

Stopper the bottle carefully to avoid dissolving any oxygen from the air. **(3)**

- (d) The main points in using the pipette are as follows:

- ⊗ Rinse out the pipette with water and also with iodine solution.
- ⊗ Fill the pipette using a pipette filler to above the graduation mark.
- ⊗ Adjust the bottom of the meniscus so that it is level with the graduation mark.
- ⊗ Remove any drop adhering to the outside of the pipette.
- ⊗ Allow the contents of the pipette to discharge into a conical flask.
- ⊗ Touch the tip of the pipette against the side of the conical flask (to add the droplet adhering to the outside of the tip).
- ⊗ Do not blow out the small amount of liquid in the tip of the burette. **Any three points: (3 × 3)**

- (e) The indicator is starch. **(3)**

The indicator is added when the solution has become pale yellow. **(3)**

The colour change observed is blue-black to colourless. **(3)**

Exam Edge

(f) GIVEN:

$$\begin{array}{lcl} V_o = 50 & I_2 & S_2O_3^{2-} \\ M_o = ? & \frac{V_o \times M_o}{n_o} = & \frac{V_r \times M_{red}}{n_r} \\ n_o = 1 & \frac{50 \times M_o}{1} = & \frac{6.0 \times 0.01}{2} \\ V_r = 6.0 & & \\ M_{red} = 0.01 & M_o = & \frac{6.0 \times 0.01}{2 \times 50} \\ n_r = 2 & & = 0.0006 \text{ moles per litre} \end{array} \quad \begin{array}{l} (3) \\ \\ \\ \\ (3) \end{array}$$

Answer: Concentration of iodine solution = 0.0006 moles per litre.

- (g) Since we are told that for every 1 mole of oxygen gas in the water sample, 2 moles of I_2 are liberated, then:

$$\begin{aligned} \text{Number of moles of oxygen} &= \frac{0.0006}{2} = 0.0003 \text{ moles/L} \\ &= 0.0003 \times 32 \text{ g/L} \\ &= 0.0096 \text{ g/L} \\ &= 0.0096 \times 1000 \text{ mg/L} \\ &= 9.6 \text{ mg/L } O_2 \\ &= 9.6 \text{ p.p.m. } O_2 \end{aligned} \quad \begin{array}{l} (3) \\ \\ \\ \\ (3) \end{array}$$

Answer: Concentration of dissolved oxygen = 9.6 p.p.m.

Note: In parts (f) and (g) of this question, the calculation of dissolved oxygen in the water was split into two separate calculations. Part (f) is a straightforward iodine/thiosulfate calculation as covered in Chapter 15 Volumetric Analysis: Oxidation Reduction. Part (b) involves a stoichiometric calculation to work out the number of moles of oxygen gas and to convert this amount to p.p.m.

2005 Question 8

- (d) What is meant by the *biological oxygen demand (BOD)* of a water sample? (6)
- (e) Describe clearly the processes involved in the primary and secondary stages of urban sewage treatment. What substances are removed by tertiary treatment of sewage? (18)

Answer

- (d) Biochemical oxygen demand is defined as the following:

- ⊗ { the amount of dissolved oxygen
- ⊗ { consumed by biological action (3)
- ⊗ { when a sample of water is kept at 20 °C
- ⊗ { in the dark (3)
- ⊗ { for five days

- (e) (i) **Primary Treatment of Sewage:** Primary treatment is a mechanical process involving screening and settlement (3). Solids that are floating in the sewage are removed (3) by passing the sewage through steel bars. This process is called **screening**. Many materials in sewage are present as suspended particles that are carried along in flowing liquids. However, these suspended particles will settle out when the flow of liquid stops. Therefore, the sewage is allowed to flow into large settling tanks where it remains for a few hours. The solids settle to the bottom of these tanks and form a sludge. This process is called **settlement** or **sedimentation**.

Secondary Treatment of Sewage: Secondary treatment is a biological (3) process in which the levels of suspended and dissolved organic materials are reduced. The sewage is pumped into

a large aeration tank where the sewage provides nutrients for a large number of growing micro-organisms. A rotor in the aeration tank continually churns the liquid and sludge with the air (3).

Tertiary Treatment of Sewage: Tertiary treatment involves the removal of phosphorus compounds [phosphates] (3) and nitrogen compounds [nitrates] (3).

2006 Question 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 cm³ of swimming pool water was filtered, the mass of the filter paper, upon drying, has increased by 0.78 g. When 250 cm³ 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)

Answer

- (a) The principle of all colorimetric experiments is based on the fact that absorbance [or intensity of colour] (4) is proportional to concentration (4).

Note: If you are unsure of the above answer, study pages 323–327 in your textbook. The relationship is summarised in Fig. 19.25 on page 326 of the textbook.

- (b) The reagent is DPD. (3)
DPD reacts with chlorine in water to give a pink colour. (3)
- (c) To estimate the concentration of free chlorine using a **comparator**:
 - ⊗ Add DPD to sample of water. (3)
 - ⊗ Note the colour that develops. (3)
 - ⊗ Compare the colour of the solution with the colour disc. (3)
 - ⊗ Match the two colours as best as possible and read off the concentration. (3)

OR

To estimate the concentration of free chlorine using a **colorimeter**:

- ⊗ Prepare a set of standard solutions. (3)
 - ⊗ Insert each solution in colorimeter and measure absorbance. (3)
 - ⊗ Plot graph of absorbance vs. concentration. (3)
 - ⊗ Use graph to obtain unknown concentration. (3)
- (d) Hypochlorous acid [HOCl] is the name of a free chlorine species in swimming pool water. (6)
Swimming pool water is more contaminated than drinking water due to the potentially harmful bacteria added by swimmers (3). Therefore, there is a need for greater concentration of chlorine in swimming pool water.

- (e) (i) Total suspended solids = $0.78 \text{ g}/1200 \text{ cm}^3$
 $= \frac{0.78}{1200} \text{ g/cm}^3$
 $= \frac{0.78 \times 1000}{1200} \text{ g/L}$
 $= 0.65 \text{ g/L}$
 $= 0.65 \times 1000 \text{ mg/L}$
 $= 650 \text{ p.p.m.}$ (3)
- (ii) Total dissolved solids = $0.32 \text{ g}/250 \text{ cm}^3$
 $= 0.32 \times 4 \text{ g/L}$
 $= 1.28 \text{ g/L}$ (3)
 $= 1.28 \times 1000 \text{ mg/L}$
 $= 1280 \text{ p.p.m.}$ (3)

Answer: Total suspended solids = 650 p.p.m.

Total dissolved solids = 1280 p.p.m.

2006 Question 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 $\text{Ca}(\text{HCO}_3)_2$. 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)

Answer

- (a) (i) Hard water is water that does not easily form a lather with soap. (5)
- (ii) Ion exchange units used to soften water contain cation exchange resin. The cation exchange resin exchanges the positive Ca^{2+} ions in the water (3) for Na^+ ions. Each calcium ion is replaced by two sodium ions (3).

OR



- (iii) **Flocculation:** The word “flocculation” means coagulation (3) or precipitation. It refers to the fact that very small suspended solids (3) in the water can be made to form larger particles (called “flocs”) by adding certain chemicals to the water.
 An example of a flocculating agent is aluminium sulfate, $\text{Al}_2(\text{SO}_4)_3$. (3)
- (iv) Calcium hydroxide is used to adjust the pH of water which is too acidic. (3)
 If the water supply is acidic it can cause corrosion of pipes. (3)

20 Electrochemistry I

Since the topic of Electrochemistry is not a major section of the course, the material in this short chapter is usually examined by means of brief questions. Whilst many students find electrochemistry a difficult topic, it is well worth spending some time studying the electrolysis reactions as the questions asked in this area are very straightforward.

2004 Question 4

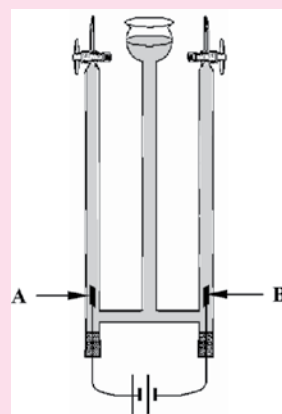
- (i) State and explain the colour observed at the negative electrode in the electrolysis of aqueous potassium iodide, containing a little phenolphthalein indicator, using inert electrodes. (6)

Answer

- (i) A pink colour is observed at the negative electrode. (3)
This is due to the fact that hydroxyl ions are produced at this electrode. (3)
[Phenolphthalein turns pink in an alkaline solution]

2005 Question 11

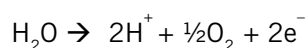
- (a) (iii) A solution of acidified water (dilute sulfuric acid) is electrolysed by passing an electric current through it using inert electrodes. At which electrode **A** or **B** does oxidation occur? Which species is oxidised? Write a balanced half equation for the oxidation reaction. (12)



Answer

- (iii) Oxidation occurs at electrode A. (3)
Water is the species oxidised. (3)

Note: Recall that oxidation always occurs at the positive electrode (OPRN) and that water is oxidised since it loses two electrons at the positive electrode.



Correct formulas: (3); Balancing: (3)

21 Fuels and Heats of Reaction

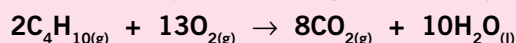
Questions on the *Fuels and Heats of Reaction* chapter feature in Section A of the paper as Mandatory Student Experiments No. 20, 21 and 22. These experiments have also been frequently examined in Section B of the paper. You will need to be able to draw labelled diagrams of the apparatus required for these experiments. For Experiments 20 and 21 you will need to be able to write balanced chemical equations for the reactions involved. It is important that you are aware of the physical appearance of many of the substances involved and of any observations made during these experiments. Knowledge of safety precautions is also regularly asked in relation to these experiments. For Mandatory Student Experiment No. 22 you should understand where inaccuracies may arise and how to minimise these. This experiment is usually associated with calculation of the heat of reaction, so you should practise this type of calculation. There are several definitions associated with this chapter so make sure you have learned these! You need to be able to name hydrocarbon compounds using the IUPAC format and be able to draw their structural formulas. The use of hydrocarbon compounds as fuels is often examined on the Leaving Certificate paper – usually in questions 6 or 8. This area includes details of how crude oil is refined, with emphasis regularly being placed on the octane number of a fuel and how this is related to the molecular structure of the hydrocarbon molecule. Details on methane as a fuel, and its environmental impact, feature regularly in questions on this chapter. Thermochemical problems usually appear in question 6 or 8 of the Leaving Certificate paper, or sometimes may be included as part of question 10 on the exam paper. These calculations are usually straightforward and you should ensure that you are able to carry out the calculations following the layout in the textbook. Details on benzene and other aromatic hydrocarbons are also included in this chapter and are examined from time to time as part of the questions discussed above.

2002 Question 8

The following hydrocarbons can all be used as fuels.

methane (CH₄) butane (C₄H₁₀) 2,2,4-trimethylpentane (C₈H₁₈)

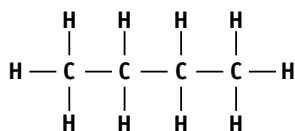
- (a) Butane is a major component of LPG. What do the letters LPG stand for? (5)
Draw two structural isomers of butane. (6)
- (b) Methane is a major component of natural gas.
Why are mercaptans often added to natural gas?
What environmental change or effect is associated with the release of methane to the atmosphere?
Apart from leaking gas pipes, name a major source from which methane is released to the atmosphere. (9)
- (c) What structural feature of 2,2,4-trimethylpentane results in it having a high octane rating?
Give one other structural feature which increases the octane number of a hydrocarbon. (6)
- (d) Define *heat of combustion* of a compound. (6)
- (e) The combustion of butane is described by the following balanced equation.



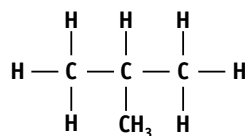
Calculate the heat of combustion of butane given that the heats of formation of butane, carbon dioxide and water are -125 , -394 and -286 kJ mol^{-1} , respectively. (18)

Answer

- (a) Liquefied petroleum gas. (5)



(a) Butane



(b) Methylpropane

(2 × 3)

- (b) Mercaptans are added to natural gas to give it an unpleasant smell in case of gas leaks. (3)

Methane is a gas which is contributing to the “greenhouse effect”. (3)

Some farmers allow manure to decay under anaerobic conditions (i.e. in the absence of air). Bacteria feed on the manure causing it to decay and methane is formed. (3)

- (c) 2,2,4-trimethylpentane has several methyl branches. The more branched the chain, the higher the octane number. (3)

Another structural feature is the short chain; the shorter the alkane chain, the higher the octane number. (3)

- (d) The heat of combustion of a substance is the heat change when one mole (3) of the substance is completely burned in excess oxygen (3).

- (e) REQUIRED: $\text{C}_4\text{H}_{10} + 6\frac{1}{2} \text{O}_2 \rightarrow 4 \text{CO}_2 + 5 \text{H}_2\text{O}$ $\Delta H = ?$

GIVEN:



REARRANGE:

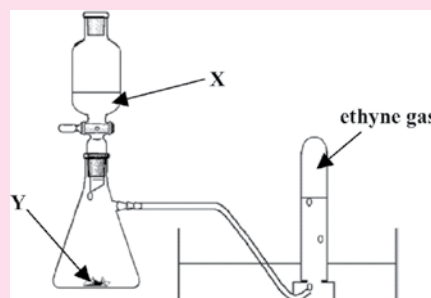


Note: If you use the equation given in the question when doing your calculation, you will get a heat change involving two moles of C_4H_{10} in the answer. However, the definition of heat of combustion refers to only one mole of a substance being burned. The answer you get in that case will therefore have to be divided by two.

2003 Question 2

The diagram shows an apparatus that can be used for the preparation of ethyne gas, C_2H_2 .

A liquid **X** is dropped onto the solid **Y** and the gas collected in test tubes as shown.



- (a) Identify the liquid **X** and the solid **Y**. (8)
- (b) Describe the appearance of the solid **Y**. (3)
- (c) Write a balanced equation for the reaction between **X** and **Y** producing ethyne. (6)
- (d) What is observed when a sample of ethyne gas is burned in air? Write a balanced equation for the combustion of ethyne in oxygen. (9)
- (e) Ethyne, C_2H_2 , is described as an *unsaturated* hydrocarbon. Describe a test you could carry out to show that ethyne is *unsaturated*. Write an equation for the reaction taking place. Name the organic product. (18)
- (f) The common name for ethyne gas, C_2H_2 , is acetylene. Give **one** major use of the gas. (6)

Answer

- (a) X = water [H_2O] (4)
Y = calcium carbide [CaC_2] (4)
- (b) Calcium carbide is a grey [grey-black] solid. (3)
- (c) $\text{CaC}_2 + 2\text{H}_2\text{O} \rightarrow \text{Ca}(\text{OH})_2 + \text{C}_2\text{H}_2$ (6)
- (d) Ethyne burns with a luminous, smoky flame, and a great deal of soot is formed. (3)
 $2\text{C}_2\text{H}_2 + 5\text{O}_2 \rightarrow 4\text{CO}_2 + 2\text{H}_2\text{O}$ (6)

Formulas: (3); Balanced: (3)

- (e) Bromine can be used to test for unsaturation. Add bromine water (6) to a test-tube of ethyne. Replace the stopper and shake (6) the test-tube a few times. The red colour of the bromine disappears (6).

Note: The equation for the reaction and the organic product are not required by the syllabus. Hence, all marks were allocated for the above part.

- (f) Ethyne is used in the welding and cutting of metals. (6)

2003 Question 4

- (j) State **two** ways, other than the addition of lead compounds, that the octane rating of a fuel can be increased.

Answer

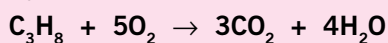
The octane rating of a fuel can be increased by (i) isomerisation (3) and (ii) catalytic cracking (3).
[Two other possible answers: (iii) reforming or dehydrocyclisation, and (iv) adding oxygenates.]

2003 Question 10

- (a) Define *heat of combustion*. (7)

Propane may be used in gas cylinders for cooking appliances.

Propane burns according to the equation

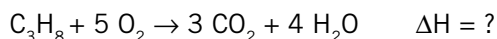


- (i) The heats of formation of propane, carbon dioxide and water are -104 , -394 and -286 kJ mol^{-1} respectively. Calculate the heat of combustion of propane. (12)
- (ii) If 500 kJ of energy are needed to boil a kettle of water, what mass of propane gas must be burned to generate this amount of heat? Express your answer to the nearest gram. (6)

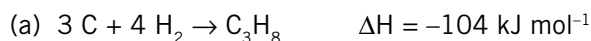
Answer

- (a) The heat of combustion of a substance is the heat change when one mole (4) of the substance is completely burned in excess oxygen (3).

- (i) REQUIRED:



GIVEN:



REARRANGE:



- (ii) 2222 kJ of energy are released when one mole of propane (44 grams) is completely burned

$$\Rightarrow 1 \text{ kJ of energy is released when } \frac{44}{2222} \text{ grams are burned} = 0.0198 \text{ kJ}$$

$$\Rightarrow 500 \text{ kJ would be released when } 500 \times 0.0198 \text{ grams are burned} \quad (3)$$

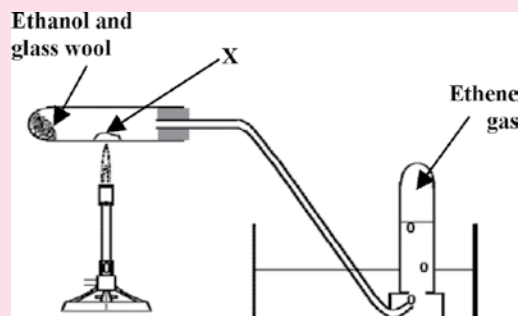
$$= 9.9 \text{ grams}$$

$$\approx 10 \text{ grams} \quad (3)$$

TIP: Note that you will lose 1 mark if you don't round off your answer as requested.

2004 Question 2

The diagram shows the experimental set-up used by a group of students to prepare a sample of ethene from ethanol and to collect the ethene produced.



- (a) What is the function of the glass wool? (5)
Identify the solid **X** and describe its appearance. (6)
- (b) State and explain **two** safety precautions which should be observed when carrying out the student experiment. (12)
- (c) Write a balanced equation for the reaction involved in this preparation. (6)
- (d) If the ethene produced is bubbled through an acidified solution of potassium manganate(VII), the solution is decolourised showing that ethene is *unsaturated*. What is meant by the term *unsaturated*? Describe how you would carry out another test to confirm that ethene is unsaturated. (12)
- (e) Describe the flame that would be observed when a combustion test is carried out on a sample of ethene gas. Write a balanced equation for the combustion of ethene in excess oxygen. (9)

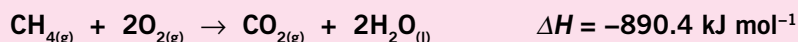
Answer

- (a) The function of the glass wool is to soak up the ethanol and to hold the ethanol in place while the reaction is occurring. (5)
Solid X is aluminium oxide. (3)
It is a white solid. (3)
- (b) Remove the delivery tube from the water before turning off the Bunsen burner (3). This prevents a suck-back of water (3).
Make sure that the stopper is air tight (3) as the gas is flammable and if it escapes it could cause a fire or an explosion (3).
[Use tongs rather than handling the glass wool (3) as glass wool is an irritant to the skin (3).]
- (c)
- $$\begin{array}{c} \text{H} & \text{H} \\ | & | \\ \text{H}-\text{C} & -\text{C}-\text{H} \\ | & | \\ \text{H} & \text{OH} \end{array} \xrightarrow[-\text{H}_2\text{O}]{\text{Al}_2\text{O}_3} \begin{array}{c} \text{H} & & \text{H} \\ & \diagdown & / \\ & \text{C}=\text{C} \\ & / & \diagdown \\ \text{H} & & \text{H} \end{array} + \text{H}_2\text{O}$$
- (6)
- (d) The term unsaturated means that a molecule contains a carbon-carbon double bond (3) [or a carbon-carbon triple bond]
Another test to confirm that ethene is unsaturated is to add (3) [bubble] ethene to bromine (3) water, the red/yellow colour (3) disappears and a new colourless (3) compound is formed.
- (e) The gas burns with a luminous [yellow] flame. (3)
The balanced equation for the combustion of ethene is as follows:
 $\text{C}_2\text{H}_4 + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 2\text{H}_2\text{O}$ Formulas: (3); Balanced: (3)

2004 Question 6

(a) Define (i) *heat of formation of a substance*, (ii) *octane number of a fuel*. (11)

(b) The combustion of methane is described by the following balanced equation.



The standard heats of formation of carbon dioxide and water are -394 and -286 kJ mol^{-1} respectively. Calculate the heat of formation of methane. (12)

(c) Methane is an excellent fuel. Give **two** properties of methane which account for its usefulness as a fuel. Natural gas is a rich source of methane. Why are mercaptans often added to natural gas? (9)

(d) Methane is often found in gas fields which occur in association with crude oil deposits. Crude oil is fractionated in order to obtain more useful products. Outline clearly how the fractionation process is carried out. (12)

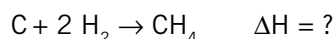
(e) Identify **two** structural features of a hydrocarbon fuel which affect its octane number. (6)

Answer

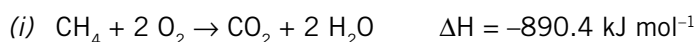
(a) (i) The heat of formation of a compound is the heat change that takes place when one mole (3) of a compound in its standard state is formed from its elements in their standard states (2).

(ii) The octane number of a fuel is a measure of the tendency of the fuel to resist knocking. (6)

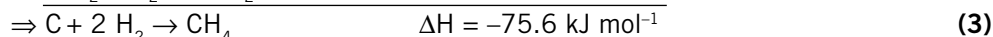
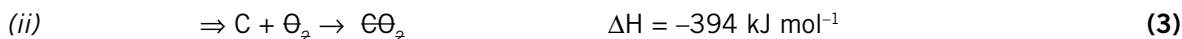
(b) REQUIRED:



GIVEN:



REARRANGE:



(c) Two properties of methane that contribute to its usefulness as a fuel are its high kilogram calorific value (3) and that it is a relatively clean fuel (3). [Also, it can be piped for distribution to houses and it is non-toxic.]

Mercaptans are added to natural gas to give an unpleasant smell in case of gas leaks. (3)

(d) The main points of the fractionation process are as follows:

- ⊗ Heating the crude oil and separating the various substances on the basis of their boiling points.
- ⊗ The crude oil passes up through fractionating column.
- ⊗ The different fractions condense at different levels in the fractionating column.
- ⊗ The lighter gases come out the top [or the larger hydrocarbons that have higher boiling points turn to liquids nearer the bottom of the column].

(The first of these points gets 6 marks, and any other two points get 3 marks each.)

TIP: Note that the first three of these points may be obtained from a diagram similar to that shown on the right. The last point may also be obtained from a diagram provided that the diagram is labelled to show what fractions come off at the top or bottom of the fractionating column.

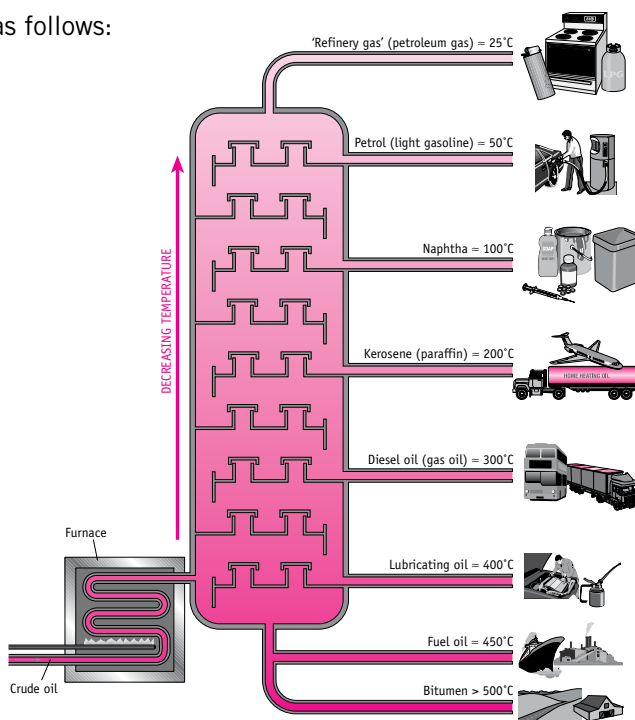


Fig. 21.1

(e) The shorter the alkane chain the higher the octane number **(3)** and the more branched the chain the higher the octane number **(3)**. [Also cyclic compounds have a higher octane number than straight chain compounds.]

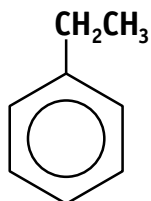
2005 Question 4

- (g) Name and draw the structure of an aromatic compound of molecular formula C_8H_{10} .
- (i) Draw the structural formula of an organic compound of molecular formula C_3H_6 . Label clearly any tetrahedrally bonded carbon atom in the molecule.

Answer

(g) Ethylbenzene

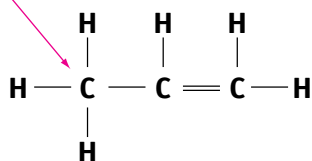
(3)



(3)

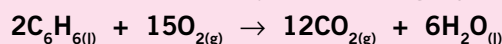
(i) Tetrahedral carbon

(3)



2005 Question 6

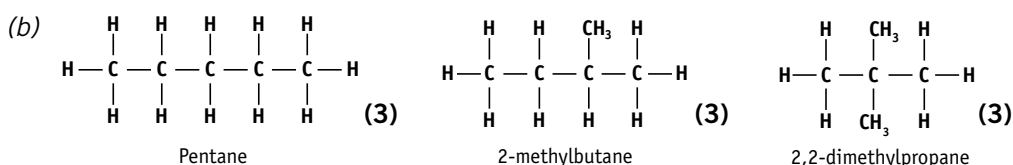
- (a) The octane number of a fuel is described as *a measure of the tendency of the fuel to cause knocking*, or as *a measure of the tendency of the fuel to resist auto-ignition*. This number is found by comparing the combustion of the fuel with the combustion of a mixture of two reference hydrocarbons using the same standard engine.
- Name **both** of the reference hydrocarbons present in the mixture used when measuring the octane number by this comparison method. (8)
 - State **two** structural features of a hydrocarbon molecule which contribute to it having a high octane number. (6)
 - Lead compounds were used in the past to increase the octane number of fuels. Why are lead compounds unsuitable as additives for petrol used in modern cars? (3)
 - Identify **one** additive *or* type of additive, other than a compound of lead, used to increase the octane number of fuels. (3)
- (b) There are **three** structural isomers of the hydrocarbon of formula C_5H_{12} . In the case of **each** of these isomers, draw the structure of the molecule and give its systematic IUPAC name. (18)
- (c) The combustion of liquid benzene is described by the following equation:



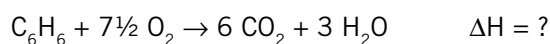
Given that the heats of formation of carbon dioxide gas, liquid water and liquid benzene are -394 , -286 and 49 kJ mol^{-1} respectively, calculate the heat of combustion of liquid benzene. (12)

Answer

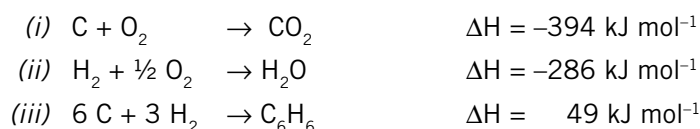
- (a) (i) The compound 2,2,4-trimethylpentane (**4**) is assigned an octane number of 100. Heptane (**4**) is assigned an octane number of 0.
- (ii) The shorter the alkane chain the higher the octane number (**3**) and the more branched the chain the higher the octane number (**3**). [Also, cyclic compounds have a higher octane number than straight chain compounds.]
- (iii) Lead compounds in the exhaust fumes poison the metal catalysts in the catalytic converter of the car (**3**) [destroy the catalytic converter.]
- (iv) Oxygenates are added to fuels to increase octane number. (3)



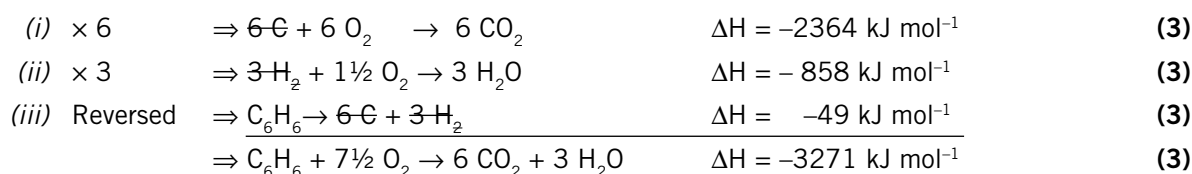
(c) REQUIRED:



GIVEN:



REARRANGE:



Note: If you use the equation given in the question when doing your calculation, you will get a heat change involving two moles of C_6H_6 in the answer. However, the definition of heat of combustion refers to only one mole of a substance being burned. The answer you get in that case will therefore have to be divided by two.

2005 Question 7

- (c) Describe with the aid of a labelled diagram how the conversion **W** may be carried out in a school laboratory and how a sample of the product may be collected. How would you test this product to show that it is unsaturated? **(18)**

Answer

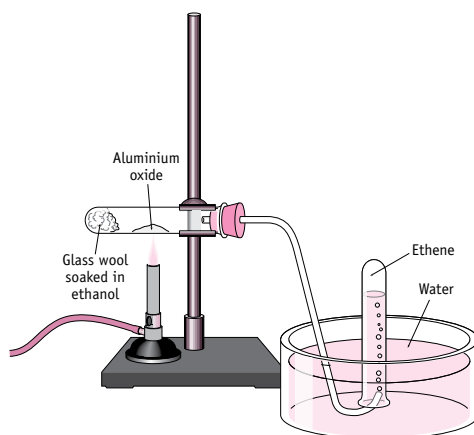


Fig. 21.2

The key points in the diagram are as follows:

- ⊗ Horizontal test-tube with delivery tube connected and collection of gas over water. **(3)**
- ⊗ Bunsen burner for heating. **(3)**
- ⊗ Aluminium oxide [Al_2O_3] **(3)**
- ⊗ Ethanol held at end of tube. **(3)**

TIP: Full marks can be obtained if all of the above points are included in the labelled diagram.

To test ethene to show that it is unsaturated, add bromine **(3)** water to ethene, the red/yellow colour disappears and a new colourless **(3)** compound is formed.

2006 Question 6

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

(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)

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

(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 kJ mol^{-1} , respectively, calculate the heat of combustion of cyclohexane. (12)

Answer

(a) (i) The octane number of a fuel is a measure of the fuel's tendency to resist knocking. (8)

(ii) The structural feature is that it is an unbranched molecule [i.e. has a straight chain]. (3)

(iii) Cyclohexane has a high octane number because it has a cyclic structure. (3)

Benzenes has a high octane number because it is aromatic. (3)

2,2,4-trimethyl pentane has a high octane number because it is highly branched. (3)

(iv) Hexane is converted to these compounds by a process called dehydrocyclisation. (3)

The use of benzene is controlled because it is a carcinogenic compound. (3)

(b) (i) Oxygenates are added to petrol in order to increase the octane number of the petrol (3). Also, oxygenates give rise to very little pollution (3) when they burn.

(ii) The lead compounds in the exhaust fumes are toxic. (3)

Also, lead compounds poison the metal catalysts in the catalytic converter of the car. (3)

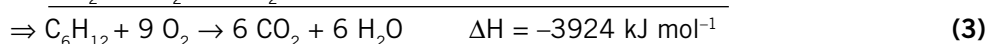
(c) REQUIRED:



GIVEN:



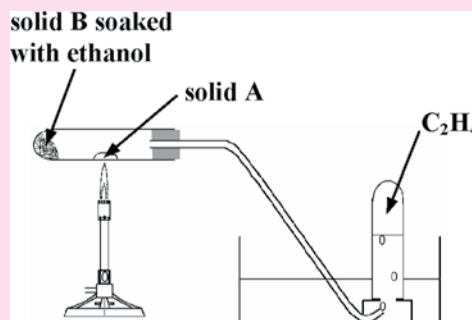
REARRANGE:



2006 Question 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. (6)
- (ii) Identify the solid **B** which is used to keep the ethanol at the end of the test tube. (3)
- (iii) What precaution should be observed when heating is stopped? Why is this necessary? (6)
- (iv) Give **one** major use of ethene gas. (3)
- (d) Draw the structures and give the systematic (IUPAC) names for **two** alkenes of molecular formula C_4H_8 . (12)



Answer

- (a) A homologous series is a series of compounds defined as the following:

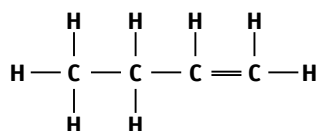
- ⊗ A series of compounds of uniform chemical type (same functional group).
- ⊗ Showing gradations in physical properties.
- ⊗ Having a general formula for its members.
- ⊗ Each member has a similar method of preparation.
- ⊗ Each member differs from the previous member by a (CH_2) unit.

Any one of these points: (5 marks)

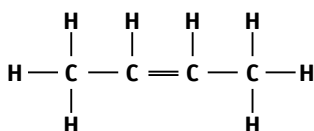
- (b) (i) Solid A is aluminium oxide (3), Al_2O_3 (3).
- (ii) Solid B is glass wool. (3)
- (iii) The tubing should be removed from the water before the heating is stopped (3) in order to prevent suck-back (3).
- (iv) Ethene is used for ripening fruit such as bananas. (3)
[Also, it is the starting material from which many other substances (particularly plastics) are made, e.g. polythene (poly-ethene), Terylene, synthetic rubber, polystyrene, PVC, etc.]

TIP: You need to specify an exact use for ethene. A general use, such as "in industry", will not suffice.

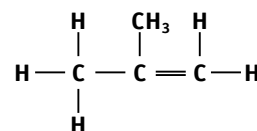
- (d) The alkene isomers of C_4H_8 are as follows:



But-1-ene (1-butene)



But-2-ene (2-butene)



Methylpropene

Any two correct formulas with matching names: (4 × 3)

TIP: "Butene" will get no marks. Also, you are only asked for two compounds. If you give a third and it is not a correct structure or name, you will lose marks because a wrong answer will cancel out a correct answer.

22 & 23 Organic Chemistry

These two *Organic Chemistry* chapters play a very important role on the Leaving Certificate Chemistry examination paper. Firstly, these chapters cover six Mandatory Student Experiments. These experiments are numbered 23–28 in the list given in the textbook. Therefore, it is possible that these two chapters can be examined in both Section A and Section B of the exam paper. Secondly, these two chapters cover a wide range of organic chemistry which is examined in one or more full questions in Section B of the exam paper. On some occasions, three full questions based on these two chapters have been given on the examination paper. As you are aware, organic chemistry is also covered in the section on Fuels (Chapter 21) but this is usually examined in a separate question. In short, Chapters 22 and 23 play an important role on the Leaving Certificate examination paper so make sure you study these chapters well!

2002 Question 2

Soap is produced by the hydrolysis of vegetable and animal fats.

- (a) What is the principal chemical difference between vegetable and animal fats? (5)

A sample of soap was prepared in a school laboratory as follows:

Approximately 3 g of lard (animal fat), 2 g of sodium hydroxide pellets (an excess), and 25 cm³ of ethanol were placed in a round-bottomed flask. A condenser was fitted to the flask and the mixture was refluxed gently for 20 minutes (Diagram 1).

Following the reflux, the apparatus was allowed to cool slightly and the arrangement of the apparatus was changed so that the ethanol could be removed by distillation (Diagram 2).

The residue from the distillation flask was then dissolved in a minimum of hot water and the solution decanted into 75 cm³ of brine. The soap was then isolated.

- (b) Apart from the lard, sodium hydroxide and ethanol, what else should be added to the reaction flask prior to the reflux? (15)
Why was the mixture refluxed?
Why was the ethanol added?
- (c) Why was it desirable to remove the ethanol after the reflux? (9)
- (d) Why was a minimum of hot water used to dissolve the residue from the distillation? What is brine? (9)
- (e) Describe how the soap could be isolated from the mixture of soap and brine. Give one precaution that helps to ensure that the soap is free of sodium hydroxide. (12)

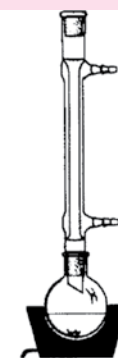


Diagram 1

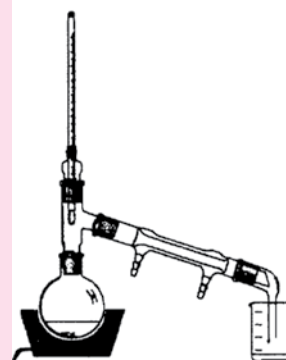


Diagram 2

Answer

- (a) Vegetable oils are unsaturated [polyunsaturated] but animal fats are saturated. (5)
- (b) Anti-bumping chips should be added. (6)
The mixture is refluxed so that the liquid is kept at its boiling point (3) without losing the vapour (3).
The ethanol acts as a solvent for the lard. (3)

Exam Edge

- (c) Distilling off most of the ethanol makes it easier to isolate the soap. (6)
Thus the yield of soap is maximised. (3)
- (d) A minimum of hot water is used to minimise the amount of soap that remains dissolved [and to avoid diluting the brine]. (6)

Note: This is an unusual question as it is not normal procedure for hot water to be added to the residue from the distillation. The normal procedure is to distil off some of the ethanol and then add the concentrated solution of soap in ethanol directly from the flask into the brine. Adding hot water is not recommended as this would result in a lather being formed inside the flask with a consequent loss of soap and lower yield.

Brine is a solution of salt in water. (3)

- (e) To isolate the soap, filter (6) it off using a filter funnel and filter paper.
Wash the soap with fresh salt solution to remove any sodium hydroxide still present. (6)

2002 Question 4

- (j) Draw the structure and give the IUPAC name for CH_3CHO .

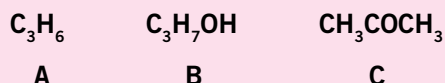
Answer



The IUPAC name for CH_3CHO is ethanal. (3)

2002 Question 6

Answer the questions (a) to (e) with reference to the compounds **A**, **B** and **C**.



- (a) Which of these three compounds would you expect to be the least soluble in water?
Give a reason for your answer. (8)
- (b) Give the IUPAC names for compounds **A** and **C**. Name both isomers of compound **B**.
Name a compound that is a structural isomer of **C**. (15)
- (c) Classify each of the compounds **A**, **B** and **C** as having one, two or three tetrahedrally bonded carbon atoms. (9)
- (d) Compound **C** can be synthesised from compound **A** in two steps, with one of the isomers of **B** as the product of the first step. Name suitable reagents for each of the steps. (12)
- (e) Compound **C** is a solvent. Give a common use for this solvent. (6)

Answer

- (a) C_3H_6 will be the least soluble in water. (4)
It is the least soluble because it is non-polar [and hence there are no intermolecular forces with water]. (4)
- (b) **A** = propene (3)
C = propanone (3)
The isomers of **B** are propan-1-ol (3), and propan-2-ol (3).
Isomer of **C** = propanal (3)

- (c) A has **one** tetrahedrally bonded carbon atom. (3)
 B has **three** tetrahedrally bonded carbon atoms. (3)
 C has **two** tetrahedrally bonded carbon atoms. (3)
 (d) The reagent to convert A to B is water. (6)
 The reagent to convert B to C is acidified (3) sodium dichromate (3).

TIP: Full marks also awarded for $\text{Na}_2\text{Cr}_2\text{O}_7 / \text{H}^+$.

- (e) Compound C is commonly used in nail varnish remover. (6)

2002 Question 7

Mass spectrometry and chromatography are widely used instrumental techniques in chemistry.

- (a) Give one application of each of these techniques. (8)
 (b) What are the main principles on which each of these techniques is based? (18)
 (c) What are the fundamental processes that occur in a mass spectrometer? (15)
 (d) HPLC is another chromatographic technique.
 What do the letters HPLC stand for? State one application of this technique. (9)

Answer

Note: Parts of this question relating to mass spectrometry have been covered in Chapter 4 of this book.

- (a) Gas chromatography is used to measure the level of alcohol in blood or urine samples [and in drug tests]. (4)
 (b) The main points in describing the principle of gas chromatography are as follows:
 ☉ It uses a gas as the mobile phase.
 ☉ The stationary phase is a high boiling point liquid spread on solid particles in a long coiled tube called the column.
 ☉ The various components in the mixture are separated out due to different attractions within the phases. (3 × 3)

TIP: The requirement for the wording in the last point above was very specific in the marking scheme, so study this point carefully!

- (c) [Answer given in Chapter 4, page 8]
 (d) HPLC stands for High Performance (3) Liquid Chromatography (3).
 HPLC is commonly used to test for the presence of growth promoters in meat [and caffeine in drinks]. (3)

2003 Question 4

- (h) Draw the structure and give the name of an *ester* of the molecular formula $\text{C}_3\text{H}_6\text{O}_2$.

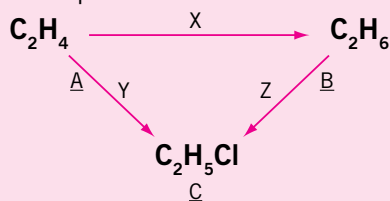
Answer

- (h)  (3)
 Ethyl methanoate

The name of the ester is ethyl methanoate. (3)

2003 Question 6

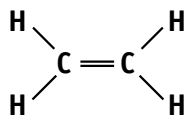
Study the reaction scheme and answer the questions which follow.



- (a) Which of the compounds **A**, **B**, and **C** has no tetrahedrally bonded carbon atoms?
Draw the structure of a molecule of this compound. (8)
- (b) Classify the conversions **X**, **Y** and **Z** as *addition*, *substitution* or *elimination* reactions. (9)
- (c) Which reagent is used to convert **A** to **C**? (3)
- (d) Which reagent and what conditions are required for the conversion of **B** to **C**? (6)
- (e) Describe the mechanism of the reaction for the conversion of **A** to **C**. (18)
State **one** piece of experimental evidence which supports the mechanism you have proposed. (6)

Answer

- (a) Compound A has no tetrahedrally bonded carbon atom. (4)



(4)

- (b) X = addition reaction (3)
Y = addition reaction (3)
Z = substitution reaction (3)
- (c) Hydrogen chloride [HCl] is used to convert A to C. (3)

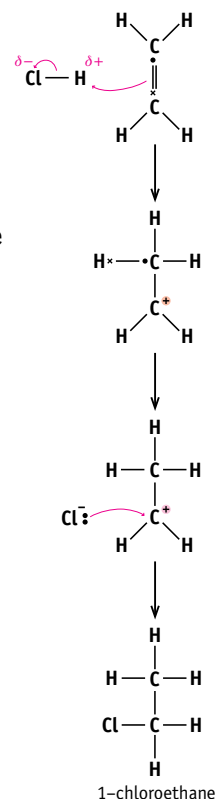
Note: It is important that you use the name “hydrogen chloride” or the formula HCl. No marks were awarded for the term “hydrochloric acid” as this would give a different product. Study page 419 in the textbook for further examples of addition reactions.

- (d) Chlorine [Cl₂] is the reagent used to convert B to C. (3)
The conditions are that ultraviolet light must be present. (3)
- (e) The main points in the addition reaction of HCl to ethene are as follows:
- ⊗ Polarised H – Cl bond. (3)
 - ⊗ HCl splits into ions [heterolytic fission]. (3)
 - ⊗ H⁺ bonds with one carbon atom using electrons of double bond. (3)
 - ⊗ Formation of carbonium ion. (3)
 - ⊗ Cl[−] attacks. (3)
 - ⊗ Bond formed between Cl[−] and carbonium ion. (3)

All of the marks for part (e) may be obtained from a suitable diagram as shown on page 369 of the textbook.

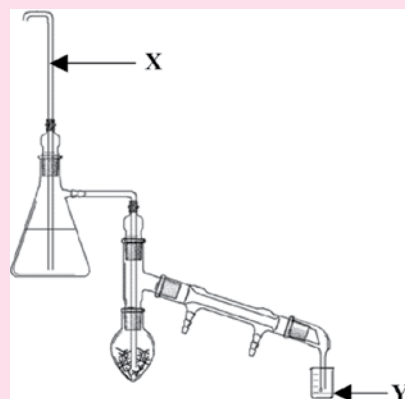
Note: When describing mechanisms of reactions, it is always a good idea to use a diagram to show the various steps and adding a few short comments of explanation as shown in the six points on the previous page.

Evidence for the proposed mechanism comes from the fact that when ethene reacts with bromine in water in the presence of sodium chloride (**3**), 1-bromo-2-chloroethane (**3**) [or 2-bromoethanol] is formed.



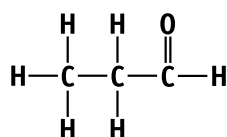
2003 Question 9

- (a) Draw the structure and state the IUPAC name for the aldehyde of the molecular formula $\text{C}_3\text{H}_6\text{O}$. (8)
 Draw the structure and give the name of another carbonyl compound that has the same molecular formula, $\text{C}_3\text{H}_6\text{O}$. Give **one** use of this compound. (12)
 Which of these two carbonyl compounds is easily oxidised to a carboxylic acid? Name the acid. (6)
- (b) The diagram shows the arrangement of glassware for the extraction of clove oil from cloves by steam distillation.
- What is the purpose of the tube marked **X**? (6)
 - What is collected at **Y**? Describe its appearance. (12)
 - State **one** use of clove oil. (6)



Answer

- (a) The structure of the aldehyde $\text{C}_3\text{H}_6\text{O}$ is as follows:

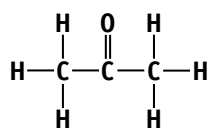


The IUPAC name for the aldehyde is propanal.

(4)

(4)

The structural formula of another carbonyl compound C_3H_6O is the following:



(3)

The name of the compound is propanone. (3)

Propanone is used in nail varnish remover. (3)

Propanal is the carbonyl compound that is easily oxidised to a carboxylic acid. (3)

The name of the acid is propanoic acid. (3)

(b) (i) X is a safety (3) tube and its purpose is to release pressure (3) [or release steam].

(ii) A mixture of clove oil (3) [eugenol] and water (3) are collected at Y.

The emulsion has a milky colour. (6)

(iii) Clove oil is used for flavouring food. (6)

2004 Question 3

(a) A sample of impure benzoic acid was recrystallised as follows: 2.5 g of the impure benzoic acid were weighed out and dissolved in the minimum amount of hot water. The hot solution was filtered and the filtrate was allowed to cool and recrystallise. The recrystallised benzoic acid was isolated by filtration. After drying, 2.25 g of purified acid were obtained.

(i) Why is it important to use the minimum amount of hot water in this procedure? (5)

(ii) Indicate clearly the stage of the recrystallisation procedure at which *insoluble* impurities were removed and how their removal was achieved. Indicate also the stage at which *soluble* impurities were removed and how their removal was achieved. (12)

(iii) How could you have ensured that the recrystallisation was complete? (3)

(iv) How could the crystals have been dried? (3)

(v) What was the percentage yield of purified benzoic acid? (3)

(b) Melting points of the impure and recrystallised benzoic acid were taken and compared.

(i) Describe with the aid of a labelled diagram how you would have measured the melting point of one of these samples. (15)

(ii) Give **two** ways in which you would expect the melting point of the impure benzoic acid to differ from that of the purified acid. (6)

(iii) State **one** use of benzoic acid and its salts. (3)

Answer

(a) (i) It is important to use the minimum amount of hot water in order to get a solution which is almost saturated (5) at the boiling point of the solvent. This maximises the final yield.

(ii) The insoluble impurities are removed when the hot solution of benzoic acid is filtered (3), i.e. in the first filtration.

The removal of the insoluble impurities is achieved because they cannot pass through the filter paper; they are left on the filter paper. (3)

The soluble impurities are removed when the recrystallised benzoic acid is filtered, i.e. in the second filtration. (3)

The removal of the soluble impurities is achieved because they are dissolved in the filtrate and so they pass through the filter paper. (3)

(iii) To ensure that the crystallisation is complete, allow the solution to cool fully [or cool the solution in iced water or in crushed ice]. (3)

(iv) The crystals are dried on filter paper [or placed in a desiccator]. (3)

(v) Percentage yield = $\frac{2.25}{2.5} \times 100 = 90\%$ (3)

- (b) (i) The main points are as follows:
- ⊗ Set up the aluminium block with the benzoic acid in the capillary tube as shown.
 - ⊗ Use the hot plate to gently heat the aluminium block.
 - ⊗ Carefully observe the crystals in the capillary tube as the temperature increases.
 - ⊗ Record the temperature range over which the crystals melt.
- [Any one of these points (including the labelled diagram) gets 6 marks; any three other points get 3 marks each]
- (ii) The impure benzoic acid melts at a lower temperature than the pure benzoic acid (3). The melting point of the impure benzoic acid is not sharp, i.e. it melts over a wider range of temperatures (3).
- (iii) Benzoic acid is used as a food preservative.

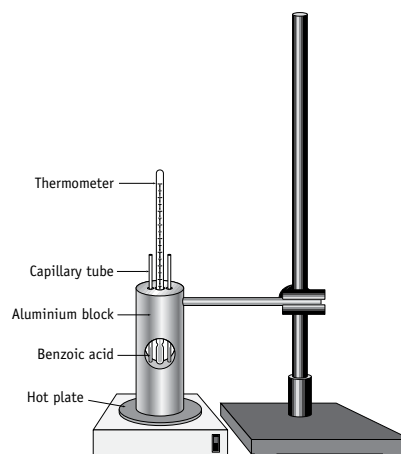
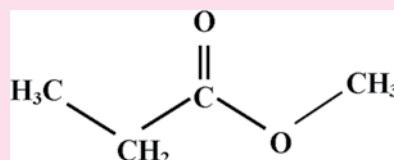


Fig. 5.2

(3)

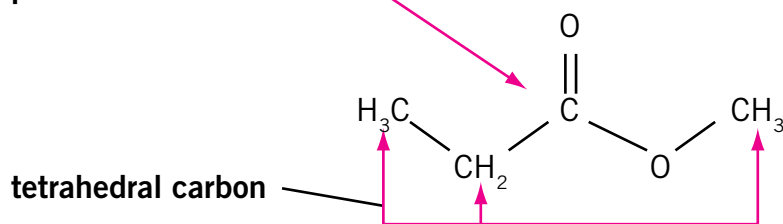
2004 Question 7

- (a) Copy into your answer book the structure of the ester shown and indicate clearly on your diagram a carbon atom which is in planar geometry in the molecule, and also a carbon atom which is in tetrahedral geometry in the molecule. (8)
- (b) Give the names of the alcohol and of the carboxylic acid from which the ester shown in the diagram is synthesised. What organic reaction type describes this esterification reaction? (15)
- (c) The carboxylic acid you were asked to name in (b) may itself be synthesised in two steps from an alcohol.
- (i) Identify the alcohol from which the carboxylic acid is derived.
 - (ii) Give the name and structure of the intermediate organic compound in this synthesis.
 - (iii) Identify the type of organic reaction involved in each step.
 - (iv) Identify the inorganic reagents which may be used in this synthesis. (21)
- (d) State **two** common uses of esters. (6)



Answer

- (a) planar carbon



Indicate the planar carbon.

(4)

Indicate a tetrahedral carbon.

(4)

- (b) Name of alcohol: Methanol.

(6)

Name of carboxylic acid: Propanoic acid.

(6)

Type of reaction: A condensation [or substitution] reaction.

(3)

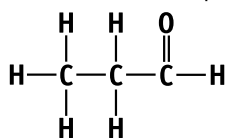
- (c) (i) The alcohol is propan-1-ol.

(6)

(ii) The name of the intermediate compound is propanal.

(3)

The structure of propanal is the following:



(3)

(iii) Each step involves an oxidation reaction.

(3)

(iv) Acidified (3) sodium dichromate (3) is used in the synthesis.

(d) Esters are used in flavourings (3) and perfumes (3).

2005 Question 2

A group of students prepared ethanal (CH_3CHO) by slowly adding an aqueous solution of ethanol ($\text{C}_2\text{H}_5\text{OH}$) and sodium dichromate(VI) ($\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$) to a hot aqueous solution of sulfuric acid (H_2SO_4). The apparatus drawn below was used. The reaction is described by the following equation.



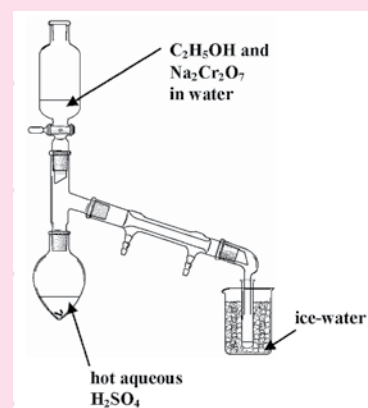
(a) Why was the receiving vessel cooled in ice-water? (5)

(b) State **two** features of the preparation that are necessary to maximise the yield of ethanal and, for each feature stated, explain why it is necessary. (12)

(c) Describe and account for the colour change which is observed during the addition of the ethanol and sodium dichromate(VI) solution to the hot acid. (9)

(d) Describe how you would carry out Fehling's test on a sample of ethanal. What observations would you expect to make in this test? (12)

(e) Assuming that all of the features needed to maximise the yield of ethanal would be present, what mass of ethanal would be produced in the preparation if the students used 8.94 g of sodium dichromate(VI) ($\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$), and a 75% yield was obtained? (12)



Answer

(a) The receiving vessel is cooled in ice as ethanal has a low boiling point. (5)

(b) Two features that are necessary to maximise the yield of ethanal are as follows:

- ⊗ By using excess ethanol (3) rather than excess sodium dichromate, further oxidation to ethanoic acid is kept to a minimum (3).
- ⊗ By distilling off the ethanal as it is formed (3), further oxidation of the ethanal to ethanoic acid is kept to a minimum (3).

(c) The colour change is from orange (3) to green (3).

The colour change is due to Cr(VI) changing to Cr(III). (3)

(d) To carry out Fehling's test on a sample of ethanal:

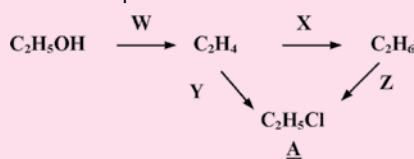
- ⊗ Mix Fehling's A with the same quantity of Fehling's B. (3)
- ⊗ Add ethanal to the Fehling's solution and place the test-tube in hot water for a few minutes. (3)

Observation: A brick red (3) precipitate (3) is observed.

(e) The solution of this problem is shown in Chapter 24.

2005 Question 7

Examine the reaction scheme and answer the questions which follow.



- (a) Name the compound labelled **A**. (5)
- (b) For each of the conversions, **W**, **X**, **Y** and **Z**, classify it as an *addition*, an *elimination* or a *substitution* reaction. (12)
- (c) Describe with the aid of a labelled diagram how the conversion **W** may be carried out in a school laboratory and how a sample of the product may be collected. How would you test this product to show that it is unsaturated? (18)
- (d) The conversion labelled **Z** is known to occur by a *free radical* mechanism. State **three** clear pieces of experimental evidence which support this mechanism. (15)

Answer

- (a) A = chloroethane (5)
- (b) W = elimination reaction (3)
 X = addition reaction (3)
 Y = addition reaction (3)
 Z = substitution reaction (3)
- (c) See Chapter 21 for the answer to this question.
- (d) Experimental evidence to support free radical mechanism:
- ☼ The initiation step requires ultraviolet light.
 - ☼ The reaction does not occur in the dark.
 - ☼ If tetraethyl lead is added there is an increase in the rate of the reaction.
 - ☼ Many thousands of molecules are produced for every photon that is absorbed.
 - ☼ Evidence for the combining of free radicals comes from the fact that small amounts of butane are found among the products.
- Any three of the above points: (2 × 6 + 3)

Note: The same mechanism holds for the chlorination of ethane as methane. The differences in terms of the products formed, radicals formed and the lead compound needed to increase the rate of reaction are described on page 414 of the textbook.

2005 Question 10

- (c) State the principle on which all chromatographic separation techniques are based. (10)
- Describe with the aid of clearly labelled diagrams how you could carry out an experiment to separate a mixture of dyes (or indicators) using paper, thin-layer or column chromatography. (15)

Answer

- (c) Chromatography is a separation technique in which a mobile phase (4) carrying a mixture moves in contact with a selectively adsorbent (3) stationary phase (3).

To separate a mixture of dyes (or indicators) using paper chromatography, the key points are as follows:

- ⊗ Place a small dot (spot) of mixture of dyes on paper. (3)
- ⊗ Just above solvent [or about 2 cm above bottom of sheet]. (3)
- ⊗ Paper placed in solvent in tank. (3)
- ⊗ Solvent moves up. Solvent front shown in diagram or described in statement. (3)
- ⊗ State or show separation of components of mixture. (3)

All of the above points may be obtained from a clearly labelled diagram.

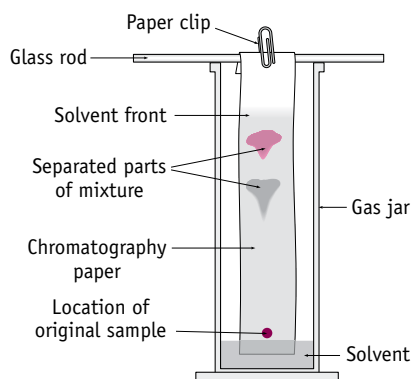
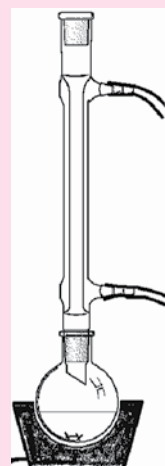


Fig. 22.1

TIP: It is always a good idea to draw a diagram when explaining how an experiment is carried out as many marks may be picked up for points which you may have forgotten to mention in your description of the experiment.

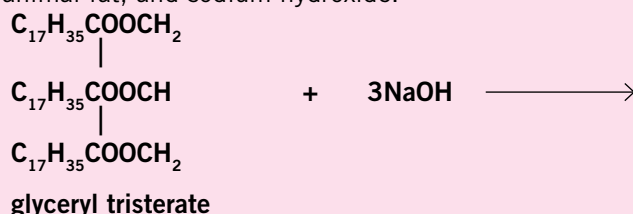
2006 Question 2

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 cm³ 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)

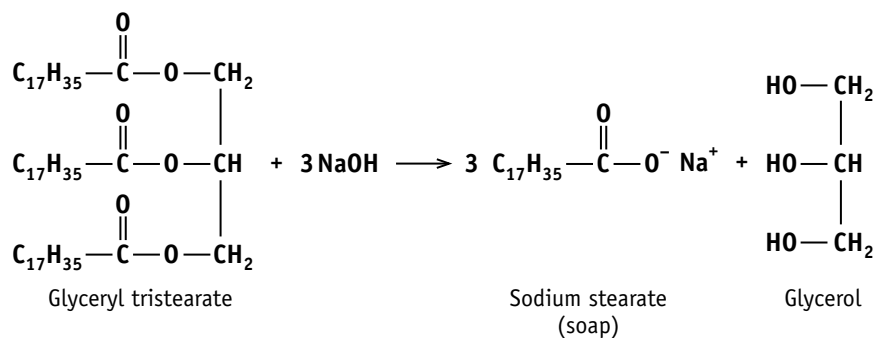
- (b) Complete and balance the equation below for the reaction between glyceryl tristearate, an animal fat, and sodium hydroxide. (9)



- (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
- of the second product of reaction,
 - of the excess sodium hydroxide? (6)
- (f) What would you observe, upon shaking, if a little of the soap prepared in the experiment is added to
- a test tube containing deionised water,
 - a test tube containing mineral water from a limestone region? (6)

Answer

- (a) Refluxing allows you to keep the liquid at its boiling point without loss of vapour [ethanol] (5). This prevents the flask boiling dry and allows sufficient time for the reaction to occur. The type of reaction that occurs is saponification [base catalysed hydrolysis]. (3)
- (b) The equation is as follows:



(9)

Products: (2 × 3); Balanced: (3)

Note: This is an important reaction. If you are unsure of the equation, study page 416 in your textbook.

- (c) The purpose of the ethanol is to act as a solvent for the lard. (3)
 Removing the ethanol after the reflux makes it easier to isolate the soap. (3)
 To remove the ethanol, set up the following apparatus:

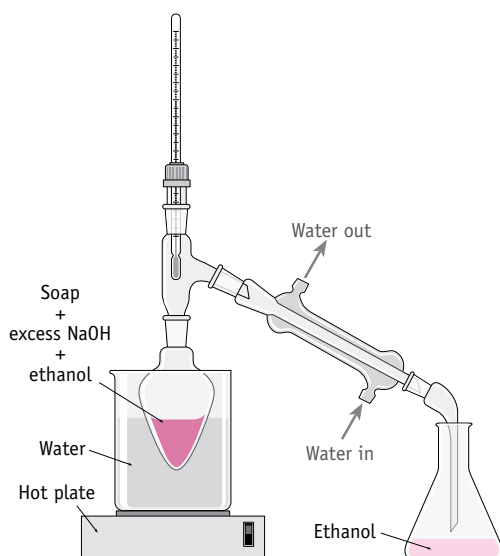


Fig. 22.2

The key points from the diagram are as follows:

- ⊗ Hot plate and thermometer.
- ⊗ Distilling flask.
- ⊗ Condenser (sloping down, showing inlet and outlet for water).
- ⊗ Collection in vessel.

Any two points: (2 × 3)

- (d) The key points in obtaining a pure sample of soap from the reaction mixture are as follows:
- ⊗ Pour the hot reaction mixture into a beaker containing brine.
 - ⊗ Filter off the soap using a filter funnel and filter paper.
 - ⊗ Wash the soap with fresh salt solution to remove any sodium hydroxide still present.

Brine: (6); Any other point: (3)

- (e) (i) The second product (glycerol) is in the brine. (3)
 (ii) The excess sodium hydroxide is in the brine. (3)
 (f) (i) When the soap is added to deionised water a lather is observed. (3)
 (ii) When the soap is added to water from a limestone region, no lather is observed but a scum is formed instead. (3)

2006 Question 4

- (g) What observation is made when a sample of ethanal is heated with Fehling's reagent?

Answer

- (g) A brick-red (3) precipitate (3) is observed.

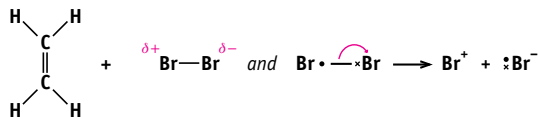
2006 Question 9

- (c) Describe the mechanism of the bromination of ethene. (9)
 State and explain **one** piece of experimental evidence to support the mechanism. (6)

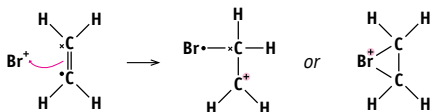
Answer

(c) The key points in the bromination of ethene are as follows:

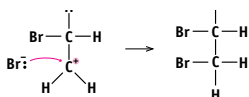
- ⊗ Polarisation of Br₂. (3)



- ⊗ Addition of Br⁺ ion across double bond. (3)



- ⊗ Attack on carbonium ion by Br⁻ ion. (3)

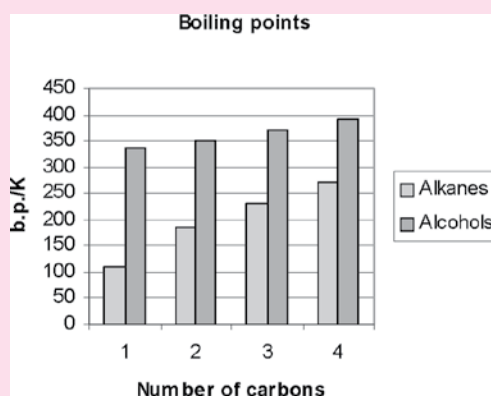


Evidence for the mechanism comes from the fact that when ethene reacts with bromine in water containing some sodium chloride (3), a compound called 1-bromo-2-chloroethane is formed. This is formed when the carbonium ion is attacked by Cl⁻ and is evidence for the existence of the carbonium ion (3).

2006 Question 10

(c) The chart compared 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)



Answer

- (c) (i) The two reasons why alcohols have higher boiling points than the corresponding alkanes are the following:
 ⊗ Hydrogen bonding exists between the alcohol molecules. (4)
 ⊗ Alcohols have a higher relative molecular mass. (3)
 (ii) The reason for the difference is because the hydrogen bonding is stronger in methanol (6) due to the shorter carbon chain, i.e. there is a smaller non-polar part of the molecule in methanol compared to butanol.

Exam Edge

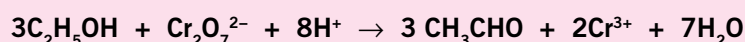
- (iii) Methane is insoluble in water. (3)
- Methanol is completely miscible (soluble) with water. (3)
- Butane is insoluble in water. (3)
- Butanol is slightly soluble in water [less soluble than methanol]. (3)

24 Stoichiometry II

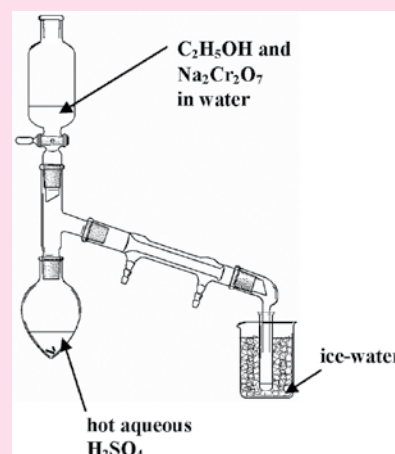
The questions in the *Stoichiometry II* chapter are based on the fundamental concepts of stoichiometry that are covered in Chapters 9–11, as well as the ideas of limiting reactant and percentage yield that are covered in this chapter. The questions on the examination paper are usually set in the context of the organic preparations which you have carried out as part of the Mandatory Student Experiments. Therefore, questions on this chapter can be given in both Section A and Section B of the examination paper.

2005 Question 2

A group of students prepared ethanal (CH_3CHO) by slowly adding an aqueous solution of ethanol ($\text{C}_2\text{H}_5\text{OH}$) and sodium dichromate(VI) ($\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$) to a hot aqueous solution of sulfuric acid (H_2SO_4). The apparatus drawn below was used. The reaction is described by the following equation.

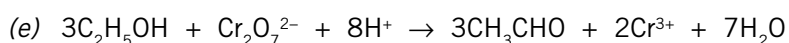


- (e) Assuming that all of the features needed to maximise the yield of ethanal were present, what mass of ethanal would be produced in the preparation if the students used 8.94 g of sodium dichromate(VI) ($\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O}$), and a 75% yield was obtained?



(12)

Answer

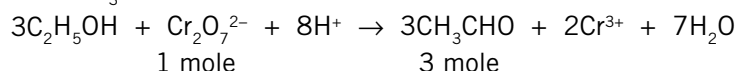


Rel. molecular mass of $\text{Na}_2\text{Cr}_2\text{O}_7 \cdot 2\text{H}_2\text{O} = 2(23) + 2(52) + 7(16) + 2(18) = 298$

$$\text{Number of moles of sodium dichromate} = \frac{\text{mass}}{\text{rel. molecular mass}} = \frac{8.94}{298}$$

$$= 0.03 \text{ moles} \quad (3)$$

Rel. molecular mass of $\text{CH}_3\text{CHO} = 12 + 3(1) + 12 + 1 + 16 = 44$



From balanced equation, 1 mole $\text{Cr}_2\text{O}_7^{2-} \rightarrow 3$ moles CH_3CHO

Therefore, 0.03 mole $\rightarrow 3 \times 0.03 = 0.09$ moles $\text{CH}_3\text{CHO} \quad (3)$

$$\begin{aligned} &= 0.09 \times 44 \text{ g} \\ &= 3.96 \text{ g} \end{aligned} \quad (3)$$

But, only 75% yield, therefore actual mass of ethanal = 75% (3.96)

$$= 2.97 \text{ g} \quad (3)$$

Answer: 2.97 g

Note: It is important that you show all the calculations. If you make an arithmetical slip, you will only lose one mark provided you have shown all calculations so that the examiner can see where the arithmetical slip has been made.

25 & 26 Industrial Chemistry and Atmospheric Chemistry (Option 1)

The *Industrial Chemistry* and *Atmospheric Chemistry* chapters are always examined as one of the parts of question 11, i.e. the equivalent of half a question (25 marks). In addition, there is always one short question given on these two chapters as part of question 4. The two chapters constitute OPTION 1, which is a far more popular option among students than OPTION 2 (Chapters 27 and 28).

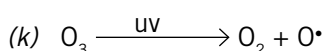
Atmospheric Chemistry (Chapter 26) is more commonly examined than *Industrial Chemistry* (Chapter 25). However, on any one year you cannot predict which chapter will be examined so it is important that you study both chapters.

2002 Question 4

(k) Write an equation for the photodissociation of ozone.

(6)

Answer



(6)

2002 Question 11 (c)

A Distinguish between a batch and a continuous production process.

(6)

Answer both of the following questions, (i) and (ii), in relation to one of the following processes:

ammonia manufacture

nitric acid manufacture

magnesium oxide manufacture

(i) In relation to your chosen chemical industry state one reason in favour of the Irish location of the industrial plant.

(4)

(ii) Give a brief outline of the processes carried out in the manufacture of the main product, giving balanced chemical equations where relevant.

(15)

Answer

A In a **batch process**, the raw materials are allowed to react for a fixed time with each other inside a vessel (batch reactor). The reaction takes place under controlled conditions of temperature and pressure for a given time. When the reaction is complete, the product is isolated.

(3)

In a **continuous process**, the raw materials are continually fed in at one end of the plant and the product is continually removed at the other end of the plant.

(3)

(i) **Reason for location:**

Manufacture of ammonia

- ⊗ The Cork-Cobh railway line was very near the plant. This was ideal for transport of ammonia.
- ⊗ The site was next to a deep water harbour for easy export of ammonia and urea by ship.
- ⊗ The site was near to the location of natural gas off Kinsale.
- ⊗ There was a good supply of skilled personnel (engineers, chemists, technicians, etc.) as the plant was only 20 km from Cork City, which has a university and an institute of technology.

Any one of the above reasons: (4)

Manufacture of nitric acid

- ⊗ The site was close to the Avoca river which supplied the plant with fresh water for cooling purposes.
- ⊗ The site was near a local harbour for ease of shipping export products.
- ⊗ The rail link connection to the main Dublin–Rosslare line made delivery of ammonia and distribution of fertiliser products very easy.

Any one of the above reasons: (4)

Manufacture of magnesium oxide

- ⊗ The site is close to a plentiful supply of seawater.
- ⊗ The site is close to a limestone quarry.
- ⊗ The plant is built next to the river Boyne for easy export of the product.
- ⊗ There is a good road network to allow easy transport of limestone to the plant.

Any one of the above reasons: (4)

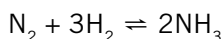
(ii) Brief outline of processes:

Manufacture of ammonia

The main points are as follows:

- ⊗ Hydrogen from methane
- ⊗ Nitrogen from air
- ⊗ $\text{CH}_4 + \text{H}_2\text{O} \rightleftharpoons \text{CO} + 3\text{H}_2$
- ⊗ $\text{CO} + \text{H}_2\text{O} \rightleftharpoons \text{CO}_2 + \text{H}_2$
- ⊗ Iron catalyst
- ⊗ High pressure (200 atm)
- ⊗ High temperature (500 °C)

Any four of the above points: (4 × 3)
(3)



Manufacture of nitric acid

The main points are as follows:

- ⊗ $4\text{NH}_3 + 5\text{O}_2 \longrightarrow 4\text{NO} + 6\text{H}_2\text{O}$
- ⊗ Platinum-rhodium catalyst
- ⊗ High temperature (900 °C)
- ⊗ High pressure (2–15 atm)
- ⊗ $2\text{NO} + \text{O}_2 \longrightarrow 2\text{NO}_2$
- ⊗ $3\text{NO}_2 + \text{H}_2\text{O} \longrightarrow 2\text{HNO}_3 + \text{NO}$

Any four of the above points: (4 × 3)

Manufacture of magnesium oxide

The main points are as follows:

- ⊗ Lime produced from limestone
- ⊗ $\text{CaCO}_3 \longrightarrow \text{CaO} + \text{CO}_2$
- ⊗ Quicklime with water gives lime
- ⊗ $\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2$
- ⊗ Reaction with MgCl_2 in seawater
- ⊗ $\text{Mg(OH)}_2 \longrightarrow \text{MgO} + \text{H}_2\text{O}$

Any four of the above points: (4 × 3)



(3)

2003 Question 4 (k)

A Describe with the aid of an equation how *nitrogen fixation* occurs in nature. (6)

Answer

A $\text{N}_2 + \text{O}_2 \longrightarrow 2\text{NO}$ (6)

2003 Question 11 (c)

A What is the chemical formula for ozone? State **one** beneficial effect of the ozone layer. (7)

CFCs are believed to be the main cause of the damage to the ozone layer.

(i) What are CFCs? What use is made of CFCs? (6)

(ii) Explain how CFCs may give rise to ozone depletion. (12)

Answer

A The chemical formula for ozone is O_3 . (4)

The ozone layer absorbs harmful UV radiation from the sun. (3)

(i) CFCs are chlorofluorocarbons, i.e. compounds of chlorine, fluorine and carbon only. (3)

CFCs are used in fridges. (3)

(ii) The main points are as follows:

⊗ Release of chlorine free radicals. (3)

⊗ Chlorine free radicals attack ozone $\text{O}_3 + \text{Cl}^\bullet \longrightarrow \text{O}_2 + \text{ClO}^\bullet$ (3)

⊗ Reaction with oxygen free radicals to give chain reaction. (3)

⊗ Regeneration of chlorine free radical. (3)

Or, last two points can be summarised by the equation:

$\text{ClO}^\bullet + \text{O}^\bullet \longrightarrow \text{Cl}^\bullet + \text{O}_2$ (6)

2004 Question 4 (k)

A How is oxygen gas produced industrially? (3)

Answer

A Oxygen is manufactured by the liquefaction and fractional distillation (3) of the air.

2004 Question 11 (c)

B The *greenhouse effect* is a natural phenomenon but its effects have been enhanced by human activity over the past 200 years.

(i) Explain the term *greenhouse effect*. (7)

(ii) Identify **one** gas in the atmosphere which makes a significant contribution to the greenhouse effect. (3)

(iii) In relation to the gas you have identified in (ii), mention a type of human activity which has been a major contributor to the increased levels of this gas in the atmosphere. (3)

(iv) Identify **one** gas, found in the atmosphere, which is not a greenhouse gas. (3)

(v) State **three** probable consequences of an increased greenhouse effect which have been suggested by environmental scientists. (9)

Answer

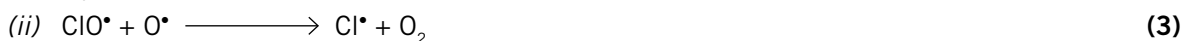
- (i) The greenhouse effect is the trapping of the sun's energy **(4)** by the atmosphere **(3)**.
- (ii) Carbon dioxide is a gas in the atmosphere that makes a significant contribution to the greenhouse effect. **(3)**
- (iii) The burning of fossil fuels has been a major contributor to the increased levels of carbon dioxide in the atmosphere. **(3)**
- (iv) Nitrogen [and oxygen] are found in the atmosphere and are not greenhouse gases. **(3)**
- (v) Three probable consequences of an increased greenhouse effect are global warming, a rise in sea levels and climate change. **(3 × 3)**

2005 Question 4 (k)

- A Describe using chemical equations the chain reaction process whereby chlorine free radicals break down ozone in the stratosphere.

Answer

- A The chemical equations are as follows:



Note: If you are unsure about how the chain reaction works, study pages 102–103 in Chapter 26 of your workbook.

2005 Question 11 (c)

- B (i) Account for the inert nature of nitrogen gas. **(7)**
- (ii) What is meant by *nitrogen fixation*?
State **two** ways by which nitrogen is fixed in nature. **(9)**
- (iii) The concentration of **NO₂** in the atmosphere has increased in the past fifty years. Describe with the aid of chemical equations how an increase in the number of cars has contributed to this change. **(9)**

Answer

- (i) Nitrogen gas is inert because of the large amount of energy needed to break the triple bond in the nitrogen molecule. **(4)**
Also, nitrogen is a non-polar molecule and this means it is only slightly soluble in water. **(3)**
- (ii) Nitrogen fixation is the conversion of atmospheric nitrogen to compounds which can be used by plants. **(3)**
In nature, nitrogen fixation takes place in two ways. Firstly, during thunderstorms, the discharge of electricity **(3)** provides enough energy for the nitrogen and oxygen in the air to react together. Secondly, nitrogen-fixing bacteria **(3)** have the ability to fix nitrogen.
- (iii) Oxides of nitrogen are produced in car engines. At high temperatures, the nitrogen in the air combines with oxygen to form nitrogen monoxide. The high temperature is produced by the spark plug **(3)**.
The equations are as follows:



2006 Question 4 (k)

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

Answer

- A Nitrogen is used in flushing out oil tanks. (3)
It is also used in preserving food, e.g. crisps. (3)

2006 Question 11 (c)

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)

Answer

Ammonia manufacture

- (i) The raw materials are air (3) and natural gas (3).
How the raw materials are treated:
⊗ Air: Natural gas is burned in air to obtain nitrogen. (3)
⊗ Natural gas: Natural gas reacts with steam [methane steam-reforming] to give hydrogen. (3)
(ii) One product which could cause pollution: ammonia. (3)
(iii) Most important use of the main product: Ammonia is used in fertilisers. (6)
Ammonia is particularly suitable for use as fertilisers as it contains nitrogen. (4)

Nitric acid manufacture

- (i) The raw materials are ammonia (3) and air (3).
How the raw materials are treated:
⊗ Ammonia: No treatment is necessary for ammonia. (3)
⊗ Air: Air is filtered to remove dust. (3)
(ii) One product which could cause pollution: nitric acid. (3)
(iii) Most important use of the main product: nitric acid is used in the manufacture of fertilisers. (6)
Nitric acid is particularly suitable for use as fertilisers as it contains nitrogen. (4)

Magnesium oxide manufacture

- (i) The raw materials are limestone (3) and seawater (3).
How the raw materials are treated:
⊗ Limestone: Limestone is heated and converted to lime and water is added to form slaked lime. (3)
⊗ Seawater: Seawater is purified – sand, silt and carbon dioxide are removed. (3)
(ii) One product which could cause pollution: magnesium oxide. (3)
(iii) Most important use of the main product: Magnesium oxide is used as a heat-resistant material for lining furnaces. (6)
Magnesium oxide is particularly suitable for this use as it has a high melting point. (4)

27 & 28 Materials and Electrochemistry II (Option 2)

The *Materials* and *Electrochemistry II* chapters are always examined as one of the parts of question 11, i.e. the equivalent of half a question (25 marks). In addition, there is always one short question given on these two chapters as part of question 4. The two chapters constitute OPTION 2 which is a far less popular option among students than OPTION 1 (Chapters 25 and 26).

The examination papers reflect a good balance of questions between Chapters 27 and 28. However, on any one year you cannot predict which chapter will be examined so it is important that you study both chapters.

2002 Question 4 (k)

B What are the structural differences between low-density and high-density poly(ethene)? (6)

Answer

B High-density polythene has very little branching (3) along the chains, and has longer chains (3) than low-density polythene.

2002 Question 11 (c)

B Aluminium is extracted from bauxite.

(i) Where in Ireland is bauxite purified to produce alumina? (4)

(ii) Outline the steps involved in the extraction of alumina from bauxite, giving balanced chemical equations where relevant. (15)

(iii) Give two reasons why it is preferable to produce aluminium by recycling rather than by extracting it from its ore. (6)

Answer

(i) Bauxite is purified at Aughinish [or Shannon estuary]. (4)

(ii) The main steps are as follows:

⊗ Digestion with sodium hydroxide to produce sodium aluminate.

⊗ $\text{Al}_2\text{O}_3 \cdot 3\text{H}_2\text{O} + 2\text{NaOH} \longrightarrow 2\text{NaAlO}_2 + 4\text{H}_2\text{O}$ (3)

⊗ Crushing and mixing.

⊗ Clarification (removal of impurities).

⊗ Precipitation of aluminium oxide.

⊗ $2\text{NaAlO}_2 + 4\text{H}_2\text{O} \longrightarrow \text{Al}_2\text{O}_3 \cdot 3\text{H}_2\text{O} + 2\text{NaOH}$

⊗ Removal of water of crystallisation.

⊗ $\text{Al}_2\text{O}_3 \cdot 3\text{H}_2\text{O} \longrightarrow \text{Al}_2\text{O}_3 + 3\text{H}_2\text{O}$

Any four points: (4 3)

(iii) Producing aluminium by recycling cans rather than by extracting it from its ore is cheaper (3) and reduces litter (3) from aluminium cans.

2003 Question 4 (k)

B State **two** properties of transition metals.

Answer

B Transition metals have variable valency **(3)** and act as catalysts **(3)**.

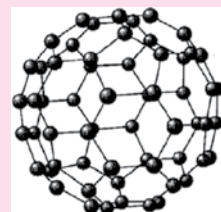
2003 Question 11 (c)

B The structure of a buckminsterfullerene with 60 atoms is drawn on the right.

(i) Atoms of what element make buckminsterfullerenes? **(4)**

(ii) Name **two** other covalent macromolecular crystals formed by this element. What are the binding forces in each of these crystals? **(15)**

Give **one** use for **each** of these two substances. **(6)**



Answer

(i) Carbon is the element. **(4)**

(ii) Other covalent macromolecular crystals formed by this element are diamond **(3)** and graphite **(3)**.

The forces in diamond are covalent bonds. **(3)**

In graphite, the forces within the sheets of carbon atoms are covalent **(3)** and between the sheets are van der Waals forces **(3)**.

Diamond is used in jewellery [and for cutting]. **(3)**

Graphite is used in pencils [and batteries]. **(3)**

2004 Question 4 (k)

B How does the anodising of aluminium protect it from corrosion? **(6)**

Answer

B The anodising forms a thick protective layer **(3)** of oxide on the surface of the aluminium **(3)**.

2004 Question 11 (c)

A Write a brief note on the contribution made to our understanding of crystal structures by

(i) Lawrence and William Bragg, **(7)**

(ii) Dorothy Hodgkin. **(6)**

What type of crystal is formed by iodine and what are the binding forces in the crystal?

Explain

(i) Why metals are generally good conductors of electricity, **(12)**

(ii) why most ionic crystals dissolve in water.



Lawrence and William Bragg



Dorothy Hodgkin

Answer

(i) Lawrence and William Bragg used X-rays to determine the arrangement of particles within crystals [or developed X-ray crystallography]. (4)

(ii) Dorothy Hodgkin carried out research into the crystal structures of complex organic molecules [e.g. vitamin B₁₂]. (3)

Iodine forms a molecular crystal. (3)

The binding forces within the crystal are van der Waals forces. (3)

(i) Metals are good conductors of electricity because the outer electrons (3) are delocalised (3) over the entire crystal lattice and the electric current can easily pass through the “sea of electrons” in the metal.

(iii) Most ionic crystals dissolve in water because of the attraction of the positive and negative ions in the crystal for the polar water molecules (3); this force overcomes the attractions between ions in the crystal (3).

2005 Question 4 (k)

B How does a sacrificial anode protect a metal from corrosion?

Answer

B The sacrificial anode is oxidised [corrodes] more easily than the protected metal does. (6)

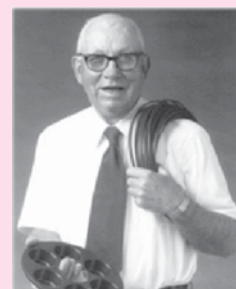
2005 Question 11 (c)

A (i) What is meant by the term *addition polymerisation*? (7)

(ii) Name the Du Pont chemist pictured on the right who discovered poly(tetrafluoroethene), PTFE. (3)

(iii) Describe using an equation how poly(tetrafluoroethene) is produced from its monomers. (9)

(iv) Give **two** common uses of PTFE. (6)



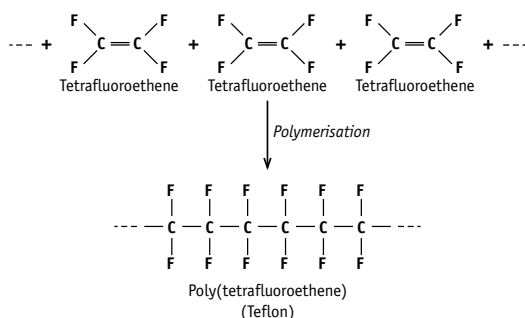
Discoverer of PTFE

Answer

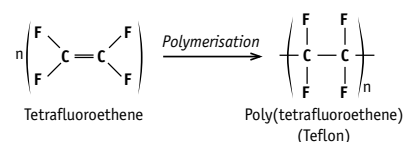
(i) Addition polymerisation occurs when monomers combine (4) to form a polymer (3).

(ii) Roy Plunkett discovered Teflon. (3)

(iii) The production of polytetrafluoroethene from its monomers may be represented as the following:



This may also be written as:



(9)

(iv) PTFE is used for the non-stick surface (3) on frying pans and also medical body parts such as knee joints and heart pacemakers (3).

2006 Question 4 (k)

- 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. (6)

Answer

- B An example of the main group metal is aluminium [or magnesium]. (3)
An example of the transition metal is nickel [or zinc]. (3)

2006 Question 11 (c)

- B A blast furnace may be used in the extraction of iron from iron ore. (12)
(i) What materials must be added to a blast furnace in operation? (3)
(ii) Name the principal reducing agent in the blast furnace and write a balanced equation for its reaction with haematite (Fe_2O_3). (9)
(iii) Why is the pig iron produced in a blast furnace further processed into steel? (4)

Answer

- (i) The materials added to a blast furnace in operation are iron ore (3), coke (3), limestone (3) and hot air (3).

TIP: Categorising hot air as a material rather than a condition for the reaction to occur is unusual. The fact that 12 marks have been allocated should alert you to the fact that four items were required.

- (ii) The principal reducing agent is carbon monoxide. (3)
 $\text{Fe}_2\text{O}_3 + 3\text{CO} \longrightarrow 2\text{Fe} + 3\text{CO}_2$ Correct formulas: (3); Balanced: (3)
(iii) Most of the pig iron is converted into steel since there is a far greater demand for steel than for pig iron. (4)