MYP Revision booklet

Chemistry

2019-2020

LIST OF CONTENT

CHEMISTRY GLOSSARY

1. Acid: A substance that produces hydrogen ions (H+) as the only positive ions when dissolved in water.
2. Activation energy: The minimum energy that molecules must possess during their collisions in order for a chemical reaction to occur.
3. Addition reaction: A reaction in which a molecule (element or compound) adds to an unsaturated compound to form a single new compound.
4. Alcohol: An organic compound containing the hydroxyl group, -OH.
5. Alkali: A base that is soluble in water.
6. Alkali metals: The elements in Group 1 of the Periodic Table.
7. Alkane: Hydrocarbon having the general formula CnH2n+2
8. Alkene: Hydrocarbon that contains one or more carbon-carbon double bonds. Alkenes with only one carbon-carbon double bond have the general formula CnH2n.
9. Alloy: A mixture of a metal with non-metals or other metals.
10. Anhydrous: Anhydrous salts are salts without water of crystallization.
11. Anion: A negatively charged ion which moves towards the anode during electrolysis.
12. Anode: A positively charged electrode in an electrolytic cell.
13. Aqueous: Describing the solution of a substance in water, i.e. the aqueous solution. In chemical equations, aqueous solutions are represented by the symbol (aq).
14. Atom: The smallest particle of an element.
15. Avogadro's constant: The number of particles in one mole of a substance. Its value is 6 x 1023.
16. Avogadro's law: At constant temperature, the volume of a gas is directly proportional to the number of moles of the gas present.
17. Base: A substance that reacts with an acid to form a salt and water only.
18. Boiling point: The temperature at which a liquid turns rapidly to its vapour.
19. Carboxylic acid: An organic acid containing the carboxyl group, -COOH.
20. Cathode: A negatively charged electrode in an electrolytic cell.
21. Cation: A positively charged ion which moves towards the cathode during electrolysis.
22. Chromatography: A method of separating the components in a mixture.
23. Collision theory: A chemical reaction can occur only if the reacting particles collide with one another.
24. Combustion: The chemical name for burning. Burning occurs when a substance reacts very rapidly with oxygen.
25. Compound: A substance formed in a chemical change when two or more elements are joined together.
26. Condensation: The process by which a vapour or a gas turns to liquid on cooling.
27. Corrosion: The wearing away of the surface of a metal by chemical reaction.
28. Covalent bond: The type of bond formed when electrons are shared between two non-metal atoms.
29. Cracking: The breaking down of long chain hydrocarbon molecules with heat and/or catalyst to produce smaller hydrocarbon molecules and/or hydrogen.
30. Decomposition: A chemical reaction that results in the breaking down of a compound into two or more components.
31. Diatomic molecule: A molecule that consists of two atoms.
32. Displacement reaction: A reaction in which an atom or molecule takes the place of another atom or molecule in a compound.
33. Distillation: A process of obtaining the pure solvent from a solution. When the solution is boiled, the solvent is vaporized and the vapour condenses to reform the pure liquid.

34. Electrode: A rod or a plate which carries electricity in or out of an electrolyte during electrolysis.

35. Electrolysis: A process in which electrical energy is used to cause a chemical reaction to occur, typically to separate the electrolyte into its elements.

36. Electron: A negatively charged sub-atomic particle that surrounds the nucleus of an atom.

37. Electronic configuration: The arrangement of electrons in the various shells of an atom or a molecule.

38. Element: A substance made from only one type of atom. It cannot be separated into simpler substances by chemical processes or by electricity.

39. Endothermic reaction: A reaction which absorbs heat from the surroundings.

40. Evaporation: The process by which a liquid changed to its vapour on the surface of the liquid.

41. Exothermic reaction: A process that gives off heat to the surroundings.

42. Fermentation: The conversion of glucose by microorganisms such as yeast into ethanol and carbon dioxide.

43. Filtrate: The clear liquid which passes through the filter during filtration.

44. Filtration: The process of separating a solid from a liquid or a solution.

45. Fossil fuels: Fuels produced many millions of years ago from the decaying remains of animals or plants, includes oil, natural gas and coal.

46. Fractional distillation: A process that separates the components in a mixture on the bases of their different boiling points. The component with the lowest boiling point boils off first and is distilled over.

47. Freezing point: The temperature at which a liquid changes to a solid.

48. Fuel: A substance that burns easily to produce energy.

49. Functional group: An atom or group of atoms that gives characteristic properties to an organic compound.

50. Giant structure: A three-dimensional network of atoms or ions packed together in a regular pattern.

51. Group: A vertical column of elements in the Periodic Table.

52. Halogen: The non-metallic elements in Group VII (7) of the Periodic Table.

53. Homologous series: A family of organic compounds with members of the family having the same functional group and similar chemical properties.

54. Hydrated salts: Salts that contain water of crystallization.

55. Hydrocarbons: Organic compounds made up from the elements hydrogen and carbon only.

56. Hydrogenation: The addition of a hydrogen molecule across a double bond.

57. Immiscible: Two liquids that do not mix.

58. Indicators: Compounds that have distinctly different colours in acidic and alkaline solutions. 59. Ion: A positively or negatively charged particle. It is formed when an atom or group of atoms loses or gains electrons.

60. Ionic bond: The electrostatic force that holds positive and negative ions together in an ionic compound.

61. Isotopes: Atoms of the same element that have the same atomic number but different mass/nucleon number.

62. Melting point: The temperature at which a solid changes to a liquid.

63. Metal: An element that is shiny and conducts electricity in the solid state. Metals burn in oxygen to form basic oxides or amphoteric oxides.

64. Mixture: A substance made by mixing other substances together. The components in a mixture can be easily separated by physical methods because they are not chemically joined together like in compounds.

65. Mole: The amount of a substance which contains 6 x 1023 particles.

66. Molecule: A group of atoms held together by covalent bonds. Molecules may be elements or compounds.

67. Nucleon number: Also known as the mass number. It is the sum of the number of protons and neutrons in the nucleus of an atom.

68. Neutralization: The reaction between an acid and a base to produce a salt and water only.

69. Neutron: A sub-atomic particle in the nucleus of an atom. It has a mass but no electrical charge.

70. Organic chemistry: The branch of chemistry that deals with carbon compounds.

71. Oxidation: A reaction where a substance gains oxygen or loses hydrogen. Oxidation is also defined as the loss of electron(s) or the increase in the oxidation state of the element.

72. Oxides: Compounds of an element with oxygen.

73. Oxidizing agent: A substance that brings about oxidation. It is itself reduced. An oxidizing agent is an acceptor of electrons.

74. Period: A horizontal row of elements in the Periodic Table.

75. Periodic table: A table that contains horizontal rows and vertical columns of elements. The elements are arranged in order of their atomic numbers and in accordance with their chemical properties.

76. pH scale: A scale that measures the acidity or alkalinity of a solution.

77. Pollution: The presence in the environment of toxic substances which are harmful to living things.

78. Polymer: A very large molecule built up of a number of repeating units called monomers.

79. Polymerization: A chemical reaction in which simple molecules, called monomers, react with each other to form larger molecules called polymers.

80. Polyunsaturated: Vegetable oils that contain many carbon-carbon double bonds in their molecules.

81. Precipitate: An insoluble solid that is produced in a solution as a result of a chemical reaction.

82. Protein: A polymer of amino acids.

83. Protons: Positively charged sub-atomic particles found in the nucleus of an atom.

84. Proton number: The number of protons in the nucleus of an atom.

85. Pure substance: A single substance which is not mixed with other substances. It has definite melting and boiling points. (e.g. pure water boils at exactly 100°C and freezes at 0°C)

86. Reactivity series: A list of elements in order of their reactivity. The more reactive the element, the higher its position in the series. An element higher up the series will displace a less reactive one from a solution of its salt.

87. Redox reaction: A reaction where both oxidation and reduction take place simultaneously.

88. Reducing agent: A substance that brings about reduction. It is itself oxidized. A reducing agent is a donor of electrons.

89. Reduction: The removal of oxygen, the addition of hydrogen, the gain of electrons, or the decrease in the oxidation state of the substance.

90. Relative atomic mass: The number of times the mass of one atom of an element is heavier than 1/12 of the mass of a carbon-12 atom.

91. Relative molecular mass: The sum of the relative atomic masses of each of the atoms in one molecule of a substance.

92. Residue: The solid which remains on the filter paper after filtration.

93. Respiration: The slow combustion of food in the cells of living organisms to release energy.

94. Rusting: The slow oxidation of iron in the presence of air and water to form hydrated iron (III) oxide (rust).

95. Salt: The ionic compound formed by the replacement of one or more hydrogen ions of an acid by a metallic ion or an ammonium ion.

96. Saturated hydrocarbons: Hydrocarbons that contain only single bonds between carbon atoms.

97. Solute: The substance that dissolves in a solvent to form a solution.

98. Solvent: The liquid in which a solute dissolves.

99. Steel: An alloy of iron and carbon.

100. Structural formula: A formula which shows how the atoms are arranged in a molecule.

101. Sublimation: The process of changing from the solid state directly to the gaseous state without passing through the liquid state.

102. Suspension: A mixture of a liquid and an insoluble solid where the insoluble solid remains suspended throughout the solution.

103. Titration: The gradual addition of a solution from a burette to another solution in a conical flask until the chemical reaction between the two solutions is complete; the 2 solutions tend to be an acid and an alkali.

104. Unsaturated molecule: Any hydrocarbon that contains one or more carbon-carbon double bonds.

105. Valence electrons: Electrons in the outer shell that are used by the atom for forming chemical bonds.

106. Water of crystallization: Water molecules that are chemically bonded in the crystals of some salts.

AN UNDERSTANDING OF KEY CONCEPTS

MYP Chemistry requires you to understand 3 key concepts

System

Relationship

Change

Systems

Systems are sets of interacting or interdependent components. Systems provide structure and order in human, natural and built environments. Systems can be static or dynamic, simple or complex.

Systems in sciences describe sets of components that function due to their interdependence or complementary nature. Common systems in science are closed systems, where resources are not removed or replaced, and open systems, where necessary resources are renewed regularly. Modelling often uses closed systems to simplify or limit variables.

Other key concepts can also be important in sciences. For example, development is an important aspect in the continual growth through change that epitomizes scientific knowledge. Science offers important perspectives on the definition, measurement and meaning of time, place and space. Creativity is always important for scientists working together to extend the limits of human understanding.

The **system** is the part of the universe we wish to focus our attention on. In the world of chemistry, the system is the chemical reaction. For example:

$2H_2 + O_2$ ---> $2H_2O$

The system consists of those molecules which are reacting.

The **surroundings** are everything else; the rest of the universe. For example, say the above reaction is happening in gas phase; then the walls of the container are part of the surroundings.

There are two important issues:

1) a great majority of our studies will focus on the change in the amount of energy, not the absolute amount of energy in the system or the surroundings.

2) regarding the direction of energy flow, we have a "sign convention."

Two possibilities exist concerning the flow of energy between system and surroundings:
1) The system can have energy added to it, which (obviously, I hope) increases its amount and lessens the energy amount in the surroundings.

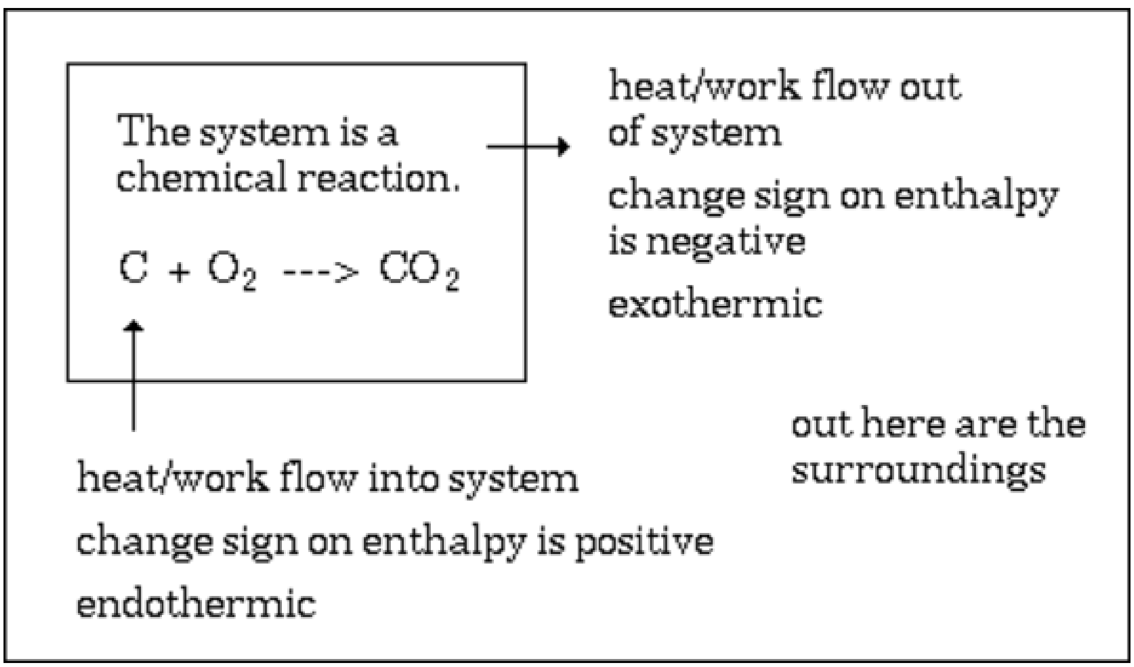
2) The system can have energy removed from it, thereby lowering its amount and increasing the amount in the surroundings.

We will signify an increase in energy with a positive sign and a loss of energy with a negative sign.

Also, we will take the point-of-view from the system. Consequently:

1) When energy (heat or work) flow out of the system, the system decreases in its amount. This is assigned a negative sign and is called exothermic.
2) When energy (heat or work) flows into the system, the system increases its energy amount. This is assigned a positive sign and is called endothermic.

We do not discuss chemical reactions from the surrounding's point-of-view. Only from the system's.



Relationships

Relationships are the connections and associations between properties, objects, people and ideas—including the human community's connections with the world in which we live. Any change in relationship brings consequences—some of which may occur on a small scale, while others may be far reaching, affecting large networks and systems such as human societies and the planetary ecosystem.

Relationships in sciences indicate the connections found among variables through observation or experimentation. These relationships also can be tested through experimentation. Scientists often search for the connections between form and function. Modelling is also used to represent relationships where factors such as scale, volume of data, or time make other methods impractical.

Change

Change is a conversion, transformation or movement from one form, state or value to another. Inquiry into the concept of change involves understanding and evaluating causes, processes and consequences.

In sciences, change is viewed as the difference in a system's state when observed at different times. This change could be qualitative (such as differences in structure, behaviour, or level) or quantitative (such as a numerical variable or a rate). Change can be irreversible, reversible or self-perpetuating.

Chemical changes occur when a substance combines with another to form a new substance, called chemical synthesis or, alternatively, chemical decomposition into two or more different substances. These processes are called chemical reactions and, in general, are not reversible except by further chemical reactions. Some reactions produce heat and are called exothermic reactions and others may require heat to enable the reaction to occur, which are called endothermic reactions. Understanding chemical changes is a major part of the science of chemistry.

When chemical reactions occur, the atoms are rearranged and the reaction is accompanied by an energy change as new products are generated. An example of a chemical change is the reaction between sodium and water to produce sodium hydroxide and hydrogen. So much energy is released that the hydrogen gas released spontaneously burns in the air. This is an example of a chemical change because the end products are chemically different from the substances before the chemical reaction.

List of Content under each concept

- Periodic table (metals and non-metals; transition metals, noble gases, trends, periods, groups)
- International Union of Pure and Applied Chemistry (IUPAC naming and classification of: alkanes, alkenes, alcohols, carboxylic acids and esters; structural formulas)
- The atmosphere (characteristics of gases; atmospheric composition, testing and treatment; extraction, emission and environmental implications)
- Matter (states and properties of matter; particle/kinetic theory, diffusion; atomic structure [including isotopes]; electron configuration and valency)
- Pure and impure substances (types of mixtures [solutions, oils, alloys, emulsions]; separation techniques, including: filtration, distillation [including crude oil], chromatography)
- Bonding (structure and bonding, properties, chemical formulas, chemical reactions and the conservation of mass; balancing equations, the mole concept and chemical calculations; reaction kinetics [rates, and factors affecting rates/collision theory]; equilibria/reversible reactions; energy changes in reactions, endo- and exothermicity; combustion of fuels)
- Types of chemical reaction (acids and bases, neutral solutions, acid/base reactions, pH and indicators, formation of salts, uses of salts; redox reactions, reactivity series; extraction of metals, and corrosion, electrochemical cells)

Criteria A

The particulate nature of matter

SUB TOPICS –

- The states of matter
- Their inter conversion in terms of the kinetic particle theory
- Diffusion
- Evidence for the movement of particles in gases and liquids
- Dependence of rate of diffusion on molecular mass.

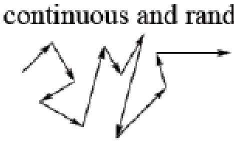
Kinetic Theory

The properties of solids, liquids and gases can be explained by kinetic theory.
Kinetic theory states that matter is made of tiny particles that move all the time.
The main points of the theory are;

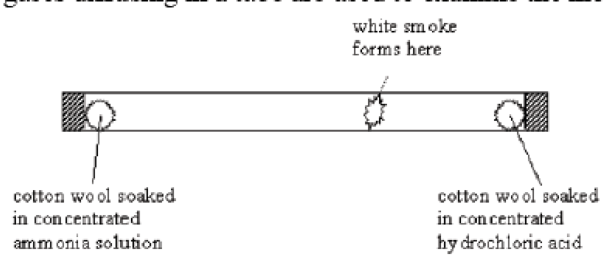
- All matter is made of tiny, invisible, moving particles.
- The particles move all the time. The higher the temperature the faster they move.
- Heavier particles move more slowly than light ones at a given temperature.

DIFFUSION: this is the mixing of atoms or molecules due to their continuous and random motion. e.g. Mixing of bromine vapour and air.

BROWNIAN MOTION: the constant random movement of tiny particles (e.g. smoke particles, or pollen on a drop of water) is caused by collision with (invisible) air or water molecules, which are themselves in continuous and random motion. e.g.



Experiments using gases diffusing in a tube are used to examine the motion of the particles.

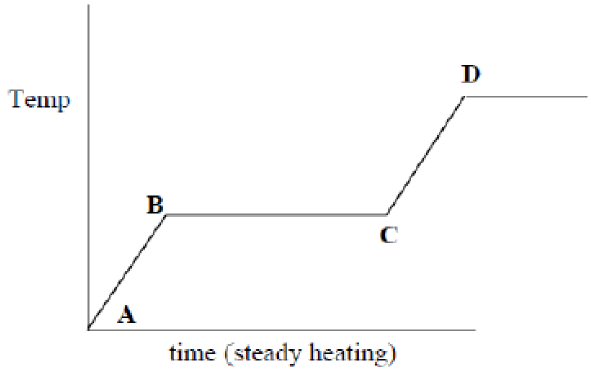


Ammonia and hydrochloric acid particles under Brownian motion as they hit air particles in the tube.
When they meet white smoke (ammonium chloride) forms.
As the ammonia travels farthest along the tube we know that;

- The ammonia particles move faster than hydrochloric acid particles.
- The ammonia particles are lighter than the hydrochloric acid particles.

Heating and Cooling Curves

We can use kinetic theory to explain changes in state when substances are heated or cooled.



- In a solid the particles are held in position by bonding to their neighbours.
  (A on the graph).
- As the solid is warmed the particles vibrate but cannot move.
  (Between A and B on the graph).
- When heated enough the particles vibrate so much that they can tear themselves free from their positions, and the substance melts to a liquid.
- At the melting point heat needs to be added to melt the solid and break the bonds between the particles, so the temperature doesn't rise until it is all liquid.
  (Between B and C on the graph).
- As the liquid is heated up the particles gain energy.
  (Between C and D on the graph).
- At the boiling point heat needs to be added to change break the forces between the particles in the liquid turning the substance into a gas.

In which substance are the particles close together and slowly moving past each other?

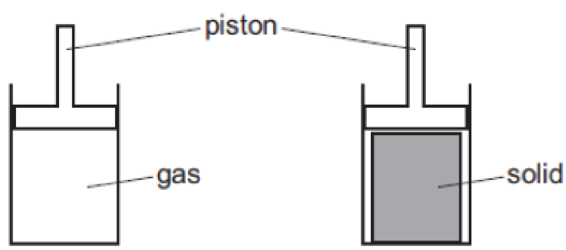
A air
B ice
C steam
D water

What are the processes W, X, Y and Z in the following diagram?

$$\text{solid} \underset{Y}{\overset{W}{\rightleftharpoons}} \text{liquid} \underset{Z}{\overset{X}{\rightleftharpoons}} \text{gas}$$

	W	X	Y	Z
A	condensing	boiling	freezing	melting
B	condensing	freezing	melting	boiling
C	melting	boiling	freezing	condensing
D	melting	freezing	condensing	boiling

An attempt was made to compress a gas and a solid using the apparatus shown.



Which substance would be compressed and what is the reason for this?

	substance	reason
A	gas	the gas particles are close together
B	gas	the gas particles are far apart
C	solid	the solid particles are close together
D	solid	the solid particles are far apart

Experimental techniques

SUB TOPICS –

- Measurement - Names of Lab apparatus used for measuring
- Criteria of purity - chromatography and importance of purity.
- Method of purification - using suitable solvent, filtration, crystallization, distillation and fractional distillation.

Elements, Mixtures and Compounds

THE ATOMIC THEORY

This theory assumes that all elements are made up of "atoms". If you were to divide a lump of an element into smaller and smaller pieces you would eventually come to a piece that could not be divided any further - a single ATOM of the element. Atoms are therefore very small. We can see this if we dilute a solution of potassium manganate(VI) many times. It is still coloured even when it is very dilute.

Definition: An atom is the smallest particle of an element that can exist or take part in a chemical change.

MOLECULES

All elements are made up of atoms. In some gaseous elements (e.g. argon) single atoms move around freely. But in other gaseous elements, single atoms cannot exist on their own at ordinary temperatures: in these elements the free-moving particles consist of pairs of atoms.
The two atoms forming a pair (a MOLECULE) are joined together by a chemical "bond". This is the case with hydrogen (H_2), oxygen (O_2) and nitrogen (N_2). Such substances are said to be diatomic.

Gaseous Argon Atoms — Gaseous Chlorine Molecules

Ar — Cl_2

An **ELEMENT** is a pure substance made up of only one type of ATOM.

A **COMPOUND** is a pure substance which contains two or more elements, chemically bonded together in a fixed proportion.

A **MIXTURE** is a group of substances that are not chemically bonded together.

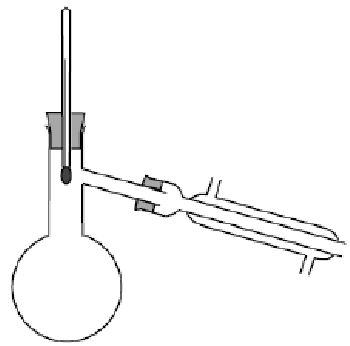
Compounds	Mixtures
Proportions of elements are fixed.	Proportions may vary.
Properties different from those of the elements.	Properties are simply those of the separate elements.
Cannot be separated into elements without chemical reaction.	Can be separated by a physical change (e.g. dissolving one of the elements)
There is usually an energy change when a compound is made from its elements.	No energy change when the elements are mixed.

Separation Techniques

1. SOLID & LIQUID

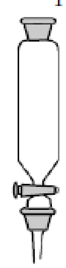
(a) If the solid has not dissolved in the liquid (i.e. is <u>suspended</u> in the liquid), the two substances can be separated by **FILTRATION**.
The liquid **<u>filtrate</u>** passes through, the undissolved solid **<u>residue</u>** stays on the paper.

(b) If the solid has dissolved in the liquid, forming a <u>solution</u>:-
(i) If only the <u>solid</u> is required, it is obtained by **EVAPORATION.**
(ii) If the <u>liquid</u> is required, it is obtained by **DISTILLATION.** The solution is placed in a flask and heated. The liquid evasporates, and its vapour passes into a condenser, where it cools and turns back to liquid. (The solid remains behind in the flask as a residue).
Note – The thermometer bulb should be at the level of the condenser.



2. LIQUID & LIQUID

(a) **Immiscible liquids** - by using a **SEPARATING FUNNEL**
When two liquids do not mix (e.g. paraffin and water), they can be separated simply by running off the denser liquid from a separating funnel by opening the tap.

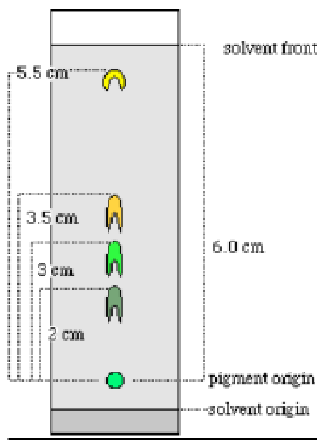


(b) **Miscible liquids** - by **FRACTIONAL DISTILLATION**
Both liquids evaporate, and their vapours pass into a fractionating column, where they are condensed and re-boiled many times. The vapour of the liquid which has the <u>lower boiling point</u>, emerges from the top of the column <u>first</u>, and passes into the condenser.
When all of this liquid has distilled, it is followed by the liquid having the higher boiling point.

3. **SOLID & SOLID** Usually one of the following methods can be used:-

(a) **DISSOLVING:** Find a solvent that will dissolve one solid but not the other.
It is often necessary to heat in order to help the process of dissolving. Then filter, wash the residue on the filter-paper with a little of the solvent, and dry it. To obtain the second solid (now in solution, in the filtrate), evaporate the filtrate, as in 1(b)(i) above.

(b) **CHROMATOGRAPHY**: This is a method for separating two solids that are both soluble in the same solvent. Find a solvent that will dissolve both solids, make a solution of the mixture and place two drops of the solution on a piece of filter-paper.
Allow a suitable liquid to spread gradually across the paper.
The solid that is more soluble in the liquid will will move through the greatest distance, and the solid that is least soluble will move least.
Thus the dissolved solids are separated on the paper.



A student was provided with only a thermometer, a stopwatch and a beaker.

What could the student measure?

A 10.5 g solid and 24.8 cm^3 liquid

B 10.5 g solid and 25 °C

C 24.8 cm^3 liquid and 45 seconds

D 25 °C and 45 seconds

A mixture of sulfur and iron filings needs to be separated. The solubilities of sulfur and iron filings in water and carbon disulfide are shown in the table below.

	solubility in water	solubility in carbon disulfide
sulfur	✗	✓
iron filings	✗	✗

What are possible methods of separating the sulfur and iron filings?

	using water	using carbon disulfide	using a magnet
A	✓	✓	✗
B	✗	✓	✓
C	✓	✗	✓
D	✗	✓	✗

Part of the instructions in an experiment reads as follows.

Quickly add 50 cm^3 of acid.

What is the best piece of apparatus to use?

A a burette

B a conical flask

C a measuring cylinder

D a pipette

Mixture 1 contains sand and water.

Mixture 2 contains salt and water.

Which method of separation could be used to obtain each of the required products from each mixture?

	mixture 1		mixture 2	
	to obtain sand	to obtain water	to obtain salt	to obtain water
A	crystallisation	distillation	filtration	filtration
B	crystallisation	filtration	filtration	distillation
C	filtration	distillation	crystallisation	filtration
D	filtration	filtration	crystallisation	distillation

A student measures the rate of two reactions.

In one reaction, there is a change in mass of the reactants during the reaction.

In the second reaction, there is a change in temperature during the reaction.

Which piece of apparatus would be essential in **both** experiments?

A balance

B clock

C pipette

D thermometer

Diagram 1 shows the paper chromatogram of substance X.

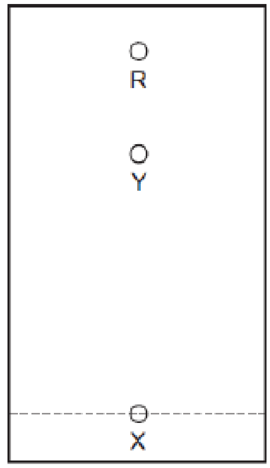


diagram 1

Diagram 2 shows the cooling curve for substance Y.

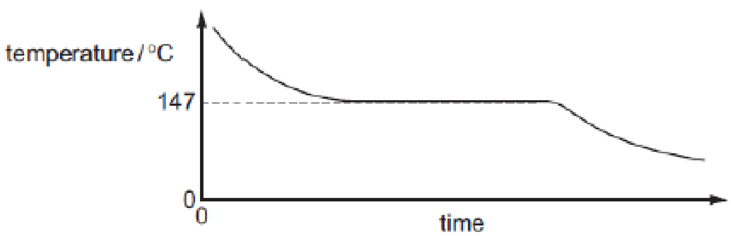


diagram 2

Which statement about X and Y is correct?

A X is a mixture and Y is a pure substance.

B X is a pure substance and Y is a mixture.

C X and Y are mixtures.

D X and Y are pure substances.

A list of techniques used to separate mixtures is given below.

filtration

diffusion

fractional distillation

simple distillation

crystallisation

chromatography

From this list, choose the most suitable technique to separate the following mixtures.
A technique may be used once, more than once or not at all.

(a) butane from a mixture of propane and butane .... [1]

(b) oxygen from liquid air .... [1]

(c) water from aqueous magnesium sulfate .... [1]

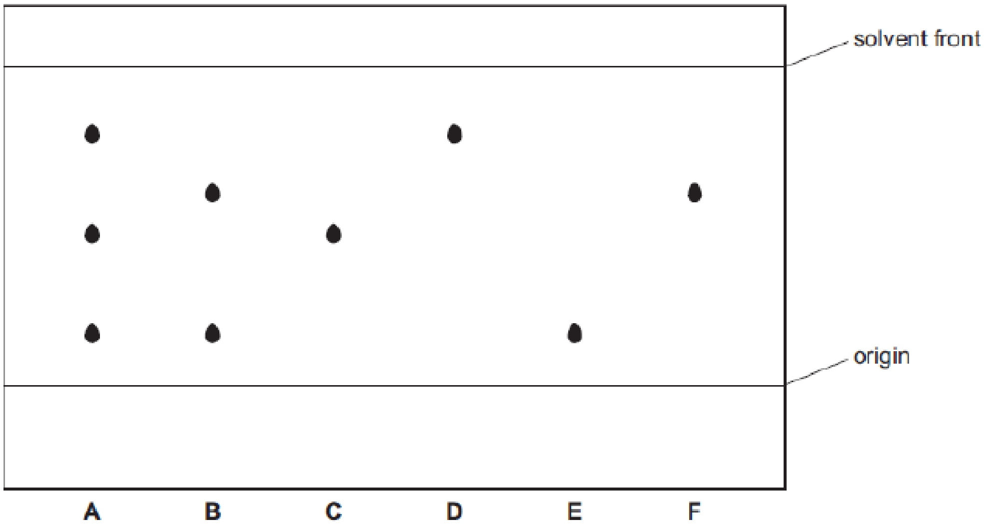
(d) potassium chloride from aqueous potassium chloride .... [1]

(e) silver chloride from a mixture of silver chloride and water .... [1]

(f) glucose from a mixture of glucose and maltose .... [1]

[Total: 6]

The diagram shows the results of an experiment to separate and identify the colours present in two coloured mixtures, **A** and **B**.
Substances **C**, **D**, **E** and **F** are single colours.



(a) Name this method of separation.

.... [1]

(b) Draw a line **on the diagram** to show the level of the solvent at the beginning of the experiment. [1]

(c) Why should a pencil be used instead of a pen to draw the origin line?

....

.... [1]

(d) State **one** difference and **one** similarity between the coloured mixtures, **A** and **B**.

difference ....

....

similarity ....

.... [2]

(e) Which substances are present in mixture **A**?

.... [1]

[Total: 6]

Seawater contains sodium chloride and other salts.
Plan an experiment to find the mass of salts in 1 dm^3 of seawater.
You will be provided with a small bottle of seawater.
You should include details of the method and any apparatus used.
(1 dm^3 = 1000 cm^3)

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....

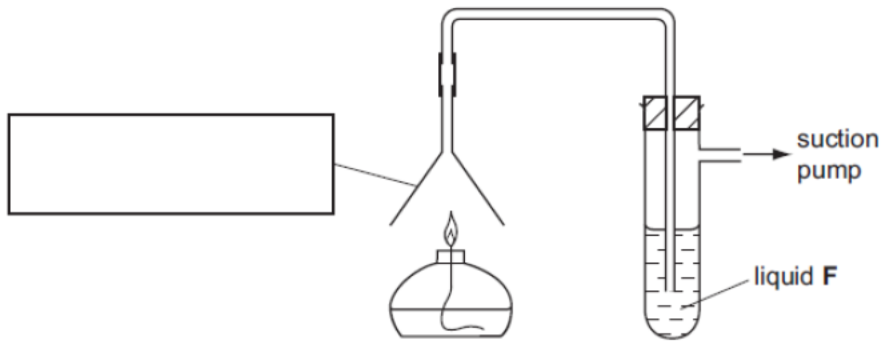
....

....

.... [6]

[Total: 6]

A student investigated the products formed when ethanol was burned using the apparatus shown.



(a) Complete the box to identify the piece of apparatus. [1]

(b) Why is a suction pump used?

.... [1]

(c) **(i)** Suggest the purpose and identity of liquid **F**.

identity ....

purpose .... [2]

(ii) Why is the end of the delivery tube below the surface of liquid **F**?

....

.... [1]

(d) Give **one** expected observation in the horizontal part of the delivery tube.
Explain your answer.

....

.... [2]

[Total: 7]

Atomic Structure

SUB TOPICS –

Atomic Structure and the Periodic Table

- State the relative charges and approximate relative masses of protons, neutrons and electrons
- Define *proton number* and *nucleon number*
- Use proton number and the simple structure of atoms to explain the basis of the Periodic Table
- Define *isotopes.*
- State the two types of isotopes as being radioactive and non-radioactive
- State one medical and one industrial use of radioactive isotopes
- Describe the build-up of electrons in 'shells' and understand the significance of the noble gas electronic structures and of valency electrons.

Bonding – The structure of matter

- Difference between elements, mixtures and compounds, and between metals and non-metals
- Describe an alloy, such as brass, as a mixture of a metal with other elements

Ions and ionic bonds

- Formation of ions by electron loss and gain
- Formation of ionic bonds between elements of Group I and VII
- Formation of ionic bonds between metallic and non-metallic elements
- Lattice structure of ionic compounds as a regular arrangement of alternating +ve and –ve ions

Molecules and covalent bonds

- Formation of single covalent bonds in H_2, Cl_2, H_2O, CH_4 and HCl
- Describe the difference in volatility, solubility and electrical conductivity between ionic and covalent compounds
- Describe the electron arrangement in more complex covalent molecules like N_2, C_2H_4, CH_3OH and CO_2

Macromolecules

- Describe the giant covalent structures of graphite and diamond
- Uses of graphite and diamond
- Structure of silicon (IV) oxide (Silicon dioxide)
- Similarity in properties between diamond and silicon (IV) oxide, related to their structures

Metallic bonding

- Describe metallic bonding as a lattice of +ve ions in a ‘sea of electrons’ and use this to describe the electrical conductivity and malleability of metals

Structure and Bonding

There are three main types of chemical bonding:

- **IONIC**
- **COVALENT**
- **METALLIC**: when a metal bonds with a metal – *a lattice of positive ions is electrostatically attracted to "a sea of" delocalized electrons.*

Ionic bonding

Ionic bonding occurs when a metal bonds with a non-metal – ***complete transfer of one or more electrons from metal to non-metal, giving charged ions that electrostatically attract.***

An IONIC BOND is defined as; the electrostatic **attraction** between oppositely charged ions.

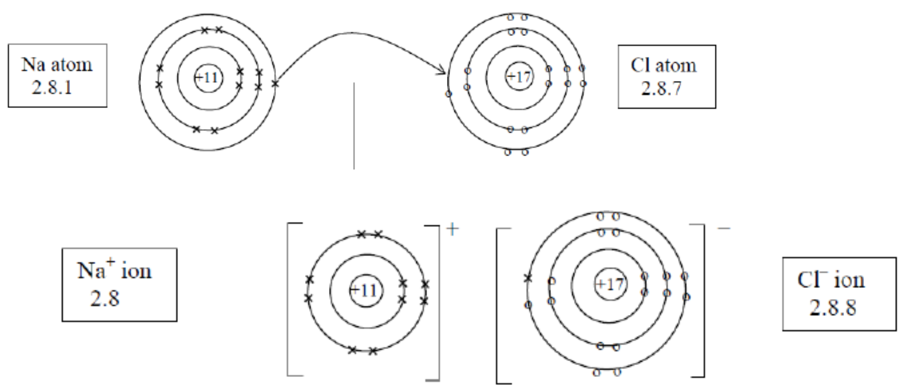
Metals in groups 1, 2 and 3 can get to a full outer shell most easily by **losing all** their outer electrons, to leave **positive ions** – **this process is called OXIDATION.**

e.g. $Na \rightarrow Na^{+} + 1e^{-}$ (2:8:1) (2.8) $\qquad$ $Mg \rightarrow Mg^{2+} + 2e^{-}$ (2.8.2) → (2.8)

Non-metals in groups 6 and 7 can get to a full outer shell by **accepting** enough electrons from a metal to make them up to 8, forming **negative ions - this process is called REDUCTION.**

e.g. $O + 2e^{-} \rightarrow O^{2-}$ (2.6) (2.8) $\qquad$ $Cl + e^{-} \rightarrow Cl^{-}$ (2.8.7) (2.8.8)

When sodium combines with chlorine, an electron is transferred completely from Na to Cl:



Covalent Bonding

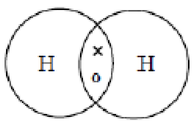
When two non-metal atoms combine they both need to gain electrons, and they can do this by **sharing** two electrons (normally one from each atom) in a **covalent bond.**

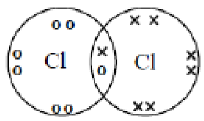
A **COVALENT BOND** is defined as; the electrostatic **attraction** between the positively charged protons in the nucleus and the negative shared pair of electrons.

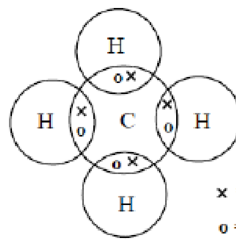
We can draw diagrams of covalent compounds between non-metal atoms by showing how the outer shells overlap, and using a dot or cross to show the electrons from the different atoms.

You need to be able to draw "dot-cross" bonding diagrams for H_2, Cl_2, NH_3, CH_4, H_2O and O_2.

In the diagrams below, notice that H atoms always have two electrons in their circles, while all the others have eight. Outer shells only are shown; a dot is used for electrons from one atom, and a cross for the other.

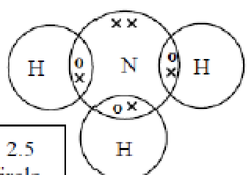




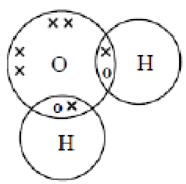


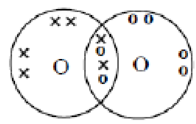
x = C electron

o = H electron



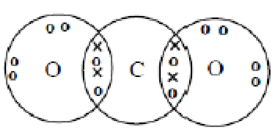
ammonia: N is 2.5 so 5 x in N's circle





O_2: O is 2.6 so forms 2 bonds
O=O is a *double* bond

In carbon dioxide, carbon (2.4) needs to form four bonds, and oxygen (2.6) needs to form two, so two *double-bonds* result (O=C=O), a linear molecule.



The covalent bond is strong, but it binds two specific atoms together (unlike the ionic attractions, which occur in all directions).
You need to know the shapes of the following molecules; CH_4 is a regular tetrahedron and CO_2 is linear.

Simple Molecular Structures

A ***molecular structure*** consists of **small** molecules, with **weak forces of attraction** (intermolecular forces) **between molecules**.
When a molecular substance is melted or boiled, it is only necessary to provide a small amount of energy to break these weak attractions, so they have low melting points and boiling points.

Molecular substances are gases, liquids, or low-melting solids at room temperature.
They usually share the following properties:

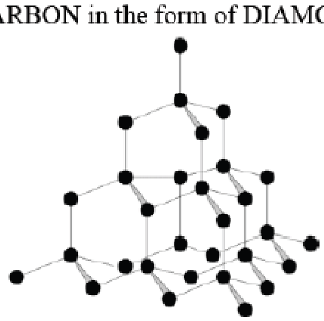
- low melting points (melting only involves breaking the weak attraction between molecules).
- low boiling points (like melting)
- don't conduct electricity in solid, or when melted, or in solution, as they have no charged particles.
- often dissolve in *non-polar solvents*, like hexane; usually insoluble in water.

As with all molecular structures these have **weak forces of attraction** between molecules so they too will have low melting points and boiling points.

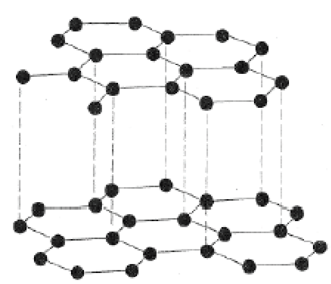
Giant Covalent structures

If a non-metal atom can form three or four bonds, it is possible for it to form giant structures linked by covalent bonds.
There are two forms of **carbon** which have giant structures.
In **diamond** each atom is covalently bonded to four neighbours, and each of those to three others, and so on throughout the whole crystal.
Graphite consists of layers of hexagons (like a honeycomb) with strong covalent bonds holding each C atom to its three neighbours.
Both diamond and graphite have very high melting points (above 4000°C) and sublimation points because it ***is*** necessary to break **all** of the **strong covalent bonds** to melt them. This requires a lot of energy.
When elements are found to exist in more than one crystalline form they are referred to as **ALLOTROPES.** Diamond and Graphite are therefore Allotropes of Carbon

CARBON in the form of DIAMOND



CARBON in the form of DIAMOND



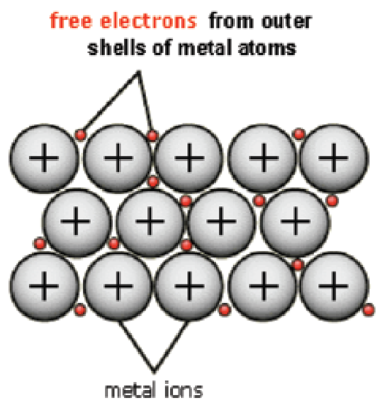
Giant covalent molecules have the following properties:

- hard (diamond is the hardest substance known)
- high melting points (some of the highest known)
- insoluble in all solvents
- don't conduct electricity in the solid, nor when molten – as they do not contain charged particles.

Metallic bonding

In metallic bonding metals give up their outer electrons to be shared with all their neighbours.
The electrons become "delocalised" in a mobile "sea" of electrons which flows between the positive ions.
The positive ions themselves pack as tightly as possible in a GAINT STRUCTURE.

A METALLIC BOND is defined as; the electrostatic **attraction** between the positively charged metal ions and the negatively charged delocalized electrons.



Giant Metallic Structures

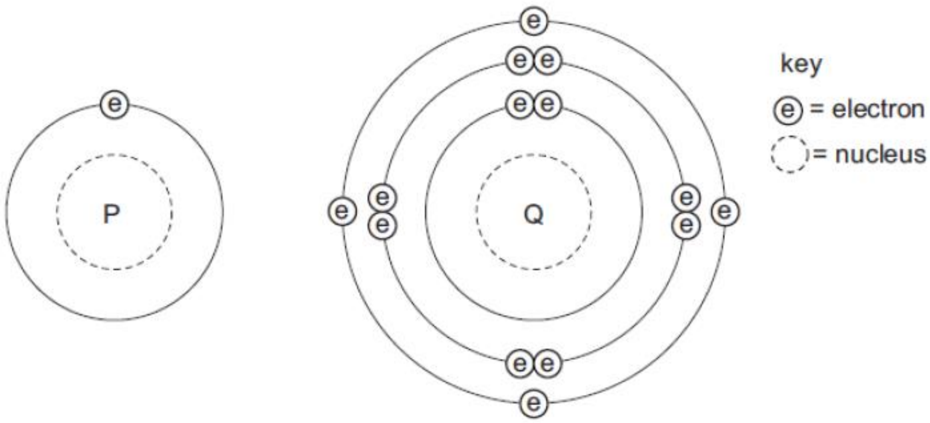
In metallic bonding metals give up their outer electrons to be shared with all their neighbours.
The electrons become "delocalised" in a mobile "sea" of electrons which flows between the positive ions.
The positive ions themselves pack as tightly as possible in a GAINT STRUCTURE.

A METALLIC BOND is defined as; the electrostatic **attraction** between the positively charged metal ions and the negatively charged delocalized electrons.

Giant metallic substances have the following properties:

- They **conduct electricity** because the electrons are free to flow between the ions and carry charge.
- They are **malleable** and **ductile** because the ions can **slide over** each other, but continue to attract each other strongly in their new positions, so that the metallic bonds do not break but distorts instead.
- They have **high melting points** because the metallic bonds are strong and require a lot of energy to break.

The diagram shows the electronic structures of atoms P and Q.



P and Q combine to form a molecule.

What is the formula of this molecule?

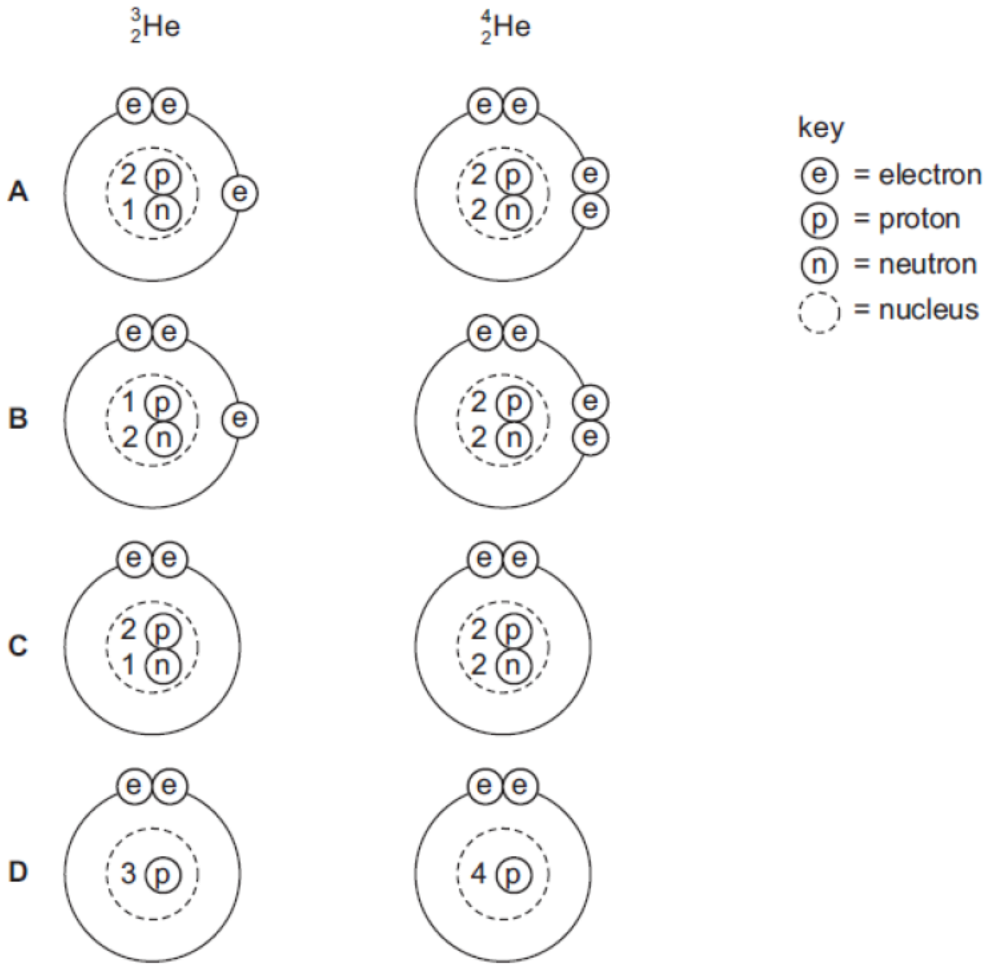
A PQ_4 **B** PQ **C** P_2Q **D** P_4Q

Which statements comparing the properties of electrons, neutrons and protons are correct?

	neutrons and protons are both heavier than electrons	only electrons and neutrons are charged
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Two isotopes of helium are ${}^{3}_{2}He$ and ${}^{4}_{2}He$.

Which two diagrams show the arrangement of particles in these two isotopes?

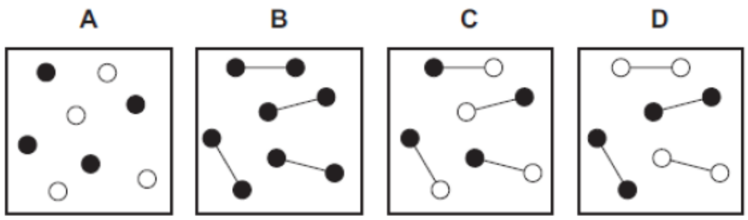


In the molecules CH_4, HCl and H_2O, which atoms use **all** of their outer shell electrons in bonding?

A C and C*l* **B** C and H **C** C*l* and H **D** H and O

Two elements, represented by ○ and ●, form a compound.

Which diagram shows molecules of the compound?



The table describes the structures of four particles.

particle	number of protons	number of neutrons	number of electrons
O	8	8	8
O^{2-}	8	8	**X**
Na	11	**Y**	11
Na^{+}	11	12	**Z**

What are the correct values of **X**, **Y** and **Z**?

	X	Y	Z
A	9	11	10
B	9	11	11
C	10	12	10
D	10	12	11

The table shows the electronic structures of four atoms.

atom	electronic structure
W	2,1
X	2,7
Y	2,8,4
Z	2,8,8

Which two atoms combine to form an ionic compound?

A W and X **B** W and Y **C** X and Y **D** X and Z

Element X has 7 protons.

Element Y has 8 more protons than X.

Which statement about element Y is correct?

A Y has more electron shells than X.

B Y has more electrons in its outer shell than X.

C Y is in a different group of the Periodic Table from X.

D Y is in the same period of the Periodic Table as X

Which row gives the number of electrons in the outer electron shell of fluorine and of neon?

	$^{19}_{9}F$	$^{20}_{10}Ne$
A	7	8
B	7	10
C	9	8
D	9	10

Lithium oxide is an ionic compound.

(i) Identify another ionic oxide in the list on page 3.

.... [1]

(ii) Draw a diagram which shows the formula of lithium oxide, the charges on the ions and the arrangement of the valency electrons around the negative ion.
Use x to represent an electron from an atom of oxygen.
Use o to represent an electron from an atom of lithium.

[2]

(b) A radioactive isotope of iodine, $^{131}_{53}\mathrm{I}$, is used to treat cancer.

(i) Define the term *isotope*.

....

.... [2]

(ii) How many protons, electrons and neutrons are there in one atom of $^{131}_{53}\mathrm{I}$?

number of protons ....

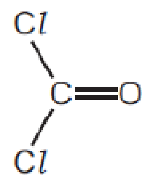
number of electrons ....

number of neutrons .... [2]

(iii) When this isotope, $^{131}_{53}\mathrm{I}$, emits radiation, a different element with a proton number of 54 is formed.
What is the name of this element?

.... [1]

The structural formula of carbonyl chloride is given below.



Draw a diagram showing the arrangement of the outer (valency) electrons in one molecule of this covalent compound.

Use o to represent an electron from a carbon atom.
Use x to represent an electron from a chlorine atom.
Use ● to represent an electron from an oxygen atom.

Silicon(IV) oxide, SiO_2, and zirconium(IV) oxide, ZrO_2, are both macromolecules.
They have similar physical properties but silicon(IV) oxide is acidic and zirconium(IV) oxide is amphoteric.

(a) Define the term *macromolecule*.

....

.... [1]

(b) **(i)** Predict **three** physical properties of these two oxides.

....

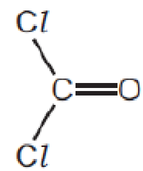
....

.... [3]

(ii) Name an element which has the same physical properties as these two oxides.

.... [1]

The structural formula of carbonyl chloride is given below.



Draw a diagram showing the arrangement of the outer (valency) electrons in one molecule of this covalent compound.

Use o to represent an electron from a carbon atom.
Use x to represent an electron from a chlorine atom.
Use ● to represent an electron from an oxygen atom.

[3]

Both strontium and sulfur have chlorides of the type XCl_2. The table below compares some of their properties.

	strontium chloride	sulfur chloride
appearance	white crystals	red liquid
formula	$SrCl_2$	SCl_2
melting point / °C	874	–120
boiling point / °C	1250	59
conductivity of liquid	good	poor
solubility in water	dissolves to form a neutral solution	reacts to form a solution of pH 1

(a) **(i)** Use the data in the table to explain why sulfur chloride is a liquid at room temperature, 25 °C.

....

.... [2]

(ii) Strontium is a metal and sulfur is a non-metal. Explain why both have chlorides of the type XCl_2.
The electron distribution of a strontium atom is 2 + 8 + 18 + 8 + 2.

....

....

.... [2]

(iii) Deduce the name of the acidic compound formed when sulfur chloride reacts with water.

.... [1]

(iv) Explain the difference in the electrical conductivity of liquid strontium chloride and liquid sulfur chloride.

....

....

.... [3]

Stoichiometry

SUB TOPICS -

- Use the symbols of the elements and write the formulae of simple compounds

- Deduce the formula of a simple compound from the relative numbers of atoms present
- Deduce the formula of a simple compound from a model or a diagrammatic representation
- Construct word equations and simple balanced chemical equations
- Define relative atomic mass, Ar
- Define relative molecular mass, Mr
- Determine the formula of an ionic compound from the charges on the ions present
- Construct equations with state symbols, including ionic equations.
- Deduce the balanced equation for a chemical reaction, given relevant information
- Define the mole and the Avogadro constant
- Use the molar gas volume, taken as 24 dm3 at room temperature and pressure
- Stoichiometric reacting masses and volumes of gases
- Concentration of solutions
- Limiting & Excess reactants
- Empirical formulae and molecular formulae
- Percentage yield and Percentage purity

Atomic and molecular masses

The relative atomic mass (Ar) of an atom is found of the periodic table.
We usually use relative atomic masses correct to the nearest whole number (or 0.5 in the case of Cl):

$H = 1$; $C = 12$; $N = 14$; $O = 16$; $Na = 23$; $Al = 27$; $S = 32$; $Cl = 35.5$; $Cu = 64$

The **relative formula mass** (RFM, Mr) of a compound is obtained by adding up the masses of all the atoms in the formula:

CO_2	r.f.m. $= 12 + 2 \times 16$	$= 44$
$Cu(NO_3)_2$	r.f.m. $= 64 + 2 \times (14 + 3\times16)$	$= 188$

Calculating percentage by mass of elements in compounds.

The percentage by mass is important to mineral prospectors, and is also useful in finding formulae.
If the formula is already known it is simply a question of finding the fraction of the formula mass made up by the element in question.

$$\% \text{ mass} = \frac{\text{Mass of the substance in the compound}}{\text{RFM of the substance}} \times 100$$

e.g. *What is the percentage by mass of aluminium in bauxite, Al_2O_3?*

RFM of Al_2O_3	$= 2\times27 + 3\times16$	$= 102$
Mass of aluminium in Al_2O_3	$= 2\times27$	$= 54$

% by mass of Al $= \frac{54}{102}\times100$ $= \underline{52.9}\%$ (3 sig. fig.)

A common problem with percentage by mass calculations occurs when an element occurs in two places in the compound.

e.g. *What is the percentage by mass of oxygen in $CuSO_4.5H_2O$?*

RFM of $CuSO_4.5H_2O$	$= 64 + 32 + 4\times16 + 5\times(2+16)$	$= 250$
Mass of oxygen in formula	$= 4\times16 + 5\times16$	$= 144$

% by mass of O $= \frac{144}{250}\times100$ $= \underline{57.6}\%$ (3 sig. fig.)

The mole and molar mass

The mole is the chemists counting unit.

One mole is the amount of substance which contains the Avogadro Constant of a specified particle (or formula).

The **Avogadro Constant** is equal to the number of atoms in 12 g of carbon–12, is about 6×10^{23}

It follows that one mole of any substance contains the same **number** of atoms (or molecules, or ions, or electrons, or other formula units).

If you work out the **relative formula mass** of a substance, the mass of 1 mole will be the same number in **g**. This is called the **molar mass**, and its units are **g/mol**.

e.g. *What is the mass of one mole of (a) CO_2 ; (b) $Cu(NO_3)_2$?*
[C=12, O=16, Cu=64, N=14]

(a) Formula mass of CO_2 $= 12 + 2 \times 16 = 44$
Molar mass of CO_2 $= 44$ g/mol

(b) Formula mass of $Cu(NO_3)_2$ $= 64 + 2 \times (14 + 3 \times 16) = 188$
Molar mass of $Cu(NO_3)_2$ $= 188$ g/mol

Amount of substance

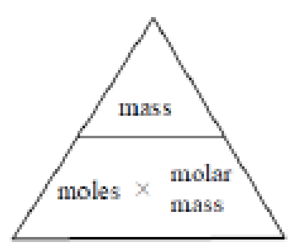
The "amount of substance" is a special name for the number of moles. If we know the mass of a substance and its molar mass, we can find the amount:

$$\textbf{Amount of substance (mol)} = \frac{\textbf{mass (g)}}{\textbf{molar mass (g/mol)}}$$

If we know the number of moles and the molar mass, we can find the mass of substance:

mass (g) = amount (mol) × molar mass (g/mol)

Some find the triangle (see right) helps them to remember: cover up one and you have the formula in terms of the other two.



Empirical and Molecular formulae

The simplest whole-number ratio of atoms in a compound is called the **empirical formula.**

We can use moles to find the formula of a substance. The steps are as follows:

(a) (Carry out an experiment to) **find the masses of all the different elements** which combine with eac other. It is the *ratio* which is important, so the total mass doesn't matter.
Sometimes data may be given as % by mass: these should be treated as the masses of the element in 10 of compound.

(b) Convert each mass to amount of substance: i.e. **divide the mass by the relative atomic mass of th element concerned**.

(c) You now have the ratio of moles of atoms of the different elements.
The ratio of atoms must be the same, since one mole of any substance contains the same number of uni
There can't be less than one atom of any element in the simplest formula, so **divide through by the number of moles which is smallest.**

(d) Then **try multiplying by small whole numbers (2, 3 or 4) to get a whole-number ratio.**

e.g. *2.88 g of magnesium is heated in nitrogen, and forms 4.00 g of magnesium nitride. Find the empirical formula of magnesium nitride. [Mg=24, N=14]*

(a) **Find the masses of all the different elements**
mass of nitrogen in sample = 4.00 – 2.88 = 1.12 g
So 2.88 g of Mg combines with 1.12 g of N

(b) **Divide the mass by the molar mass of the element concerned**
Molar masses: Mg = 24 g/mol, N = 14 g/mol
So amounts which combine are:
$\frac{2.88g}{24g/mol}$ of Mg atoms combines with $\frac{1.12g}{14g/mol}$ of N atoms
i.e. 0.120 mol of Mg atoms combines with 0.0800 mol of N atoms

(c) **Divide through by the number of moles which is smallest**
$\frac{0.120}{0.0800}$ mol of Mg combines with $\frac{0.0800}{0.0800}$ mol of N
i.e. 1.50 mol of Mg atoms combines with 1.00 mol of N atoms

(d) **Try multiplying by small whole numbers to get a whole-number ratio**
multiply by 2:
3 mol of Mg atoms combine with 2 mol of N atoms
so 3 Mg atoms combine with 2 N atoms

So simplest formula = Mg_3N_2

The molecular formula is the formula showing the actual number of each type of atom in one molecule.

For example, butene, like all alkenes has the empirical formula CH_2, but its molecular formula is C_4H_8

The molecular formula must be a whole number × the empirical formula:

We find the whole number using the RFM and the mass of the empirical formula.

$$\text{Whole number} = \frac{\text{RFM}}{\textbf{mass of empirical formula}}$$

e.g. *A compound of carbon, hydrogen and oxygen is found to be 40.0% carbon and 6.7% hydrogen by mass. Its relative molecular mass is 120.*
Find (a) its empirical formula; and (b) its molecular formula.

In 100 g of compound there is: 40.0 g of C and 6.7 g of H and (100 – 40.0 – 6.7) = 53.3 g of O

Convert to moles of atoms)	$\frac{40.0}{12}$ mol C :	$\frac{6.7}{1}$ mol H :	$\frac{53.3}{16}$ mol O
	3.33 mol C :	6.7 mol H :	3.33 mol O
Divide by smallest	1.0 mol C	2.0 mol H	1.0 mol O

(a) empirical formula = CH_2O

molecular formula = $(CH_2O)_y$

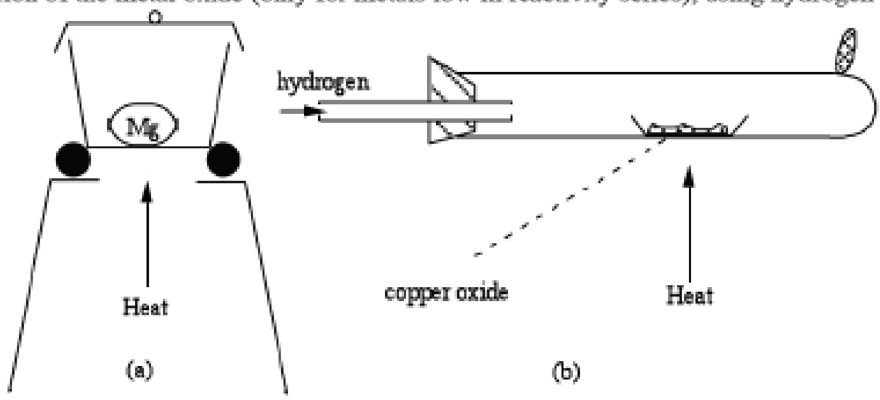
formula mass of CH_2O = 12 + 2 + 16 = 30
r.m.m. = 30 × y = 120 (given in question)
y = 4

(b) molecular formula = $C_4H_8O_4$

Finding the formulae of metal oxides

There are two main methods available for determination by experiment;
(a) by direct reaction between the element and oxygen/air.

(b) by reduction of the metal oxide (only for metals low in reactivity series), using hydrogen or methane.



Example (a): a crucible was weighed with its lid (6.20g). A coiled piece of magnesium ribbon was added, and it was weighed again (6.68g). The crucible was heated [diagram (a)] until the magnesium started to burn, then the lid was raised for brief periods until it had ceased to burn. It was heated to constant mass with the lid off. The final mass of crucible, lid and magnesium oxide was 7.00g. Find the empirical formula of the magnesium oxide. [Mg=24; O=16]

mass of magnesium used	= 6.68 – 6.20	= 0.48g
mass of oxygen gained	= 7.00 – 6.68	= 0.32g

amount of magnesium : amount of oxygen

$$= \frac{0.48\text{ g}}{24\text{ g/mol}} : \frac{0.32\text{ g}}{16\text{ g/mol}}$$

$$= 0.020\text{ mol} : 0.020\text{ mol}$$

$$= 1 : 1$$

Empirical formula = MgO

Example (b): a porcelain boat was weighed empty (3.620g). Some red copper oxide was added, and it was weighed again (4.832g). The apparatus was assembled as in the diagram (b). The hydrogen cylinder was switched on, and the air was swept out of the test tube before igniting the gas at the exit hole. The boat was heated until all the copper oxide had been reduced, and the hydrogen flow was left on while the boat cooled, to prevent air reaching the hot copper. On reweighing the boat, its mass was 4.696g. Find the empirical formula of the copper oxide. [Cu=63.5; O=16]

Water of crystallisation – Calculations

Some salts contain water molecules that form part of their crystal structure. This water is called water o crystallisation and is written in the formula of the salt.

e.g. $MgSO_4.H_2O$ refers to hydrated magnesium sulphate.
$Na_2SO_4.6H_2O$ refers to hydrated sodium sulphate.

The % water of crystallisation and the formula of the salt are calculated as follows:

- A known mass of hydrated salt is heated gently in a crucible until it reaches constant mass.
- The mass of anhydrous salt remaining and the mass of water lost are then calculated.
- These are converted to moles and the formula of the hydrated salt can be found from the **mole ratio**.

Worked example.

6.25g of blue hydrated copper(II) sulphate, $CuSO_4.xH_2O$, (x unknown) was gently heated in a crucible until the mass remaining was 4.00g. (This is the white anhydrous copper(II) sulphate $CuSO_4$).

The mass of anhydrous salt ($CuSO_4$) = 4.00g,
Mass of water (of crystallisation) driven off = 6.25 - 4.00 = 2.25g

RFM (Mr) $CuSO_4$ = 64 + 32 + (4x18) = 160 and RFM (Mr) H_2O = 1+1+16 = 18

Moles of $CuSO_4$ = 4 / 160 = 0.025
Moles of H_2O = 2.25 / 18 = 0.125

The mole ratio of $CuSO_4$: H_2O is 0.025 : 0.125 or **1 : 5**

So the formula of the hydrated salt is **$CuSO_4.5H_2O$**

Calculations from Equations: Reacting masses

Normally you will be given an equation, and asked a question which concerns only two substances, for one of which the mass is given, and for the other the mass needs to be calculated:

e.g. (6) What mass of substance A is needed to give ***x*** *g of substance B?*
What mass of substance C is produced from ***y*** *g of substance D?*

The steps involved in a calculation are as follows :

(a) Convert the information given to moles of one substance.
(b) Use the chemical equation to find moles of other substance needed.
(c) Convert back from moles to mass (or concentration, volume etc.)

e.g. *What mass of oxygen is needed to burn 3.00 kg of propane, C_3H_8?*

In this case the chemical equation is not given, so we must start by writing it down:

$$C_3H_8 + 5\ O_2 \rightarrow 3CO_2 + 4H_2O$$

(a) Convert the information given to moles of one substance.
We are given the mass of propane, so it is this we must convert to moles.

Formula mass of C_3H_8 = $3 \times 12 + 8 \times 1 = 44$ So molar mass = 44 g/mol

Number of moles of C_3H_8 in 3.00kg $= \frac{3000g}{44g/mol}$ = 68.2 mol

(b) Use the chemical equation to find moles of other substance needed.
1 mol of propane reacts with 5 mol of oxygen molecules.

So 68.2 mol of propane reacts with $5 \times 68.2 = 341$ mol of oxygen molecules

(c) Convert back from moles to mass (or concentration, volume etc.)
Molar mass of O_2 molecules = 2×16 = 32 g/mol
So mass of oxygen needed = 341 mol × 32 g/mol = 10.9 kg

e.g. *In the reaction $Fe_3O_4 + 4CO \rightarrow 3Fe + 4CO_2$ what mass of the iron oxide is needed to form 2.50 g of carbon dioxide?*

(a) molar mass of CO_2 = $12 + 2 \times 16$ = 44 g/mol
amount of CO_2 $= \frac{2.50g}{44g/mol}$ = 0.0568 mol

(b) From equation, 4 mol of CO_2 is formed from 1 mol Fe_3O_4
So 0.0568 mol of CO_2 is formed from $\frac{0.0568}{4}$ = 0.0142 mol Fe_3O_4

(c) Molar mass of Fe_3O_4 = $3 \times 56 + 4 \times 16$ = 232 g/mol
mass of 0.0142 mol Fe_3O_4 = 0.0142 mol × 232 g/mol = 3.29g (3 s.f.)

Percentage Yield

The % yield of a reaction is the percentage of the product obtained compared to the theoretical maximum as calculated from the balanced equation.

However in any chemical process it is almost impossible to get 100% of the product because;

- the reaction may not go to completion (it may be reversible)
- there may be side-reactions which use up some of the starting material
- it may not be possible to separate all of the product from the mixture

$$\textbf{\% yield} = \frac{\textbf{actual amount of desired chemical obtained}}{\textbf{maximum theoretical amount formed}} \times 100$$

Worked Example

2.8g of iron was heated with excess sulphur to form iron sulphide. **Fe + S → FeS**

The excess sulphur was dissolved in a solvent and the **FeS** filtered, washed and dried.

4.1g of purified iron sulphide was finally obtained, what was the % yield?

- Calculate the maximum theoretical amount.
  - 2.8g of iron = 2.8/RAM = 2.8/56 moles = 0.05
  - **Fe + S → FeS** 0.05 moles of Fe theoretically makes 0.05 moles of FeS
  - Theoretical amount of FeS = 0.05 x RFM = 0.05 x 88 = 4.4g

- $\% \text{ yield} = \dfrac{\text{actual amount of chemical obtained}}{\text{maximum theoretical amount}} \times 100$

  $= \dfrac{4.1}{4.4} \times 100 =$ **93.2%** (to 1dp)

Paper 2 Questions

1. 2011

The relative formula mass, M_r, of copper(II) sulfate, $CuSO_4$, is 160.

Which mass of sulfur is present in 160 g of copper(II) sulfate?

A 16 g **B** 32 g **C** 64 g **D** 128 g

2. 2012

A compound has the formula CH_3CO_2H.

How should the relative molecular mass, M_r, of this compound be calculated?

A $12 + 1 + 16$

B $3(12 + 1) + 2(12 + 16) + 1$

C $(4 \times 12) + (2 \times 1) + 16$

D $(2 \times 12) + (4 \times 1) + (2 \times 16)$

3. 2013

Which relative molecular mass, M_r, is **not** correct for the molecule given?

	molecule	M_r
A	ammonia, NH_3	17
B	carbon dioxide, CO_2	44
C	methane, CH_4	16
D	oxygen, O_2	16

CHAPTER 4 -

Stoichiometry

Paper 4 Questions (4)

1. 2011 Question 7 (c)

There are three possible equations for the thermal decomposition of sodium hydrogencarbonate.

$2NaHCO_3(s) \rightarrow Na_2O(s) + 2CO_2(g) + H_2O(g)$ **equation 1**

$NaHCO_3(s) \rightarrow NaOH(s) + CO_2(g)$ **equation 2**

$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g)$ **equation 3**

The following experiment was carried out to determine which one of the above is the correct equation.

A known mass of sodium hydrogencarbonate was heated for ten minutes. It was then allowed to cool and weighed.

Results
Mass of sodium hydrogencarbonate = 3.36 g
Mass of the residue = 2.12 g

Calculation
M_r for $NaHCO_3$ = 84 g; M_r for Na_2O = 62 g; M_r for NaOH = 40 g
M_r for Na_2CO_3 = 106 g

(i) Number of moles of $NaHCO_3$ used = ... [1]

(ii) If residue is Na_2O, number of moles of Na_2O = ...

If residue is NaOH, number of moles of NaOH = ...

If residue is Na_2CO_3, number of moles of Na_2CO_3 = ... [2]

(iii) Use the number of moles calculated in **(i)** and **(ii)** to decide which one of the three equations is correct. Explain your choice.

....

....

.... [2]

2. 2012 Question 2 (c)

Fluorine, the most reactive halogen, forms compounds with the other halogens. It forms two compounds with bromine.
Deduce their formulae from the following information.

compound 1
The mass of one mole of this compound is 137 g.

Its formula is .... [1]

compound 2
0.02 moles of this compound contain 0.02 moles of bromine atoms and 0.1 moles of fluorine atoms.

Its formula is .... [1]

3. 2012 Question 7 (c)

In the above experiment, 50.0 cm^3 of hydrochloric acid of concentration 2.0 mol / dm^3 was used. 6.4 g of $SrCl_2.6H_2O$ was made.
Calculate the percentage yield.

number of moles of HCl used = ...

number of moles of $SrCl_2.6H_2O$ which could be formed = ...

mass of one mole of $SrCl_2.6H_2O$ is 267 g

theoretical yield of $SrCl_2.6H_2O$ = ...g

percentage yield = ...% [4]

4. 2013 Question 4 (d)

Calculate the maximum mass of carbon dioxide given off when 20.0 g of small lumps of calcium carbonate react with 40 cm^3 of hydrochloric acid, concentration 2.0 mol / dm^3.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

number of moles of HCl used =

mass of carbon dioxide = ... g [4]

5. 2013 Question 6 (c)

Basic lead(II) carbonate has a formula of the type $xPbCO_3.yPb(OH)_2$ where x and y are whole numbers.
Determine x and y from the following information.

$$PbCO_3 \rightarrow PbO + CO_2$$

$$Pb(OH)_2 \rightarrow PbO + H_2O$$

When heated, the basic lead(II) carbonate gave 2.112 g of carbon dioxide and 0.432 g of water.

Mass of one mole of CO_2 = 44 g
Mass of one mole of H_2O = 18 g

Number of moles of CO_2 formed = ... [1]

Number of moles of H_2O formed = ... [1]

x = ... and y = ...

Formula of basic lead(II) carbonate is .... [1]

CHAPTER 5 –

Electricity and Chemistry

SUB TOPICS -

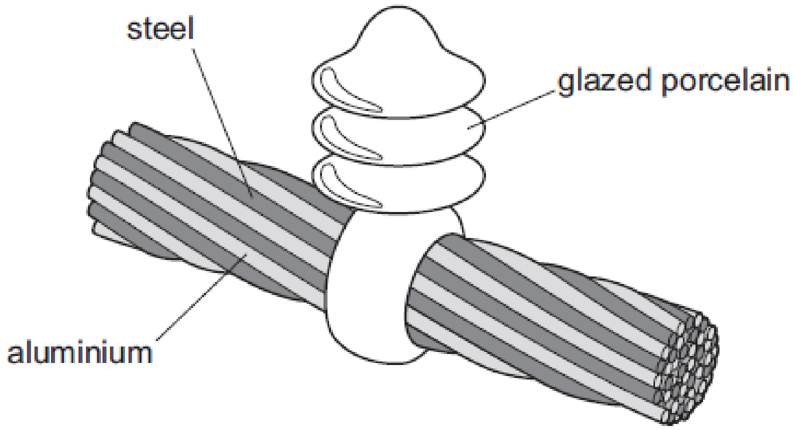
- .Describe the electrode products using inert electrodes in the electrolysis of -
  Molten lead (II) bromide
  Concentrated hydrochloric acid
  Concentrated aqueous sodium chloride
- State the general principle that metals or hydrogen are formed at the negative electrode (cathode), and that non-metals (other than hydrogen) are formed at the positive electrode (anode)
- Predict the products of the electrolysis of a specified binary compound in the molten state
- Describe the electroplating of metals

- Name the uses of electroplating
- Describe the reasons for the use of copper and (steel-cored) aluminium in cables, and why plastics and ceramics are used as insulators
- **Relate the products of electrolysis to the electrolyte and electrodes used, exemplified by the specific examples in the Core together with aqueous copper (II) sulfate using carbon electrodes and using copper electrodes.**
- **Describe electrolysis in terms of the ions present and reactions at the electrodes in the examples given.**
- **Predict the products of electrolysis of a specified halide in dilute or concentrated aqueous solution.**
- **Describe, in outline, the manufacture of**
  - **aluminium from pure aluminium oxide in molten cryolite**
  - **chlorine and sodium hydroxide from concentrated aqueous sodium chloride**

Paper 2 Questions (5)

1. 2011

The diagram shows a section of an overhead power cable.



Which statement explains why a particular substance is used?

A Aluminium has a low density and is a good conductor of electricity.

B Porcelain is a good conductor of electricity.

C Steel can rust in damp air.

D Steel is more dense than aluminium.

2. 2011

Metals could be extracted from their molten chlorides using electrolysis.

Which substances are formed at each electrode?

	anode	cathode
A	chlorine	hydrogen
B	chlorine	metal
C	hydrogen	metal
D	metal	chlorine

3. 2011

Concentrated aqueous potassium bromide solution is electrolysed using inert electrodes.

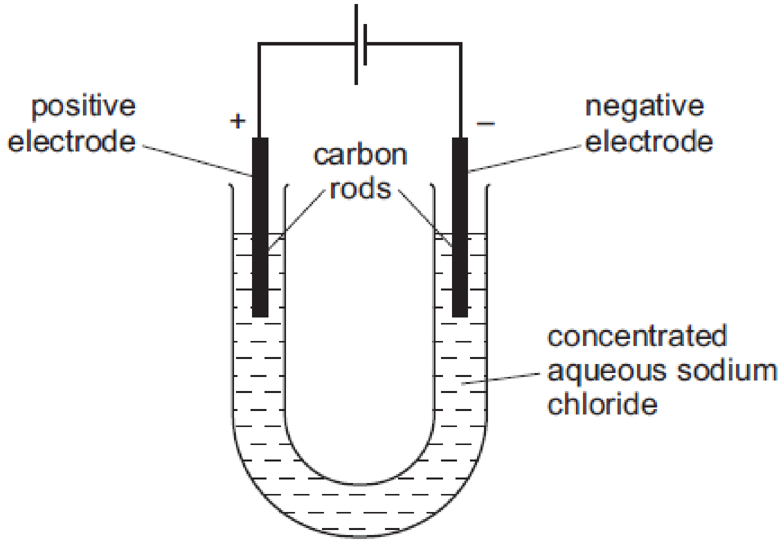
The ions present in the solution are K^+, Br^-, H^+ and OH^-.

To which electrodes are the ions attracted during this electrolysis?

	attracted to anode	attracted to cathode
A	Br^- and K^+	H^+ and OH^-
B	Br^- and OH^-	H^+ and K^+
C	H^+ and K^+	Br^- and OH^-
D	H^+ and OH^-	Br^- and K^+

4. 2012

The diagram shows the electrolysis of concentrated aqueous sodium chloride.

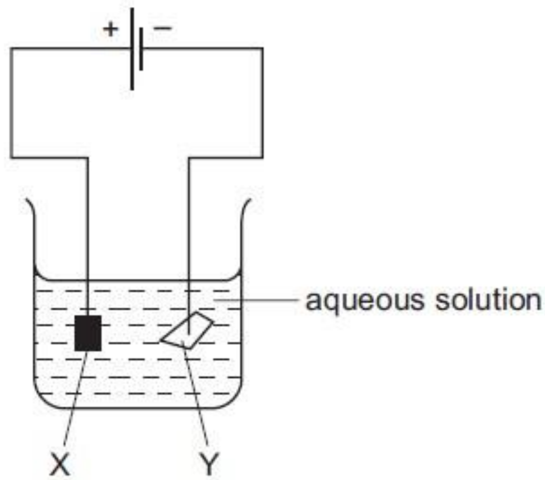


What is produced at each of the electrodes?

	product at cathode	product at anode
A	hydrogen	chlorine
B	hydrogen	oxygen
C	sodium	chlorine
D	sodium	oxygen

5. 2012

The diagram shows an electrolysis experiment using metals X and Y as electrodes.



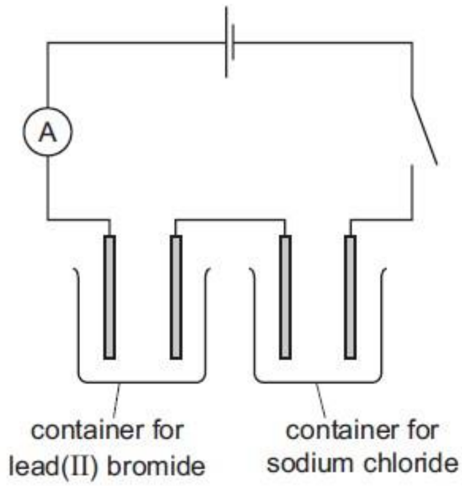
One of the metals becomes coated with copper.

Which metal becomes coated and which aqueous solution is used?

	metal	aqueous solution
A	X	$CrCl_3$
B	X	$CuCl_2$
C	Y	$CrCl_3$
D	Y	$CuCl_2$

6. 2013

The diagram shows the circuit for electrolysing lead(II) bromide and sodium chloride to liberate the metal.

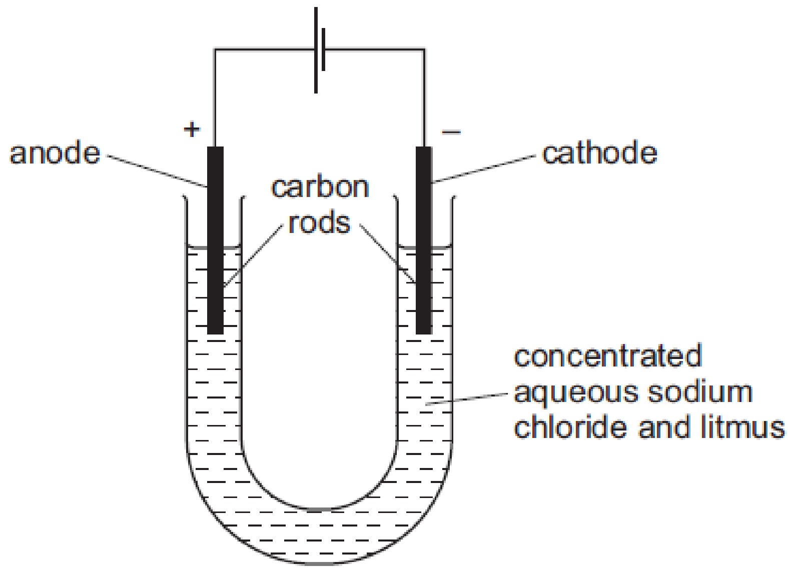


In what form are these salts electrolysed for liberating the metal?

	lead(II) bromide	sodium chloride
A	concentrated solution	concentrated solution
B	concentrated solution	molten
C	molten	concentrated solution
D	molten	molten

7. 2013

The diagram shows the electrolysis of concentrated aqueous sodium chloride.



What is the colour of the litmus at each electrode after five minutes?

	colour at anode	colour at cathode
A	blue	red
B	red	blue
C	red	colourless
D	colourless	blue

CHAPTER 5 –

Electricity and Chemistry

SUB TOPICS -

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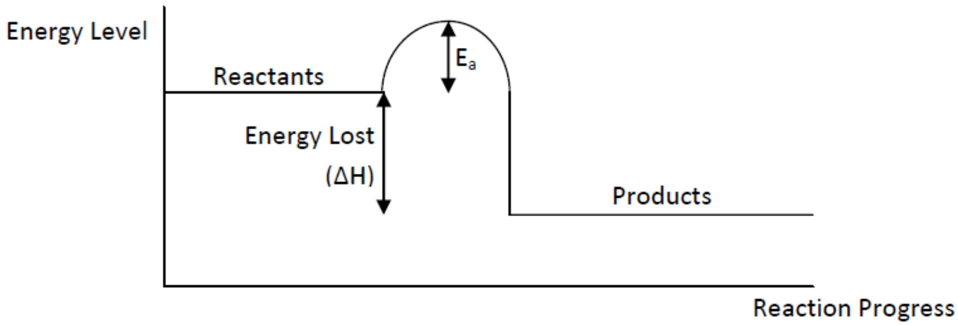
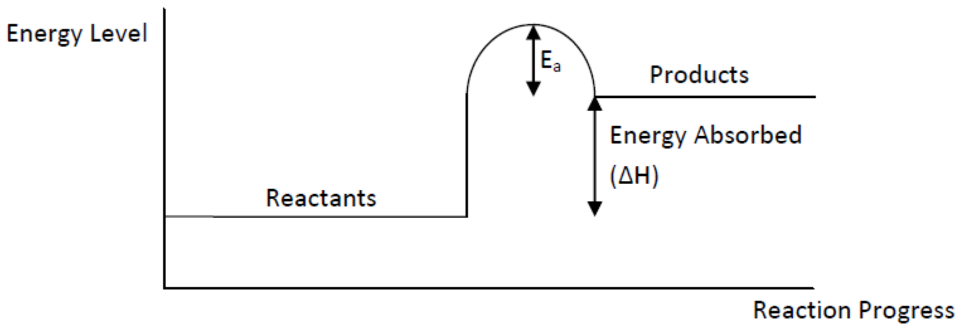
- Describe electrolysis in terms of the ions present and reactions at the electrodes in the examples given.
- Predict the products of electrolysis of a specified halide in dilute or concentrated aqueous solution.
- Describe, in outline, the manufacture of
  - aluminium from pure aluminium oxide in molten cryolite
  - chlorine and sodium hydroxide from concentrated aqueous sodium chloride

Exothermic Reactions – when energy is lost to the surroundings during a reaction.

Energy Profile for an Exothermic Reaction – the graph shows the variation in energy during the course of a chemical reaction where heat is released.



Endothermic Reactions – when energy is absorbed by the products from the surroundings during a reaction.



Activation Energy (E_a) – the initial energy that is required for a reaction to begin

Calorimeter – determines the amount of heat generated in a chemical reaction by the rise in temperature of the reaction chamber and the water jacket around the reaction vessel.

Bond Energy – the amount of energy needed or released to break or form a bond. Bond breaking is endothermic. Bond forming is exothermic.

$$\Delta H = \text{Total bond energy of all bonds broken} - \text{Total bond energy of all bonds formed}$$

Equilibrium Reactions

Irreversible Reactions – reactions that has products that cannot turn back into their reactants.

Reversible Reactions – reactions that has products that can react back into the original reactants.

Dynamic Equilibrium – when there is no overall change in the amount of products and reactants even though the reaction is ongoing. Dynamic Equilibrium can only take place in a closed system. The position of dynamic equilibrium is not always at a half-way point, as in it may be at a position where there are more products than reactants.

Le Chatelier's principle – if a closed system at equilibrium is subject to a change then the system will adjust in such a way as to minimise the effect of the change.

Factors affecting Equilibrium

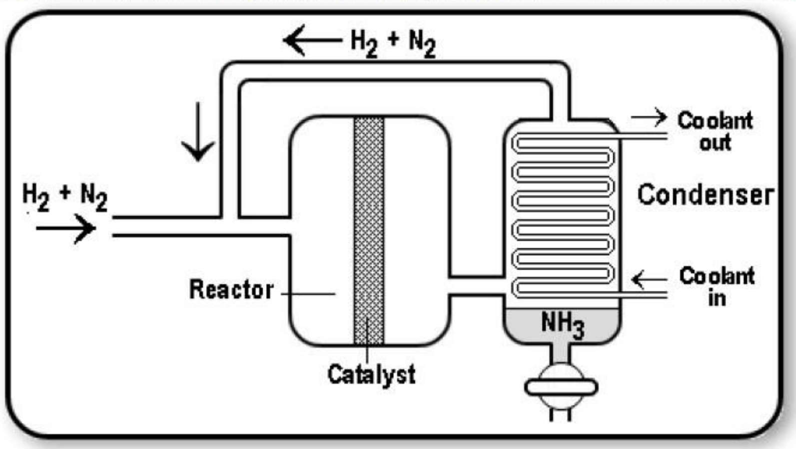
Factor	Increase of Factor	Decrease of Factor
Temperature	Equilibrium shifts to decrease the temperature so it shifts to the endothermic direction	Equilibrium shifts to increase the temperature so it shifts to the exothermic direction
Concentration	Equilibrium shifts to decrease the concentration	Equilibrium shifts to increase the concentration
Pressure	Equilibrium shifts to decrease the pressure so it shifts in the direction of the least molecules	Equilibrium shifts to increase the pressure so it shifts in the direction of the most molecules
Catalyst	Speeds up the time it takes to reach equilibrium but does not change the position	-

Haber Ammonia Process

Haber Process – the process by which ammonia is made from nitrogen and hydrogen. Nitrogen is obtained from air and hydrogen is obtained from methane. It follows the following equation:

$$N_2 + 3\ H_2 \leftrightarrow 2\ NH_3 \qquad \Delta H = -92\ kJ/mol$$

- Increasing the temperature will produce less ammonia because this will use up the added heat. Lowering the temperature will produce a greater yield of ammonia but will decrease the rate of the overall reaction.
- Increasing pressure should move the equilibrium to the right to produce more ammonia. However this will increase the cost because of the thickness of the walls of the plant needed to contain the reaction and it means the temperature will increase and its disadvantages.



Conditions of the Haber Process

- Pressure of 200 atm and Temperature between 380 and 450 °C
- Ground Iron catalyst to increase the rate of reaching equilibrium at the lower temperature
- The equilibrium mixture is cooled, allowing ammonia to liquefy and be removed.
- Unused Nitrogen and Hydrogen is continuously recycled back into the system.

Electricity and chemistry http://www.youtube.com/watch?v=AktF-vwTYo8

Using inert electrodes (platinum or carbon)

Electrolyte	Product at cathode	Product at anode
lead (II) bromide (l)	lead	bromine
concentrated HCl (aq)	Hydrogen	Chlorine
concentrated NaCl (aq)	Hydrogen	chlorine

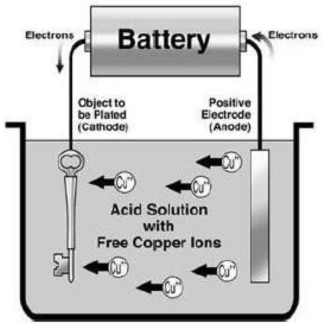
In general, metals or hydrogen are formed at the cathode (negative electrode, where reduction takes place), non-metals except hydrogen are formed at the anode (positive electrode, where oxidation takes place).

How to predict products: (FIY they will only give compounds with 2 elements like water, but no more i.e. not $CuSO_4$)

-if molten compound (e.g. iron oxide) then you just get the two components produced in elemental form (e.g. molten iron + oxygen gas)

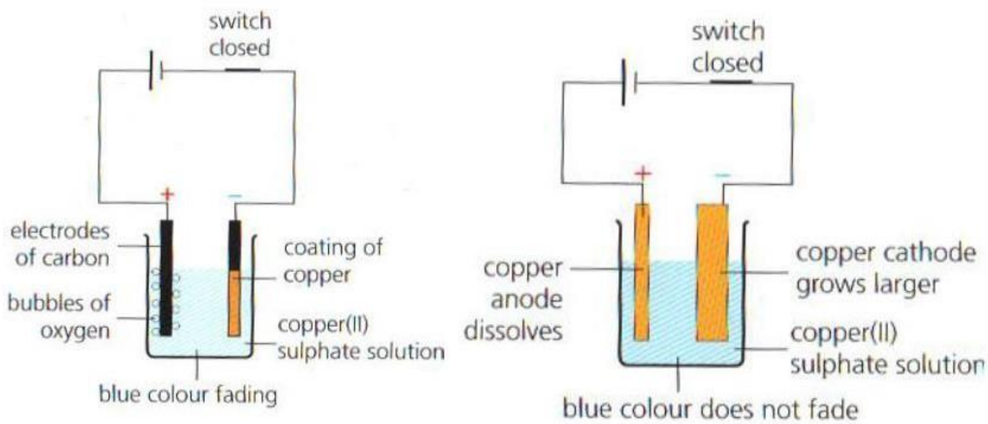
-in a solution, at the cathode, the less reactive of the two positive ions will be reduced e.g. aqueous NaCl will produce hydrogen, not sodium

-if a halide is in a concentrated solution, then a halogen will be produced at the anode. If the solution is dilute, oxygen, from the water, will be produced.



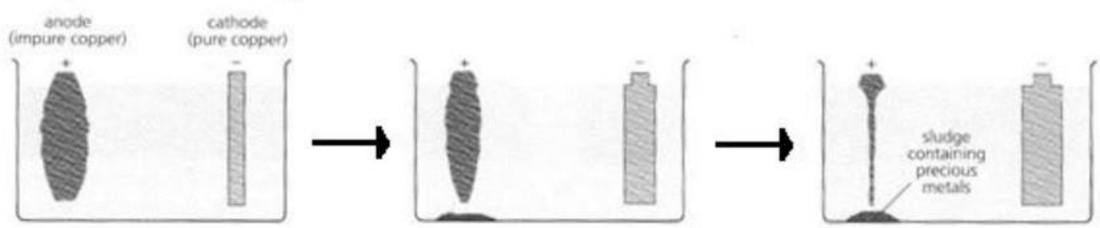
Electroplating (picture on the left)
For electroplating, you need:
-an anode made of the metal that you want to electroplate the object with,
-ions of the same metal as the anode in solution
-object to be plated at cathode
It is used to make things look better (coating a watch with gold) or to prevent corrosion.

Aluminium is used for electricity cables (not wires) because it is light, does not corrode, is a good conductor, and cheaper and much lighter than copper. The cables have a steel core, for strength. Plastics and ceramics are used as insulators because they do not conduct electricity, and conduct heat poorly. Plastic is used for the casing in electric plugs. Ceramics (inorganic, non-metallic solids) are used to support the cables in electricity pylons. Copper is used in electrical wires (not cables) as it is a very good conductor and ductile (not as good a silver, which is the best conductor used in circuitry in keyboards, phones etc., but it is cheaper)
Choice of electrodes can affect products. An example:



Electrolyte	Cathode made of:	Anode made of:	Product at cathode:	at anode:
dilute copper (II) sulphate	inert material	copper	copper	Nothing but Copper is **lost**
dilute copper (II) sulphate	inert material	inert material	copper	oxygen

This concept is used to refine copper



Electrolysis is a way to decompose compounds, using electricity. Reduction of positive cations happens at the cathode (CATions at CAThode. Oxidation of negative anions happens at the anode (ANions at ANode) for example:

At the anode: $2Cl^- \longrightarrow Cl_2 + 2e^-$

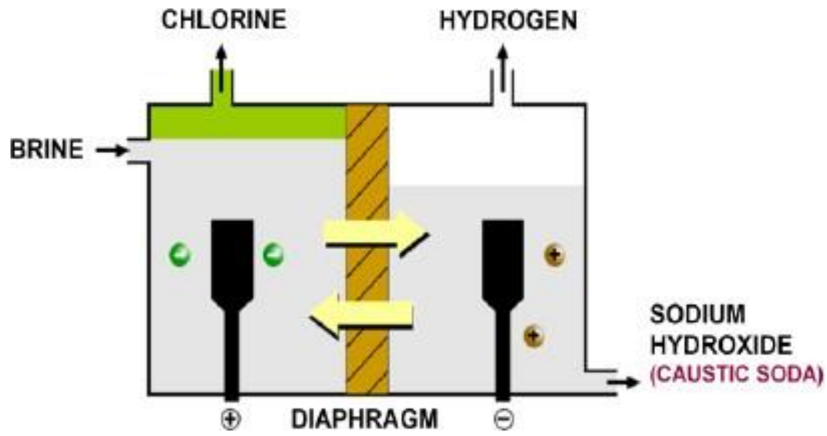
At the cathode: $2H^+ + 2e^- \longrightarrow H_2$

Electrolysis of **Brine**

Brine is concentrated NaCl solution

Chlorine (making PVC, medical drugs, pesticides, sterilising water, making chemicals) is produced at the **titanium anode**, hydrogen (making nylon, make hydrogen peroxide, hardening vegetable oils, fuel) is produced at the **steel/nickel cathode** (in the book it says steel in the text and nickel in the diagram) and sodium hydroxide (soaps, detergents, medical drugs, dyes, paper, ceramics) is left over.

Asbestos diaphragm lets ions pass through, but keeps gases apart.



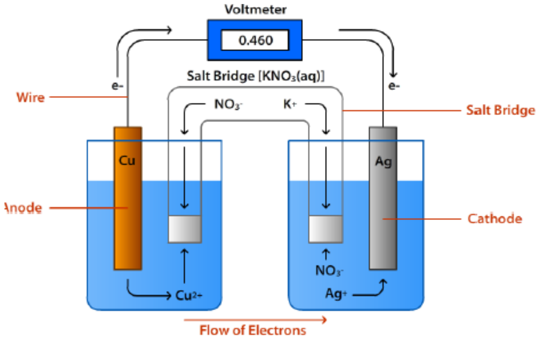
Production of energy

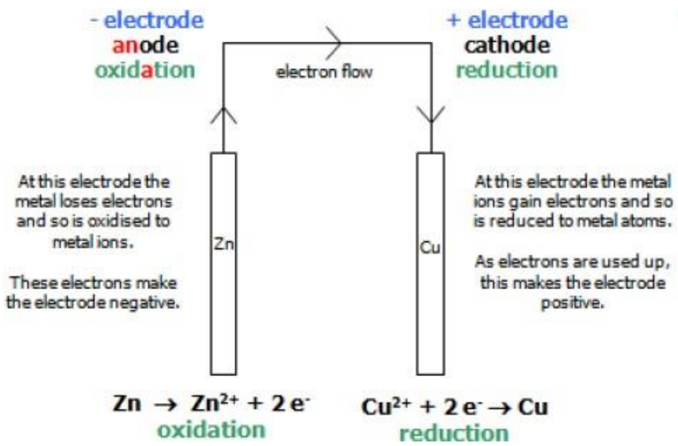
A **fuel** is a substance used to obtain energy. Burning fuels (like oil, coal etc.) to form oxides is an exothermic reaction (gives out heat). The heat from burning fuels is used in power plants to create steam from water and turn turbines.

Hydrogen: Burns explosively with oxygen, so it is used in rockets. But in a **fuel cell** it combines with oxygen without burning (details of the construction and operation of a fuel cell are **not** required).

Nuclear fuels: are not burned. The radioisotopes (unstable atoms) decay naturally giving out energy. In a nuclear power station they are forced to break down by shooting neutrons at them. They heat up water to make steam which turns turbines. Advantages: lots of energy is made from a small amount of fuel and no carbon dioxide is produced. Disadvantage is that it produces radiation and radioactive waste, non-renewable

Electrochemical cell: consists of a negative pole (the more reactive metal) and a positive pole (less reactive metal) and an electrolyte. The greater the difference in reactivity of the two metals, the greater the voltage will be. The electrons flow because one metal is more reactive, so it has a stronger drive to give up its electrons. The atoms give up electrons and enter the solution as ions.





Paper 4 Questions (5)

1. 2011 Question 3 (b)

The ions which are involved in the electrolysis are Al^{3+} and O^{2-}. The products of this electrolysis are given on the diagram.
Explain how they are formed. Use equations where appropriate.

....

....

....

.... [4]

CHAPTER 7 –

Chemical reactions

SUB TOPICS -

Rate (speed) of reaction

- Describe the effect of concentration, particle size, catalysts (including enzymes) and temperature on the rate (speed) of reactions
- Describe a practical method for investigating the rate (speed) of a reaction involving gas evolution
- Describe the application of the above factors to the danger of explosive combustion with fine powders (e.g. flour mills) and gases (e.g. mines)

Devise a suitable method for investigating the effect of a given variable on the rate (speed) of a reaction

Interpret data obtained from experiments concerned with rate (speed) of reaction

Describe and explain the effects of temperature and concentration in terms of collisions between reacting particles

Describe the role of light in photochemical reactions and the effect of light on the rate (speed) of these reactions

Describe the use of silver salts in photography as a process of reduction of silver ions to silver; and photosynthesis as the reaction between carbon dioxide and water in the presence of chlorophyll and sunlight (energy) to produce glucose and oxygen

Reversible reactions

Describe the idea that some chemical reactions can be reversed by changing the reaction conditions (Limited to the effects of heat on hydrated salts. Concept of equilibrium is not required.) including hydrated copper(II) sulfate and hydrated cobalt(II) chloride.

Predict the effect of changing the conditions (concentration, temperature and pressure) on other reversible reactions

Concept of equilibrium

Redox

Define oxidation and reduction in terms of oxygen loss/gain. (Oxidation state limited to its use to name ions, e.g. iron (II), iron(III), copper(II), dichromate (VI), manganate(VII).)

Define redox in terms of electron transfer

Identify redox reactions by changes in oxidation state and by the colour changes involved when using acidified potassium manganate(VII), and potassium iodide. (Recall of equations involving KMnO4 is not required.)

Unit 7 – Chemical reactions

Notes

Identify physical and chemical changes, and understand the differences between them

● **Physical changes**

o undergoing a physical change means no new substances are made, but

there is a change to the appearance of a substance

o e.g. being made into a mixture is a physical change since you can

separate mixtures – since they consist of elements/compounds that are

not chemically combined

▪ E.g. fractional distillation of crude oil

o also includes changing state or dissolving

● **Chemical changes**

o A process of re-arrangement of the atoms present in the reactants to

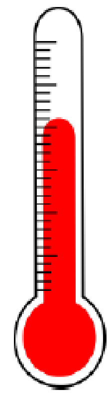
form one or more products, which have the same total number of each

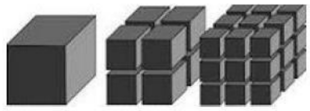
type of atoms as the reactants .

o E.g. neutralisation reactions to produce salts or rusting

Rate (speed) of reaction

Describe and explain the effect of concentration, particle size, catalysts (including enzymes) and temperature on the rate of reactions





● Increasing the temperature increases the rate of reaction. As increasing temperature increases the speed of the moving particles, so they collide more frequently and energetically.

● Increasing concentration of reacting solutions increases the rate of reaction, as it increases the frequency of collisions.

● Increasing the surface area of solid reactants (same as decreasing the particle size) increases the rate of reaction, as it increases the frequency of collisions.

● Catalysts are substances that speed up chemical reactions without being changed or used up during the reaction.

(Extended only) Devise and evaluate a suitable method for investigating the effect of a given variable on the rate of a reaction

● Use equations below to find the rate of reaction to compare the effect of changes in surface area/particle size, concentration, temperature, use of a catalyst etc…

● Rates of reactions can be measured using the amount of product used, or amount of product formed over time:

Rate of reaction = amount of reactant used/Time

Rate of reaction = amount of product formed/Time

o Quantity of reactant or product can be measured by the mass in grams or by a volume in cm^3

o if a gas is produced in a reaction, you can either collect the gas produced in a gas syringe and measure the volume over time, or let the gas escape and measure loss of mass over time

o Units of rate of reaction may be given as g/s or cm^3/s

o Use quantity of reactants in terms of moles and therefore, units for rate of reaction in mol/s

● You would want to do multiple experiments changing the variable e.g. if it was temperature do the experiment at 20˚C, 25˚C, 30˚C etc… measuring the rate at regular intervals each time to then compare (possibly graphically)

<u>*Describe the application of the above factors to the danger of explosive combustion with fine powders (e.g. flour mills) and gases (e.g. methane in mines)*</u>

● Flour mills

o Particle size is very small

o Therefore, surface area is very large

o Could easily combust causing an explosion due to these flammable substances that have a large surface area

● Methane in mines

o Increase in pressure

o Same as increasing the concentration of the reactants – because now the volume has decreased, therefore there are more particles per unit volume

o Increases chance of successful collisions

<u>*Demonstrate knowledge and understanding of a practical method for investigating the rate of a reaction involving gas evolution*</u>

● Measure the volume of a gas (if the gas is a product) using a gas syringe or an upside down measuring cylinder or burette

● Record the total volume of gas collected at regular intervals and plot a graph

● Use the rate of reaction equation above

o In the example of production of a gas, you would do: volume of gas / time taken to find the rate at the specific time

<u>*(Extended only) Describe and explain the effects of temperature and concentration in terms of collisions between reacting particles (An increase in temperature causes an increase in collision rate **and** more of the colliding molecules have sufficient energy, activation energy, to react whereas an increase in concentration only causes an increase in collision rate)*</u>

● Increasing temperature: causes an increase in collision rate **and** more of the

colliding molecules have sufficient energy, activation energy, to react, so rate of

reaction increases

● Increasing concentration: causes an increase in collision rate, increasing

frequency of successful collisions and so rate of reaction increases

Interpret data obtained from experiments concerned with rate of reaction (use the term rate over speed)

● use information above, look out for the key changes that lead to increased

reaction rate: temperature, concentration, pressure, surface area (particle size)

and use of a catalyst

(Extended only) Describe and explain the role of light in photochemical reactions and the effect of light on the rate of these reactions

● Photochemical reactions = reactions that are initiated by light

● The brighter the light/the greater the light intensity, the faster the rate of

reaction

(Extended only) Describe the use of silver salts in photography as a process of reduction of silver ions to silver; and photosynthesis as the reaction between carbon dioxide and water in the presence of chlorophyll and sunlight (energy) to produce glucose and oxygen

● Use of silver salts in photography:

o Silver halide salts are used in black and white photography

o AgCl is sensitive to light & breaks down to form metallic silver Ag + -> Ag

▪ Appears black

o Brighter the light on the film, the faster the reaction & the darker that

part of the photograph appears i.e. improves efficiency & accuracy of

photos

● Photosynthesis:

o Chemical change that occurs in the leaves of green plants

o Chlorophyll, a green pigment in the plants, absorbs light energy

▪ CO 2 reacts with H 2O to produce C6 H 12O 6 (glucose) and O 2

reaction conditions (Limited to the effects of heat and water on hydrated *and anhydrous copper(II) sulfate and cobalt(II) chloride) (Concept of equilibrium is **not** required)*

● In some chemical reactions, the products of the reaction can react to produce the original reactants

o These are called reversible reactions

o The direction of the reaction can be changed by changing the conditions

● Dehydration of hydrated copper(II) sulfate

o Anhydrous copper(II) sulfate + water ⇌ hydrated copper(II) sulfate

o White solid turns blue in presence of water

o Forward reaction add water

o Reverse reaction heat the hydrated copper(II) sulfate (water evaporates)

The above is exactly the same with hydrated and anhydrous cobalt(II) chloride

(Extended only) Predict the effect of changing the conditions (concentration, temperature and pressure) on other reversible reactions

● The relative amounts of all the reacting substances at equilibrium depend on the conditions of the reaction.

● If a system is at equilibrium and a change is made to any of the conditions, then the system responds to counteract the change.

o Effects of changing conditions on a system at equilibrium can be predicted using Le Chatelier's Principle.

● Effect of changing concentration:

● If the concentration of one of the reactants or products is changed, the system is no longer at equilibrium and the concentrations of all the substances will change until equilibrium is reached again.

● If concentration of reactants is increased : position of equilibrium shifts towards products (right) so more product is produced until equilibrium is reached again

● if concentration of products is increased : position of equilibrium shifts towards reactants (left) so more reactant is produced until equilibrium is reached again

● Effect of changing pressure:

● In gaseous reactions, an increase in pressure will favour the reaction that produces the least number of molecules as shown by the symbol equation for that reaction.

if a reaction produces a..	...larger volume of gas (more moles on product side)	...smaller volume of gas (fewer moles on product side)
an increase in pressure...	decreases yield of reaction- equilibrium shifts left	increases yield of reaction- equilibrium shifts right
a decrease in pressure...	increases yield of reaction- equilibrium shifts right	decreases yield of reaction- equilibrium shifts left

● Effect of changing temperature:

● If temperature is increased : equilibrium moves in the direction of the endothermic reaction (e.g. if forwards reaction is endothermic and temperature is increased, equilibrium shifts right to produce more product)

● If temperature is decreased : equilibrium moves in the direction of the exothermic reaction

● For the forwards being exo/endothermic and yield meaning the amount of product from the forwards reaction:

	Exothermic	Endothermic
An increase in temperature...	Decreases yield of reaction- equilibrium moves left	Increases yield of reaction- equilibrium moves right
A decrease in temperature...	Increases yield of reaction- equilibrium moves right	Decreases yield of reaction- equilibrium moves left

(Extended only) Demonstrate knowledge and understanding of the concept of equilibrium

● When a reversible reaction occurs in a closed system, equilibrium is reached when the reactions occur at exactly the same rate in each direction.

limited to its use to name ions, e.g. iron(II), iron(III), copper(II), manganate(VII))

● oxidation is gain of oxygen

● Reduction is loss of oxygen

● Roman numerals are used after an element in the name of a compound to refer to its oxidation state – normally used in reference to metals showing what + charge they have

o E.g. iron(II) is Fe 2+ , iron(III) is Fe 3+ , copper(II) is Cu 2+ , manganate(VII) is Mn 7+ etc…

(Extended only) Define redox in terms of electron transfer

● Oxidation Is Loss (of electrons)

● Reduction Is Gain (of electrons)

*(Extended only) Identify redox reactions by changes in oxidation state and by the colour changes involved when using acidified potassium manganate(VII), and potassium iodide (recall of equations involving KMnO 4 is **not** required)*

● A redox reaction is one where both oxidation and reduction take place

● If an element is gaining electrons and another is losing electrons, then the reaction is a redox reaction (or losing/gaining oxygen)

● Potassium manganate(VII):

o Deep purple and when reduced, it becomes colourless

o E.g. react with iron(II) chloride and colourless Mn 2+ ions are formed

o E.g. react with sulphur dioxide and the same thing happens

o This is because, potassium manganate(VII) is an oxidising agent and therefore is reduced itself (see below)

● Potassium iodide:

o Colourless solution is oxidised by an oxidising agent to form brown iodine

solution

o React with hydrogen peroxide (which oxidises the iodide ions to iodine,

which is brown in colour)

(Extended only) Define oxidising agent and reducing agent

● Oxidising agent is a substance which oxidises another substance during a redox

reaction

● Reducing agent is a substance which reduces another substance during a redox

reaction

(Extended only) Identify oxidising agents and reducing agents from simple equations

● Oxidising agents are reduced

● Reducing agents are oxidised

● they can be identified by recognising which elements have been

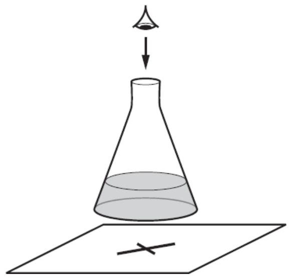
oxidised/reduced in an equation

Practice questions :

1. When aqueous sodium thiosulfate and dilute hydrochloric acid are mixed, a precipitate of insoluble sulfur is produced. This makes the mixture difficult to see through.

$$Na2S2O3(aq) + 2HCl\ (aq) \rightarrow S(s) + 2NaCl\ (aq) + H2O(l) + SO2(g)$$

A student adds the following volumes of aqueous sodium thiosulfate, dilute hydrochloric acid and distilled water to the conical flask.



The time taken for the formation of the precipitate of sulfur to make the cross disappear from view is recorded.

experiment number	volume of sodium thiosulfate / cm^3	volume of hydrochloric acid / cm^3	volume of distilled water / cm^3	time taken for cross to disappear from view / s
1	10	10	40	56
2	20	10	30	28
3				

(a) State the order in which the aqueous sodium thiosulfate, hydrochloric acid and distilled water should be added to the flask.

(b) In experiment 3 the student wanted the sodium thiosulfate to be double the concentration used in experiment 2.

(i) Complete the table to show the **volumes** which should be used and the **expected** time taken for the cross to disappear from view in experiment 3.

(ii) Use collision theory to explain why increasing the concentration of sodium thiosulfate would change the rate of reaction.

(c) The student repeated experiment 1 at a higher temperature.

Use collision theory to explain why the rate of reaction would increase.

2. Hydrogen can be manufactured from methane by steam reforming.

CH4(g) + H2O(g) CO(g) + 3H2(g)

The reaction is carried out using a nickel catalyst at temperatures between 700 °C and 1100 °C and using a pressure of one atmosphere.

The forward reaction is endothermic.

(a) What is meant by the term *catalyst*?

(b) Suggest **two** reasons why a temperature lower than 700 °C is not used.

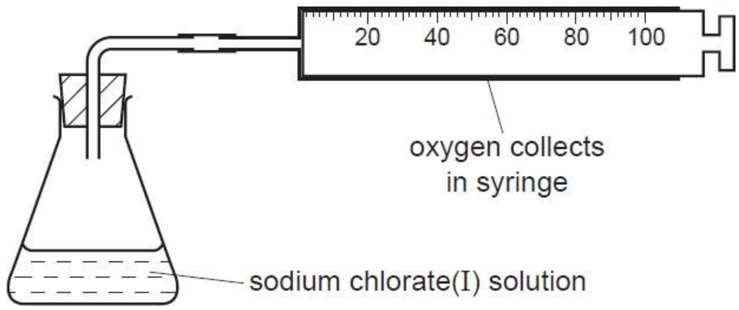
(c) Suggest **one** advantage of using a pressure greater than one atmosphere.

(d) Suggest **one** disadvantage of using a pressure greater than one atmosphere.

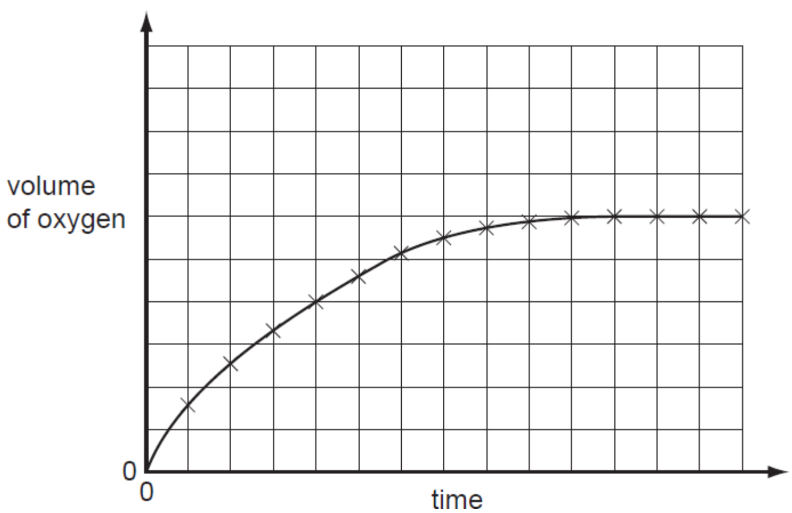
3. **(a)** Sodium chlorate(I) decomposes to form sodium chloride and oxygen. The rate of this reaction is very slow at room temperature provided the sodium chlorate(I) is stored in a dark bottle to prevent exposure to light.

$$2NaClO \rightarrow 2NaCl + O2$$

The rate of this decomposition can be studied using the following experiment.



Sodium chlorate(I) is placed in the fl ask and 0.2 g of copper(II) oxide is added. This catalyses the decomposition of the sodium chlorate(I) and the volume of oxygen collected is measured every minute. The results are plotted to give a graph of the type shown below.



(i) Explain why the gradient (slope) of this graph decreases with time.

(ii) Cobalt(II) oxide is a more effi cient catalyst for this reaction than copper(II) oxide.

Sketch, on the grid, the graph for the reaction catalysed by cobalt(II) oxide. All other conditions were kept constant.

(iii) What can you deduce from the comment that sodium chlorate(I) has to be shielded from light?

(iv) Explain, in terms of collisions between particles, why the initial gradient would be steeper if the experiment was repeated at a higher temperature.

4. Iodine reacts with chlorine to form dark brown iodine monochloride.

$$I2 + Cl2 \rightarrow 2ICl$$

This reacts with more chlorine to give yellow iodine trichloride.

An equilibrium forms between these iodine chlorides.

$$ICl(l) + Cl2(g) \rightleftharpoons ICl3(s)$$

dark brown yellow

(a) What do you understand by the term equilibrium?

(b) When the equilibrium mixture is heated, it becomes a darker brown colour.

Suggest if the reverse reaction is endothermic or exothermic. Give a reason for your choice.

(c) The pressure on the equilibrium mixture is decreased.

(i) How would this affect the position of equilibrium? Give a reason for your choice.

It would move to the

reason

(ii) Describe what you would observe.

(d) Calculate the overall energy change for the reaction between iodine and chlorine using the bond energy values shown.

$$I_2 + Cl_2 \rightarrow 2ICl$$

Bond	Energy / kJ per mol
I–I	151
Cl–Cl	242
I–Cl	208

Show your working.

(e) Draw a labelled energy level diagram for the reaction between iodine and chlorine using the information in (d).

5. The following reactivity series shows both familiar and unfamiliar elements in order of decreasing reactivity. Each element is represented by a redox equation

$$Rb \rightleftharpoons Rb^{+} + e^{-}$$
$$Mg \rightleftharpoons Mg^{2+} + 2e^{-}$$
$$Mn \rightleftharpoons Mn^{2+} + 2e^{-}$$
$$Zn \rightleftharpoons Zn^{2+} + 2e^{-}$$
$$H_2 \rightleftharpoons 2H^{+} + 2e^{-}$$
$$Cu \rightleftharpoons Cu^{2+} + 2e^{-}$$
$$Hg \rightleftharpoons Hg^{2+} + 2e^{-}$$

Two of the uses of the series are to predict the thermal stability of compounds of the metals and to explain their redox reactions.

(a) Most metal hydroxides decompose when heated.

(i) Complete the equation for the thermal decomposition of copper(II) hydroxide.

Cu(OH)2 → .... + ...

(ii) Choose a metal from the above series whose hydroxide does not decompose when heated

(b) Define in terms of electron transfer the term *oxidation*.

(ii) Explain why the positive ions in the above equations are oxidising agents.

(c) Which metals in the series above do not react with dilute acids to form hydrogen?

(ii) Describe an experiment which would confi rm the prediction made in **(c)(i)**.

(ii) Describe what you would observe when zinc, a reducing agent, is added to this pink / purple solution.

Unit 8 – Acid Base

Notes

Describe the characteristic properties of acids as reactions with metals, bases, carbonates and effect on litmus and methyl orange

acid + metal ⟶ salt + hydrogen

acid + base ⟶ salt + water

acid + metal ⟶ salt + water + carbon dioxide

● All three of these above reactions are neutralisation reactions

● The salt produced depends on the acid used:

▪ Hydrochloric acid produces chlorides

▪ Nitric acid produces nitrates

▪ Sulfuric acid produces sulfates

o It also depends on the positive ions in the base, alkali or carbonate i.e. the metal

● Red litmus (for the above reactions would just stay red- assuming that you start with the acid and add the metal)

o Stays red in acidic

o Stays red in neutral

o Turns blue in alkaline

● Blue litmus (for the above reactions would just stay blue- assuming that you start

with the base then add the acid)

o Turns red in acidic

o Stays blue in neutral

o Stays blue in alkaline

● Methyl orange (for the above reactions would change from red to yellow)

o Red in acidic

o Yellow in neutral

o Yellow in alkaline

(Extended only) Define acids and bases in terms of proton transfer, limited to aqueous solutions

● Protons are H^+ ions

● Acids are proton donors and bases are proton acceptors, therefore there is a proton transfer from acids to bases

Describe the characteristic properties of bases as reactions with acids and with ammonium salts and effect on litmus and methyl orange

● Acid + ammonia -> ammonium salt

o E.g. $HCl + NH_3 \rightarrow NH_4Cl$

● Effect would be going from alkaline to neutral (assuming that you start with the base or ammonia then add the acid)

o Methyl orange: stays yellow

o Red litmus: blue to red

o Blue litmus: stays blue

(Extended only) Describe the meaning of weak and strong acids and bases

● Acids release H^+ ions in aqueous solution

● Strong acid = completely dissociates to release H^+ ions in aqueous solution

o Hydrochloric, nitric and sulfuric acids

● Weak acid = partially dissociates to release H^+ ions in aqueous solution

o Ethanoic, citric and carbonic acids

● Stronger an acid, lower the pH (for a given conc. of aq. solutions)

● As the pH decreases by one unit, the H^+ conc. of the solution increases by a factor of 10.

● bases release OH^- ions in aqueous solution

● strong bases fully dissociate to release OH^- ions in aqueous solution

● weak bases partially dissociate to release OH^-

ions in aqueous solution

● the stronger a base, the higher the pH

● Strong and weak is NOT the same as concentrated and dilute – the latter refers to the amount of substance whereas, the former refers to the above – the H^+ ion conc. in aq. solutions

Describe neutrality and relative acidity and alkalinity in terms of pH measured using Universal Indicator paper (whole numbers only)

● Acids produce H^+ ions in aqueous solutions

● Alkalis produce OH^- ions in aqueous solutions

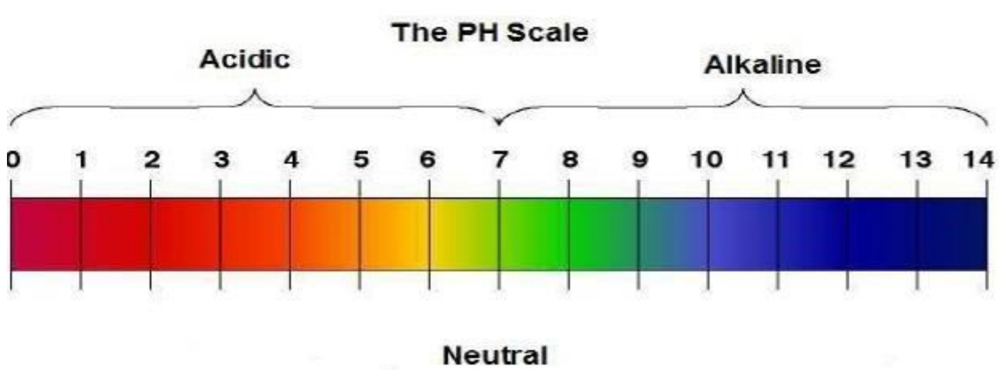
● The pH scale (0 to 14) measures the acidity or alkalinity of a solution, and can be measured using universal indicator of a pH probe

o pH 7 is neutral

o < pH 7 is acidic

o > pH 7 is alkaline

● $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ is the ionic equation of any neutralisation reaction



Describe and explain the importance of controlling acidity in soil

● If the pH of soil is too low i.e. too acidic, this would mean that crops would be unable to grow in these acidic soils

● Farmers use lime (calcium oxide) to neutralise acid soils

Classify oxides as either acidic or basic, related to metallic and non-metallic character

● Many metals and non-metals react with oxygen in the air when they are heated to produce metal oxides and non-metal oxides

● The nature of these oxides is either acidic or basic

o E.g. MgO basic (metal oxide)

o E.g. CO_2 acidic (non-metal oxide)

o E.g. SO_2 acidic (non-metal oxide)

(Extended only) Further classify other oxides as neutral or amphoteric

● Neutral oxides also exist

o These are non-metal oxides that neither neutralise acids or base

o E.g. H_2O water, CO, NO etc…

● Amphoteric oxides exist as well

o Metal-oxides that neutralise both acid and bases to form a salt and water (amphoteric means having both acidic and basic properties)

o E.g. Al_2O_3 and ZnO

Demonstrate knowledge and understanding of preparation, separation and purification of salts

salts from insoluble bases:

● add excess base to the acid

● filter to remove any unreacted base that’s been added

● heat solution or leave in warm place to evaporate the water, allowing crystals of the salt to form

salts from soluble bases:

● use an acid-base titration to find the exact volume of the soluble base that reacts with the acid

● mix these correct proportions, which will produce a solution of the salt and water only (as all the acid and base has reacted)

● warm solution to evaporate the water to leave crystals of the salt

(Extended only) Demonstrating knowledge and understanding of the preparation of insoluble salts by precipitation

An insoluble salt is formed as a precipitate, which means it is a solid in the water/solution.

● making insoluble salts:

○ mix the two solutions that will form the salt

○ filter to remove the solid insoluble salt

○ wash salt with distilled water

○ leave salt to dry on filter paper (or dry more quickly in an oven)

(Extended only) Suggest a method of making a given salt from a suitable starting material, given appropriate information

● If reactants are soluble = titration

● If base is insoluble = add an excess of it to acid to ensure neutralisation and then remove the excess

● if you are given the name of a salt and need to identify suitable starting materials:

o first part of name e.g. sodium → use a base containing sodium e.g. sodium hydroxide

o second part of name → acid:

▪ sulfate →sulfuric acid

▪ nitrate → nitric acid

▪ chloride → hydrochloric acid

Describe the following tests to identify:

aqueous cations: aluminium, ammonium, calcium, chromium(III), copper(II), iron(II), iron(III) and zinc (using aqueous sodium hydroxide and aqueous ammonia as appropriate) (Formulae of complex ions are ***not*** *required)*

in dilute NaOH

ammonium	ammonia gas given off (how to test is further below)	no change
calcium	White precipitate forms	No change
chromium (III)	grey-green precipitate forms	dark green solution forms
copper (II)	blue precipitate forms	no change
iron (II)	green precipitate forms	no change
iron (III)	orange-brown precipitate forms	no change
zinc	White precipitate forms	Precipitate re dissolves

in ammonia (not needed to identify the other ions)

ions	few drops ammonia	excess ammonia
aluminium	White precipitate forms	No change
calcium	No precipitate formed	No precipitate formed
zinc	White precipitate forms	Precipitate re dissolves

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cations: use of the flame test to identify lithium, sodium, potassium and copper(II)

Lithium	Crimson
Sodium	Yellow
Potassium	Lilac
Copper(II)	Green

anions: carbonate (by reaction with dilute acid and then limewater), chloride, bromide and iodide (by reaction under acidic conditions with aqueous silver nitrate), nitrate (by reduction with aluminium), sulfate (by reaction under acidic conditions with aqueous barium ions) and sulfite (by reaction with dilute acids and then aqueous potassium manganate(VII))

● Carbonates

o Carbonates react with dilute acids to create carbon dioxide.

o This gas can be bubbled through limewater, if the limewater goes cloudy, the gas is CO_2

● Halides

o First add dilute nitric acid, followed by silver nitrate solution

o Chloride gives a white precipitate

o Bromide gives a cream precipitate

o Iodine gives a yellow precipitate

● Nitrates

o React copper/aluminium/zinc alloy with nitrate in sodium hydroxide solution then add aluminium and ammonia is released

o Turns damp red litmus paper blue

● Sulfates

o First add dilute hydrochloric acid, followed by barium chloride solution

o A white precipitate will form when sulfate ions are in this solution

● Sulfites

o React with dilute acids

▪ Sulfur dioxide gas is given off

▪ Use aqueous potassium manganate(VII) and there will be a colour change from purple to colourless

gases: ammonia (using damp red litmus paper), carbon dioxide (using limewater), chlorine (using damp litmus paper), hydrogen (using lighted splint), oxygen (using a glowing splint), and sulfur dioxide (using aqueous

● Ammonia

o Turns damp red litmus paper blue

● Carbon dioxide

o bubble the gas through the limewater (calcium hydroxide) and it will turn milky (cloudy)

● Chlorine

o When damp litmus paper is put into chlorine gas the litmus paper is bleached and turns white

● Hydrogen

o Use a burning splint held at the open end of a test tube of the gas

▪ Creates a ‘squeaky pop’ sound

● Oxygen

o Uses a glowing splint inserted into a test tube of the gas

▪ Splint relights in oxygen

● Sulfur dioxide

o Use aqueous potassium manganate(VII) and there will be a colour change

from purple to colourless

Practice questions :

1. What is the correct sequence of steps for the preparation of a pure sample of copper(II) sulfate

crystals from copper(II) oxide and sulfuric acid?

A dissolving → crystallisation → evaporation → filtration

B dissolving → evaporation → filtration → crystallisation

C dissolving → filtration → crystallisation → evaporation

D dissolving → filtration → evaporation → crystallisation

2. Salts can be made by adding different substances to dilute hydrochloric acid.

For which substance could any excess not be removed by filtration?

A copper(II) oxide

B magnesium

C sodium hydroxide

D zinc hydroxide

3. A salt is produced in each of the following reactions.

P magnesium + dilute hydrochloric acid

Q zinc oxide + dilute sulfuric acid

R sodium hydroxide + dilute hydrochloric acid

S copper carbonate + dilute sulfuric acid

Which statements about the products of the reactions are correct?

1.A flammable gas is produced in reaction P.

2. Water is formed in all reactions.

3. All the salts formed are soluble in water.

A 1, 2 and 3   B 1 and 2 only   C 1 and 3 only   D 2 and 3 only

4. Zinc sulfate is a soluble salt and can be prepared by reacting excess zinc carbonate with dilute

sulfuric acid.

Which piece of equipment would not be required in the preparation of zinc sulfate crystals?

A beaker

B condenser

C evaporating dish

D filter funnel

5. Which of the following methods are suitable for preparing both zinc sulfate and copper sulfate?

1. Reacting the metal oxide with warm dilute aqueous sulfuric acid.

2. Reacting the metal with dilute aqueous sulfuric acid.

3. Reacting the metal carbonate with dilute aqueous sulfuric acid.

A 1 and 2 only

B 1 and 3 only

C 2 and 3 only

D 1, 2

6. Which two processes are involved in the preparation of magnesium sulfate from dilute sulfuric

acid and an excess of magnesium oxide?

A neutralisation and filtration

B neutralisation and oxidation

C thermal decomposition and filtration

D thermal decomposition and oxidation

7. How many different salts could be made from a supply of dilute sulfuric acid, dilute hydrochloric

acid, copper, magnesium oxide and zinc carbonate?

A 3 B 4 C 5 D 6

8. Which salt preparation uses a burette and a pipette?

A calcium nitrate from calcium carbonate and nitric acid

B copper(II) sulfate from copper(II) hydroxide and sulfuric acid

C potassium chloride from potassium hydroxide and hydrochloric acid

D zinc chloride from zinc and hydrochloric acid

9. Which acid reacts with ammonia to produce the salt ammonium sulfate?

A hydrochloric B nitric C phosphoric D sulfuric

10. Anhydrous copper(II) sulfate can be made by heating hydrated copper(II) sulfate.

$$CuSO4.5H2O \rightarrow CuSO4 + 5H2O$$

What can be added to anhydrous copper(II) sulfate to turn it into hydrated copper(II) sulfate?

A concentrated sulfuric acid

B sodium hydroxide powder

C sulfur dioxide

D water

11. Three ways of making salts are

● titration using a soluble base or carbonate

● neutralisation using an insoluble base or carbonate

● precipitation.

(a) Complete the following table of salt preparations.

method	reagent 1	reagent 2	salt
titration	....	....	sodium nitrate
neutralisation	nitric acid	....	copper(II) nitrate
precipitation	....	....	silver(I) chloride
neutralisation	sulfuric acid	zinc(II) carbonate	....

(b) Write an ionic equation with state symbols for the preparation of silver(I) chloride.

(ii) Complete the following equation.

$ZnCO_3 + H_2SO_4 \rightarrow$ .... + ... + ...

Energy changes in reactions

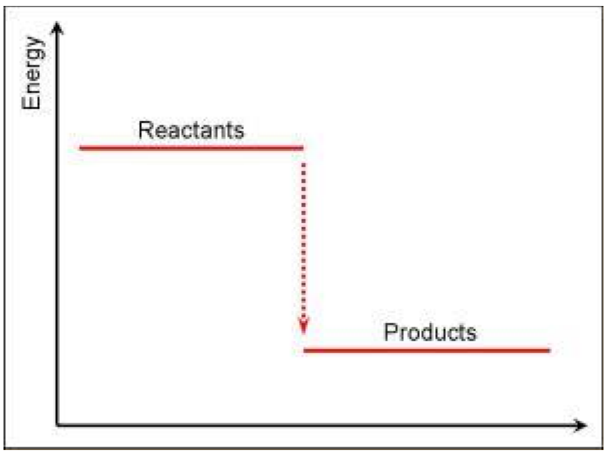
An energy change takes place during a chemical reaction - in most cases energy is given out or taken in in the form of heat. Depending on the energy exchanged, reactions can be divided into **exothermic** (*exo* = outside, *therm* = heat; give heat to the environment, release energy) and **endothermic** (*endo* = inside, *therm* = heat; take in heat from the environment, absorb energy).

Exothermic reactions

Neutralisation reactions, respiration in living organisms and combustions release heat. They are hence exothermic reactions. An exothermic reaction can be summarised as follows:

reactants --> products + energy

The energy is given out to the environment, hence the products have a lower energy than the reactants. The energy change can be shown on an **energy level diagram**:



The energy change can be measured. Since we talk about energy, the measurement unit is the joule (J) or kilojoule (kJ):

FeS + S --> FeS energy change = -100kJ/mole

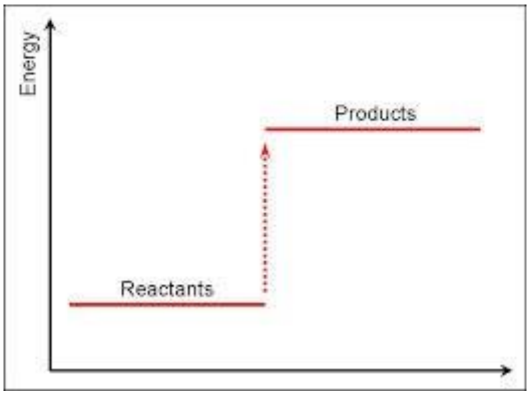
This basically means that when 1 mole of iron reacts with 1 mole of sulphur 100kJ of energy is released. The **minus** sign shoes that energy is **given out**. Thus, the temperature increases.

Endothermic reactions

Thermal decompositions, sherbet (citric acid and sodium hydrogencarbonate) or ammonium salts reacting with strong alkalis are examples of endothermic reactions. Photosynthesis falls in this category as well. They absorb heat from the environment, thus decreasing the temperature. They can be summarized as follows:

reactants + energy --> products

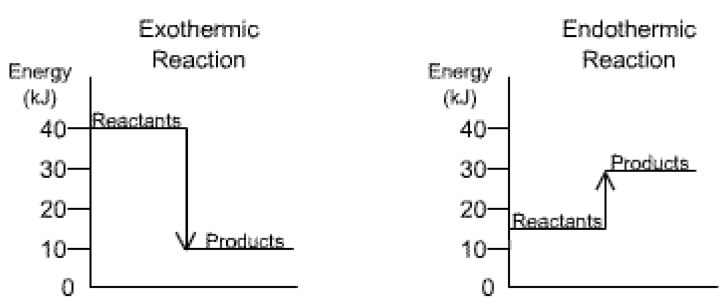
The energy is absorbed from the environment be it the air, our tongue or the Bunsen burner. Since energy is taken in, the energy level of the products is greater than the one of the reactants:



An example of an endothermic reaction:

$CaCO_3$ --> CaO + CO_2 energy change = +178 kJ/mole

178 kJ of energy is needed to decompose 1 mole of calcium carbonate. The **plus** sign shows that energy is **taken in**.



Comparison between an exothermic reaction and an endothermic reaction in terms of energy change.

Explaining energy changes

Whenever a chemical reaction takes place, bonds are broken in order to form new bonds. Basically, you need energy in order to break the bonds from the reactants (so this is an endothermic process). When the new bonds are formed in the products, they stabilize the compounds, so energy is released (so formation of bonds is an exothermic process) Depending on which one of these two energy values is greater, the overall process is endothermic or exothermic:

energy change = energy in - energy out

where energy in stands for the energy needed to break the bonds and energy out is the energy released from the formation of new bonds.

In an exothermic reaction the energy released in bond making is greater than the energy used for bond breaking.

In an endothermic reaction, the energy released in bond making is less than the energy needed for bond breaking.

The energy needed to break a bond or released when the bond is formed is called **bond energy.** You can find tables with the bond energies for the most common bonds. The bond energy is given in **kJ/mole:**

Average Bond Energies (kJ/mol)

Single Bonds						Multiple Bonds	
H—H	432	N—H	391	I—I	149	C=C	614
H—F	565	N—N	160	I—Cl	208	C≡C	839
H—Cl	427	N—F	272	I—Br	175	O=O	495
H—Br	363	N—Cl	200			C=O*	745
H—I	295	N—Br	243	S—H	347	C≡O	1072
		N—O	201	S—F	327	N=O	607
C—H	413	O—H	467	S—Cl	253	N=N	418
C—C	347	O—O	146	S—Br	218	N≡N	941
C—N	305	O—F	190	S—S	266	C≡N	891
C—O	358	O—Cl	203			C=N	615
C—F	485	O—I	234	Si—Si	340		
C—Cl	339			Si—H	393		
C—Br	276	F—F	154	Si—C	360		
C—I	240	F—Cl	253	Si—O	452		
C—S	259	F—Br	237				
		Cl—Cl	239				
		Cl—Br	218				
		Br—Br	193				

*C=O(CO_2) = 799

Using the values found in tables you can calculate the overall energy change:

Calculating energy changes (enthalpy)

The energy in is also called the **activation energy** - this is the energy needed to trigger the reaction (or to break the bonds from the reactants).

Reversible reactions

A reversible reaction is a reaction in which the products can be converted back into the reactants

- under certain conditions. In order to show that a reaction is reversible, instead of using the arrow we are using a double half arrow (one pointing to the right, one to the left). The two reactions are called the **forward**reaction (towards the right) and the **backward** reaction (towards the left):

A reaction that is not reversible will have the usual arrow pointing towards the right:

Here are some examples of reversible reactions:

a. the thermal decomposition of ammonium chloride

Thermal decomposition of ammonium chloride

When ammonium chloride is heated, it decomposes into ammonia and hydrogen chloride, following the reaction:

word equation: ammonium chloride + heat --> ammonia + hydrogen chloride
formula equation: $NH_4Cl(s)$ **+ heat -->** $NH_3(g)$ **+** $HCl(g)$

NH3

As ammonia and hydrogen chloride start to diffuse in the test tube, they get farther away from the source of heat (the bottom of the test tube is hotter since it's kept in the burner flame). At lower temperatures they react again to form solid ammonium chloride (the white solid formed on the walls of the test tube):

word equation: ammonia + hydrogen chloride --> ammonium chloride + heat

formula equation: **NH**$_3$(g) **+ HCl**(g) **--> NH**$_4$**Cl**(s) **+ heat**

The amount of heat consumed in the decomposition reaction is equal with the amount of energy released from the second reaction. In other words, for a reversible reaction the energy change is equal for the two reactions, but it has opposite signs. If one reaction is exothermic (energy having a negative value), the opposite reaction is endothermic (energy has a positive value).

Equilibrium

If a reversible reaction is carried out in a closed system/container, a point is reached in which the concentration of he substances does not change any more. This state is called an **equilibrium.** This does not mean that the reactions stopped, it just means that the forward reaction and the backward reaction take place at the same rate. There seems to be no overall change, when in reality both reactions carry on. For this reason we refer to this state as a **dynamic equilibrium**.

The amount of reactants and products at equilibrium depend on more factors, like temperature, pressure, concentrations at the start of the reaction. The equilibrium can also be shifted towards one side or the other using **Le Chatelier's principle**:

If we act with a constraint upon a system in an equilibrium state, the system will try to diminish that constraint. In other words, the system will oppose the change, shifting the equilibrium so that the change is diminished.

CHEMICAL REACTIONS (paper 6 components)

FACTORS THAT AFFECT REACTION RATES

1. Nature of Reactants
2. Surface Area of Solid Reactants
3. Concentration of Reactants
4. Temperature of Reaction System
5. Presence of a Catalyst

1. Nature of Reactants

Some generalizations:

a) State of Matter

(i) Gases tend to react faster than aqueous solutions, liquids or solids

b)

Reactions involving ionic species tend to proceed faster than reactions involving molecular compounds (covalent bonds are stronger)

c)

- Reactions involving the breaking of weaker bonds proceed faster than reactions involving the breaking of stronger bonds.
- Reactions involving the breaking of fewer bonds per reactant proceed faster than those involving the breaking of a larger number of bonds per reactant.

d)

- Strong acids which are completely ionized in solution react faster than weak acids which are only partially ionized in solution

e) **Metals on the top of reactivity series react faster than those below.**

*

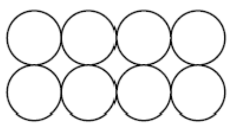
2. Surface Area of Solid Reactants

- An increase in the exposed matter of a solid reactant increases the reaction rate

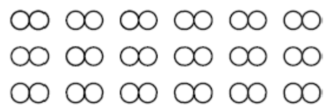
- Explained by the *Collision Theory*
  - The greater the surface area the better the chance of collision. The more collisions per unit of time, the faster the reaction.

  - That's why many solids are powdered using a mortar and pestle before being used in a reaction.

- Examples of reactions where surface area is important are:
  - active metals with acids, e.g. HCl with zinc
  - coal dust with oxygen gas
  - grain dust with oxygen gas



A lump of solid (1.0g)



Powdered solid (1.0g)

3. Concentration of Reactants

- An increase in the concentration in one or more of the reactants will increase the reaction rate (as long as the reactant is in the rate-determining step).

- Explained by the *Collision Theory*
  - An increase in the number of particles in a given volume of space (mol/L) will result in a better chance of collision. The more collisions per unit of time, the faster the reaction.

eg. $2\ HCl_{(aq)} + Zn_{(s)} \rightarrow H_{2(g)} + ZnCl_{2(aq)}$

  - 2.00 g of Zinc will react faster in 6.00 mol/L HCl than the same mass of Zinc in 2.00 mol/L HCl

4. Temperature of the Reaction System

- In general, as temperature increases, reaction rate increases.

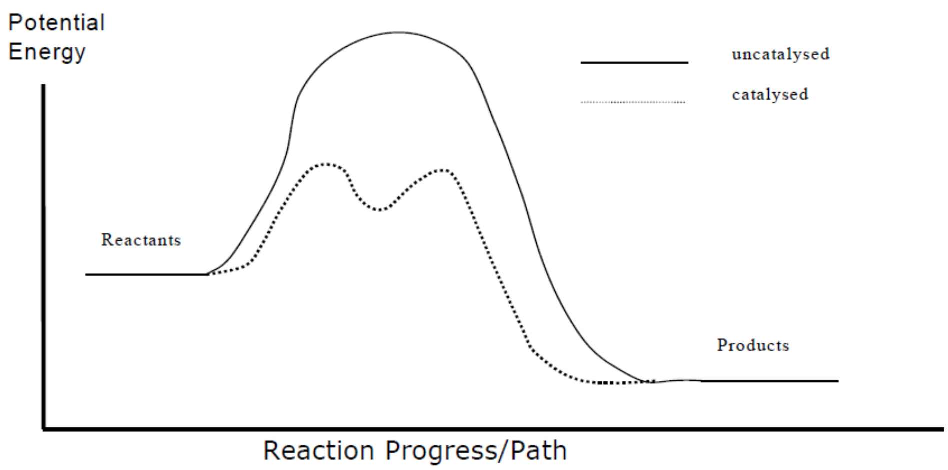
- Explained by the *Kinetic Molecular Theory and the Collision Theory*
  - At higher temperatures, particles have greater kinetic energy so they are moving faster

5. Catalyst

- A catalyst is a substance that speeds up the reaction rate without being used up in the process. (Therefore it can be regenerated and used again)

- It provides a new pathway (reaction mechanism) to form products, one that requires less energy
  - It lowers the activation energy for the reaction, therefore more particles have the required energy to overcome the activation energy barrier and form products

- The effect of adding a catalyst on a reaction can be demonstrated on a Potential Energy Diagram:



6. The effect of Pressure

- **If one or more of the reactants is a gas** then **increasing pressure** will effectively increase the concentration of the reactant molecules and **speed up the reaction**.

7. The effect of Stirring

- If the reacting mixture is not stirred 'evenly', the reactant tends to settle out at the bottom of the flask and delay the reaction .

8. LIGHT:

Some chemical reactions need light energy. They speed up if the intensity of light is more.

$2AgBr$…...........> $2Ag + Br_2$

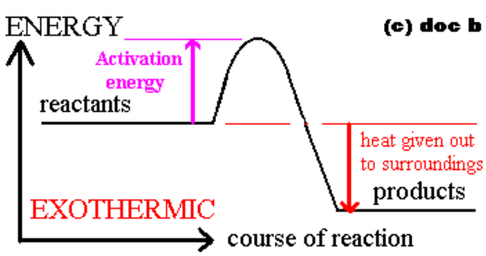
9.**Enzymes**:

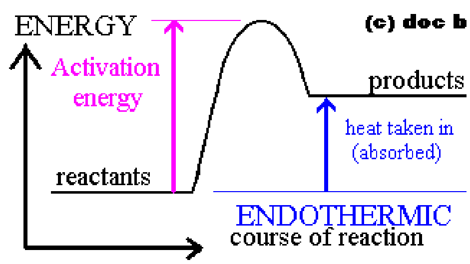
They are biocatalyst and they work well between 25 to 45 degree celcius temperature.

Raw liver

$2H_2O_2$ .... $>2H_2O + O_2$

EXOTHERMIC AND ENDOTHERMIC REACTIONS





EXOTHERMIC & ENDOTHERMIC REACTIONS

1. **Bonds are broken in reactants. Bond breaking needs/absorbs energy.**
2. **Bonds are formed in products. Bond making releases energy.**
3. **If** bond breaking needs more energy **than bond making then the reaction is endothermic.**
4. **If** bond making releases more energy **than bond breaking then the reaction is exothermic.**
5. **Fe + S….>FeS Heat energy= - 100 KJ/Mol**

   The minus sign shows energy is given out.(Exothermic)

   CaCO3….>CaO + CO2 Heat energy= +178 KJ/Mol.

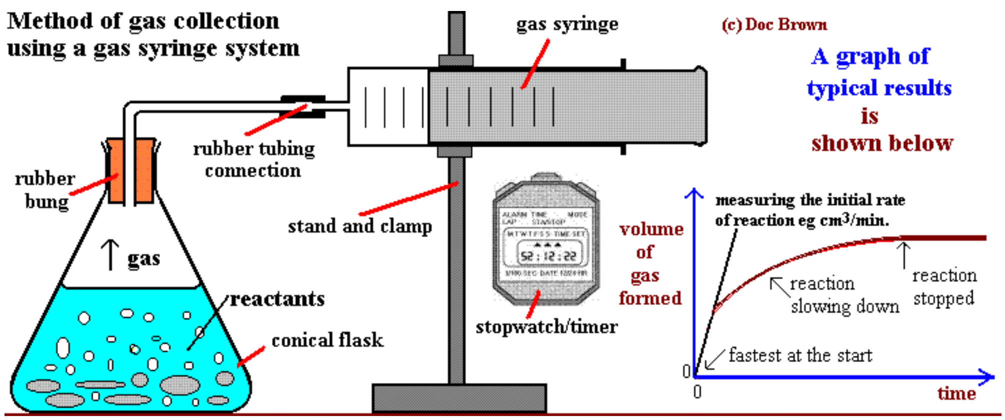
The plus sign shows energy is taken in (Endothermic)

The importance of "Rates of Reaction knowledge":

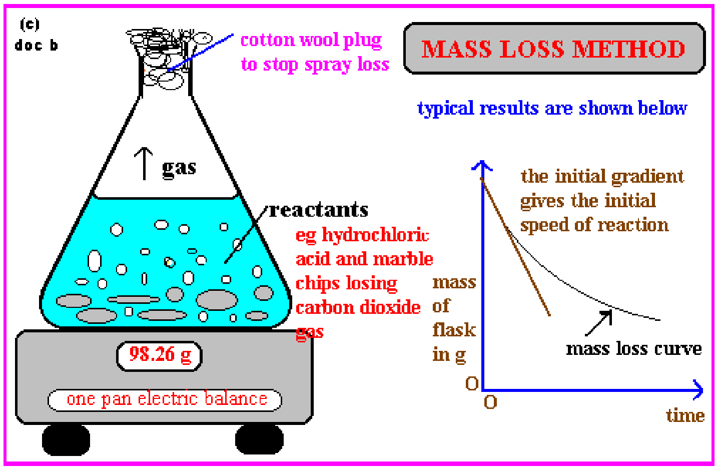
- **To save time .Time is money in industry**, the faster the reaction can be done, the more economic it is.
- **For - Health and Safety Issues.**

To measure the 'speed' or 'rate' of a reaction

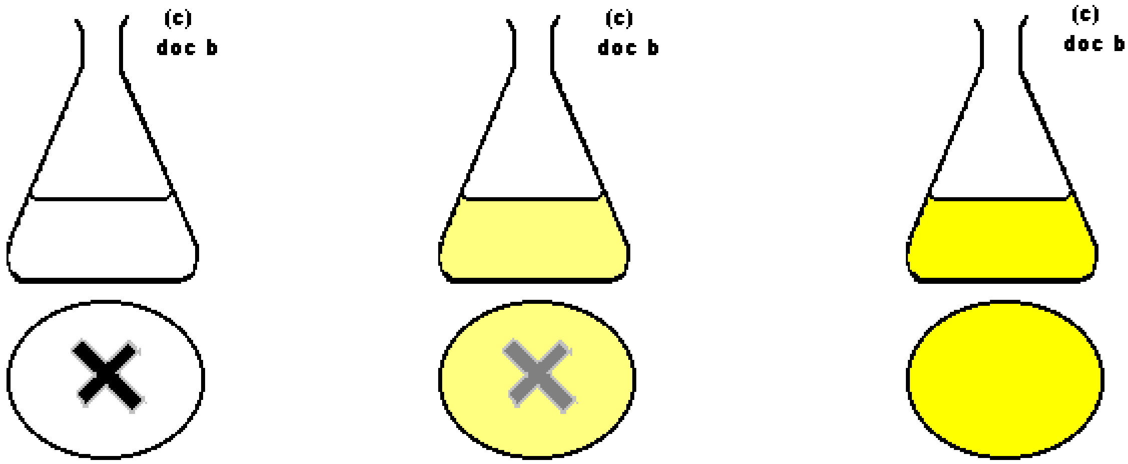
1. **When a gas is formed** from a solid reacting with a solution, it can be collected in a **gas syringe** and a graph can be plotted to measure the rate of the reaction.



- The **rate of a reaction that produces a gas can also be measured by following the mass loss as the gas is formed and escapes** from the reaction flask.
- The method is ok for reactions producing carbon dioxide or oxygen, but not very accurate for reactions giving hydrogen (too low a mass loss for accuracy).
- The reaction rate is expressed as the rate of loss in mass from the flask in e.g. g/min


(c)
doc b
cotton wool plug
to stop spray loss
MASS LOSS METHOD
typical results are shown below
gas
reactants
eg hydrochloric
acid and marble
chips losing
carbon dioxide
gas
98.26 g
one pan electric balance
the initial gradient
gives the initial
speed of reaction
mass
of
flask
in g
mass loss curve
O
O
time

3. When **sodium thiosulphate reacts with an acid, a yellow precipitate of sulphur is formed**



the reciprocal of the time is taken as a measure of the relative rate of reaction.

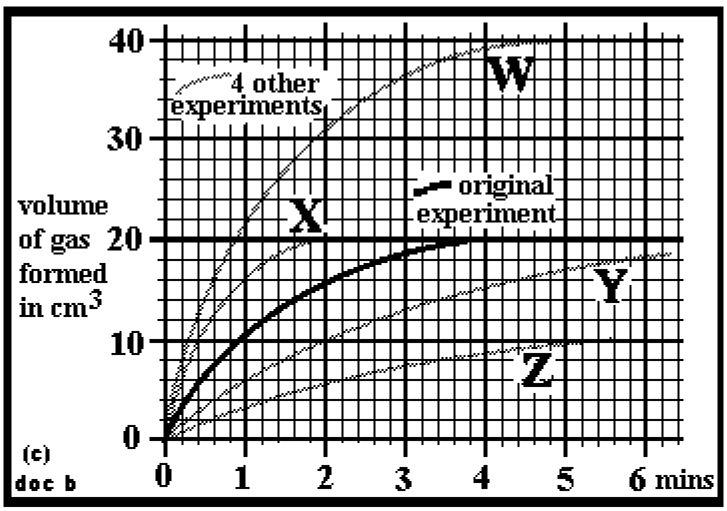
- GRAPHS IN RATE OF REACTIONS

	Graph 1 shows the decrease in the amount of a solid reactant with time. The graph is curved, becoming less steep as the gradient decreases because the reactants are being used up, so the speed decreases.
	Graph 2 shows the increase in the amount of a solid product with time. The graph tends towards a maximum amount possible when all the solid reactant is used up and the graph becomes horizontal. This means the speed has become zero as the reaction has stopped.
	Graph 3 shows the decrease in reaction time with increase in temperature as the reaction speeds up. The reaction time can represent how long it takes to form a fixed amount of gas in e.g. in the first few minutes of a metal/carbonate - acid reaction
	Graph 4 shows the increase in speed of a reaction with increase in temperature as the particles have more and more kinetic energy.

Graph 5 shows the increase in the amount of a gas formed in a reaction with time. Here the gradient is a measure of the rate of the reaction.

Graph 6 shows the effect of increasing concentration, which decreases the reaction time**, as the speed increases because the greater the concentration the greater the chance of fruitful collision.**

Graph 7 shows the rate/speed of reaction is often proportional to the concentration of one particular reactant.



Graph 4.8 A set of results for the same reaction

(i) The **graph lines W, X, original, Y and Z on the left diagram are typical of when a gaseous product is being collected**. The middle graph might represent the original experiment 'recipe' and temperature. Then the experiment repeated with variations e.g.

(ii) X could be the same recipe as the original experiment but a **catalyst added**, forming the same amount of product, but faster.

(iii) Initially, the increasing order of **rate of reaction represented on the graph by curves Z to W i.e. W > X > original > Y > Z** might represent**progressively increasing concentrations of reactant** or **progressively higher temperature of reaction** or **progressively smaller lumps-particle/increasing surface area of a solid reactant**. All three trends in changing a reactant/reaction condition variable **produce a progressively faster reaction shown by the increasing gradient in cm^3/min which represents the rate/speed of the reaction**.

(iv) Z could represent taking **half the amount of reactants or half a concentration**. The reaction is slower and only half as much gas is formed.

(v) W might represent taking **double the quantity of reactants**, forming twice as much gas e.g. same volume of reactant solution but doubling the concentration, so producing twice as much gas, initially at double the speed (gradient twice as steep).

-

1.

practice questions

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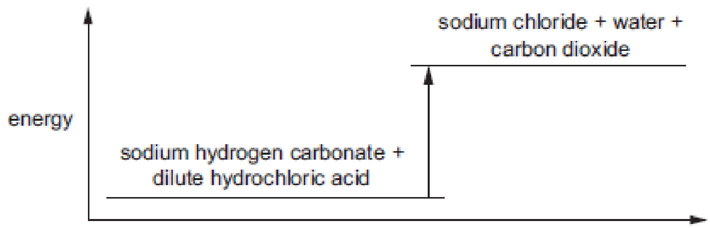
1.

The energy level diagram for the reaction between sodium hydrogen carbonate and dilute hydrochloric acid is shown.



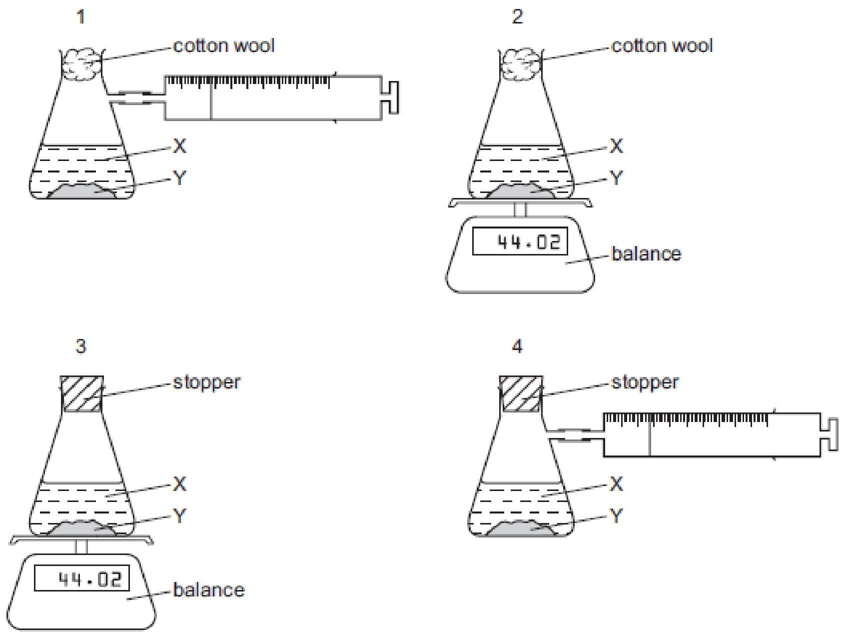
Which row correctly describes the type of reaction and the energy of the reactants and products?

	type of reaction	energy of the reactants and products
A	endothermic	the products have more energy than the reactants
B	endothermic	the reactants have more energy than the products
C	exothermic	the products have more energy than the reactants
D	exothermic	the reactants have more energy than the products

2.

A liquid X reacts with solid Y to form a gas.

Which two diagrams show suitable methods for investigating the rate (speed) of the reaction?



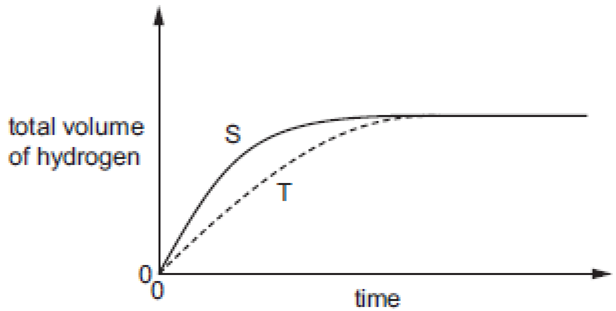
A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

3.

An experiment, S, is carried out to measure the volume of hydrogen produced when excess dilute sulfuric acid is added to zinc.

A second experiment, T, is carried out using the same mass of zinc but under different conditions.

The results of the two experiments are shown.



Which changes in the conditions between experiments S and T give curve T?

	addition of a catalyst	the zinc is in large pieces not powdered
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

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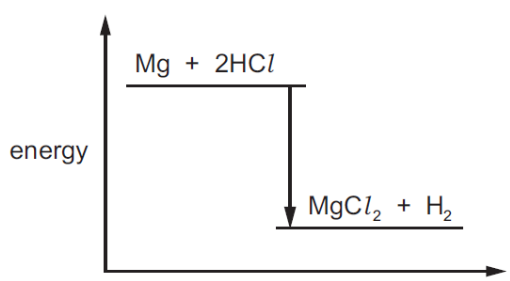
4.

The energy level diagram for the reaction between magnesium and hydrochloric acid is shown.



Which statement about the reaction is **not** correct?

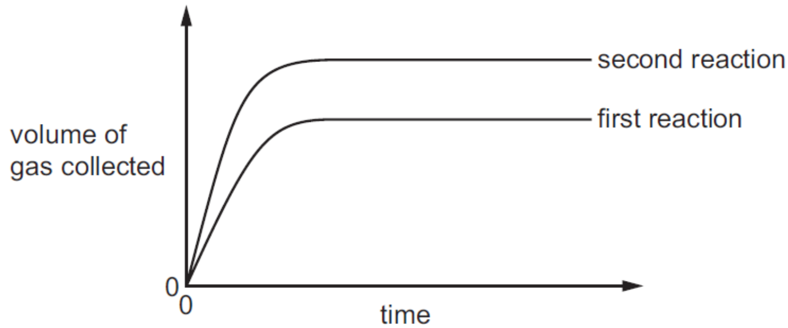
A Energy is given out during the reaction.

B The products are at a lower energy level than the reactants.

C The reaction is endothermic.

D The temperature increases during the reaction.

The results of two separate reactions between excess calcium carbonate and hydrochloric acid are shown.



Which statement explains the differences between the reactions?

A More calcium carbonate was used in the second reaction.

B The same volume of more concentrated acid was used in the second reaction.

C The second reaction was allowed to react for longer.

D The temperature was higher in the second reaction.

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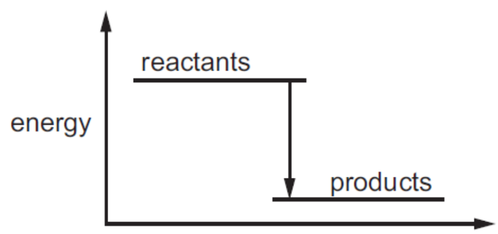
The energy level diagram shows the energy of the reactants and products in a chemical reaction.



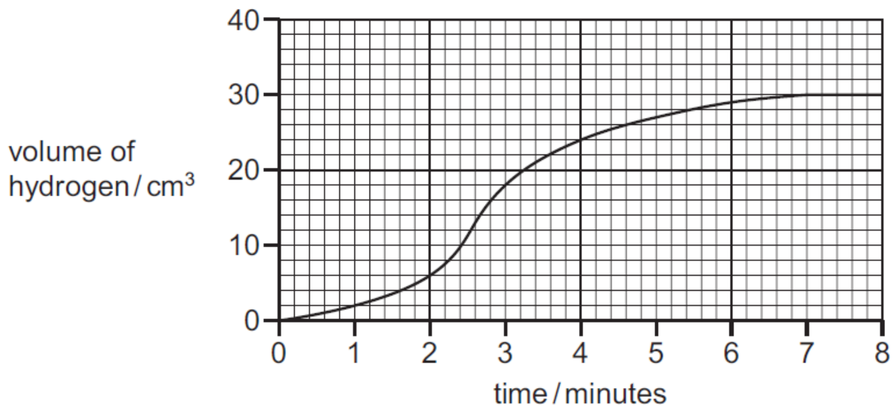
Which row correctly describes the energy change and the type of reaction shown?

	energy change	type of reaction
A	energy is given out to the surroundings	endothermic
B	energy is given out to the surroundings	exothermic
C	energy is taken in from the surroundings	endothermic
D	energy is taken in from the surroundings	exothermic

Magnesium is reacted with a dilute acid.

The hydrogen gas is collected and its volume measured.

The results are shown on the graph.



Between which times was the reaction fastest?

A 0 and 1 minute

B 1 and 2 minutes

C 2 and 3 minutes

D 7 and 8 minutes

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Hydrogen burns exothermically in oxygen.

The equation for the reaction is:

$$2H_2 + O_2 \rightarrow 2H_2O$$

The table shows the bond energies involved.

bond	bond energy in kJ / mol
H–H	436
O=O	498
O–H	464

What is the energy given out during the reaction?

A –3226 kJ / mol

B –884 kJ / mol

C –486 kJ / mol

D –442 kJ / mol

Which statements explain why increasing temperature increases the rate of a chemical reaction?

1 Heat makes the molecules move faster and collide more often.

2 Heat makes the molecules collide with more energy so they are more likely to react.

3 Increasing temperature lowers the activation energy for the reaction.

A 1 and 2 **B** 1 and 3 **C** 1 only **D** 2 only

Steam reacts with carbon in an endothermic reaction.

$$C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$$

Which conditions of temperature and pressure would give the largest yield of hydrogen?

	temperature	pressure
A	high	high
B	high	low
C	low	high
D	low	low

The equation for an exothermic reaction in the Contact process is shown.

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

Which effects do increasing the temperature and using a catalyst have on the rate of formation of sulfur trioxide, SO_3?

	increasing the temperature	using a catalyst
A	rate decreases	rate decreases
B	rate decreases	rate increases
C	rate increases	rate decreases
D	rate increases	rate increases

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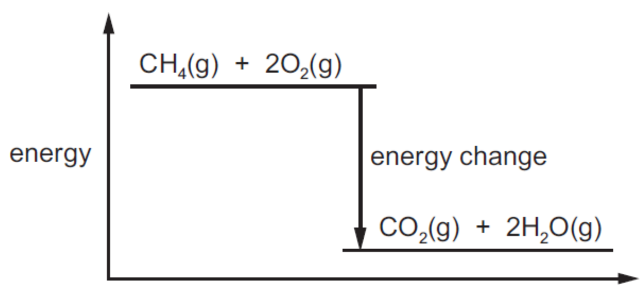
The energy level diagram for the combustion of methane is shown.



Which row gives the equation and energy change for this reaction?

	equation	energy change in kJ / mol
A	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$	+891
B	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$	–891
C	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	+891
D	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	–891

Ammonia is manufactured by a reversible reaction.

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

The forward reaction is exothermic.

What is the effect of increasing the pressure on the percentage yield and rate of formation of ammonia?

	percentage yield	rate of formation
A	decreases	decreases
B	decreases	increases
C	increases	decreases
D	increases	increases

Which row describes how the energy of collision between particles changes when concentration and temperature are increased?

	concentration	temperature
A	increases	increases
B	increases	no change
C	no change	increases
D	no change	no change

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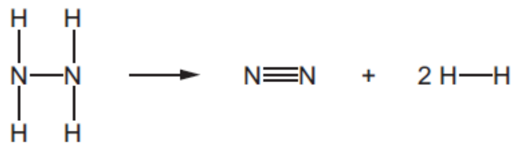
Hydrazine, N_2H_4, decomposes as shown.



The energy change for this reaction is –95 kJ / mol.

The table shows some bond energies involved.

bond	bond energy in kJ / mol
N≡N	945
N–H	391
H–H	436

What is the bond energy of the N–N bond?

A 158 kJ / mol **B** 315 kJ / mol **C** 348 kJ / mol **D** 895 kJ / mol

Which row explains why increasing temperature increases the rate of reaction?

	particles collide more often	particles collide with more energy
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Methanol is manufactured by reacting carbon monoxide and hydrogen together in the presence of an aluminium oxide catalyst.

The equation for the reaction is shown.

$$CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$$

The reaction is a reversible reaction.

The forward reaction is exothermic.

Which change in conditions increases the yield of methanol?

A decreasing the concentration of the carbon monoxide

B increasing the pressure

C increasing the rate of the reaction

D increasing the temperature

Ammonia is produced by the Haber process.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g).$$

Which statement about the Haber process is **not** correct?

A An iron catalyst is used to increase the rate of reaction.

B The reaction is carried out at high temperature to increase the rate of reaction.

C The reaction is carried out at low pressure to increase the yield of ammonia.

D The reaction is reversible.

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Ammonia is manufactured by the reaction of nitrogen with hydrogen in the presence of a catalyst.

(a) What is the purpose of a catalyst?

.... [1]

(b) The reaction is reversible.

Complete the equation below by adding the sign for a reversible reaction.

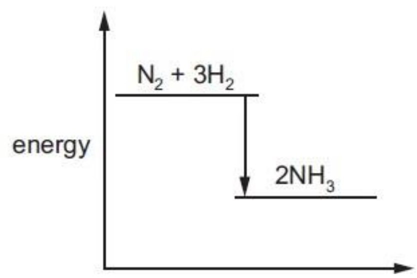
$$N_2 \; + \; 3H_2 \qquad\qquad\qquad 2NH_3$$

[1]

(c) The energy level diagram for this reaction is shown.

Is this reaction exothermic or endothermic?

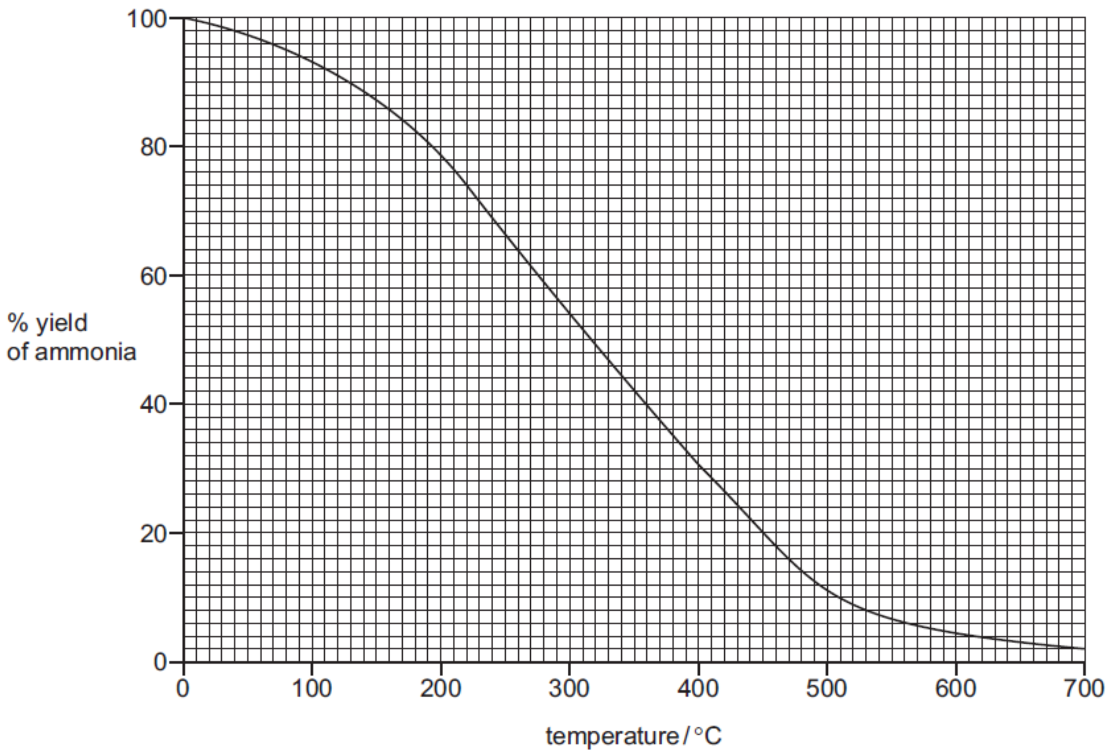
Give a reason for your answer.



....

.... [1]

(d) The graph shows how the percentage yield of ammonia changes with temperature when the pressure is kept constant.



(i) Describe how the percentage yield of ammonia changes with temperature.

.... [1]

(ii) Determine the percentage yield of ammonia at 350 °C.

.... [1]

Calcium carbonate reacts with dilute hydrochloric acid.

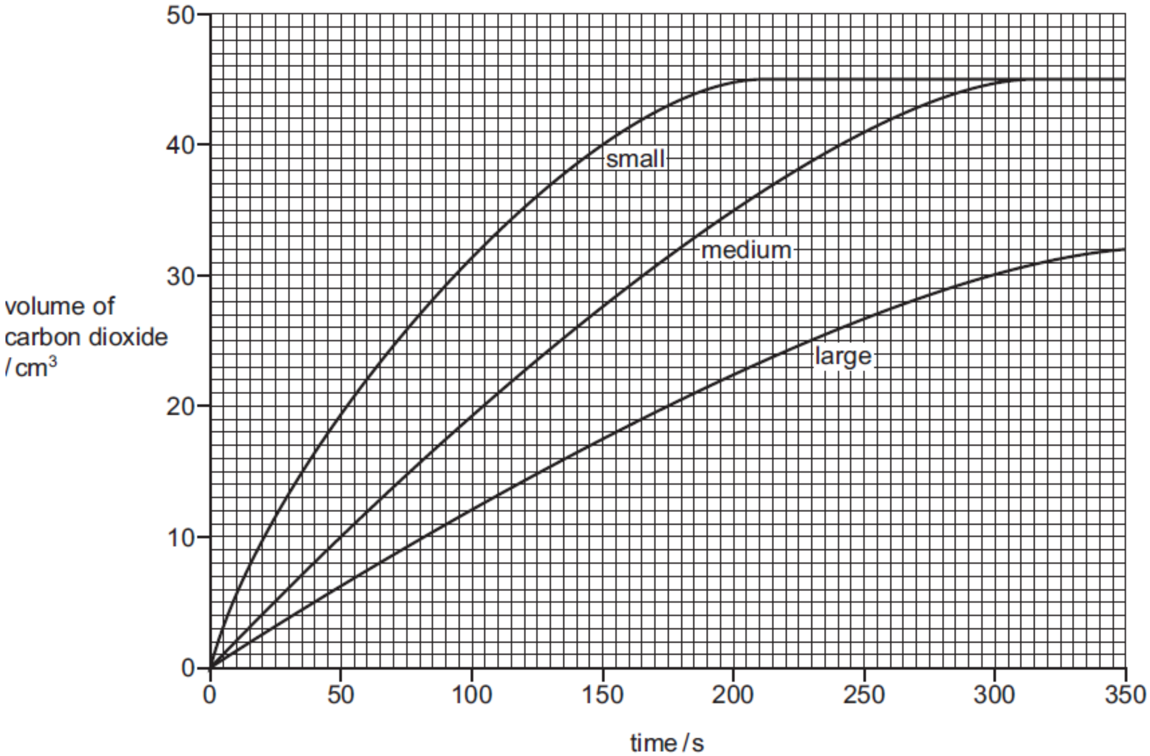
$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

A student investigated this reaction by measuring the volume of carbon dioxide released every minute at constant temperature.

(a) Draw a diagram of the apparatus that the student could use to investigate this reaction.

[2]

(b) The graph shows the results of this reaction using three samples of calcium carbonate of the same mass: large pieces, medium-sized pieces and small pieces.



(i) Which sample, large, medium or small pieces, gave the fastest initial rate of reaction?

Use the graph to explain your answer.

....

.... [2]

(ii) The experiment was repeated using powdered calcium carbonate of the same mass. Draw a line **on the grid above** to show how the volume of carbon dioxide changes with time for this experiment. [2]

(iii) At what time was the reaction just complete when small pieces of calcium carbonate were used?

.... [1]

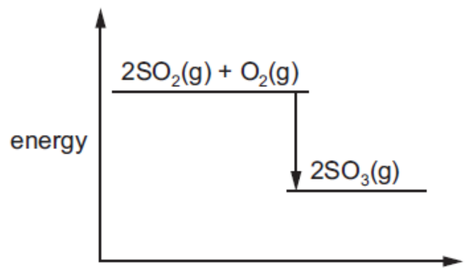
Sulfur dioxide reacts with excess oxygen to form sulfur trioxide.

$$2SO_2(g) \quad + \quad O_2(g) \quad \rightleftharpoons \quad 2SO_3(g)$$

(a) What is the meaning of the symbol $\rightleftharpoons$?

.... [1]

(b) The energy level diagram for the reaction is shown.



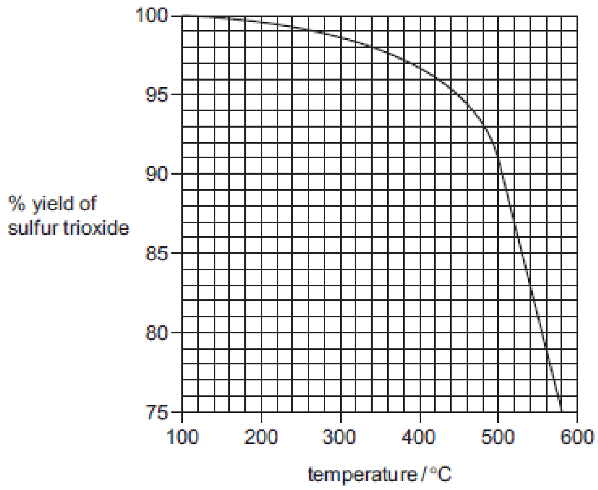
Is this reaction exothermic or endothermic?

Give a reason for your answer.

....

.... [1]

(c) The graph shows how the percentage yield of sulfur trioxide changes with temperature when the pressure is kept constant.



(i) Describe how the percentage yield of sulfur trioxide changes with temperature.

.... [1]

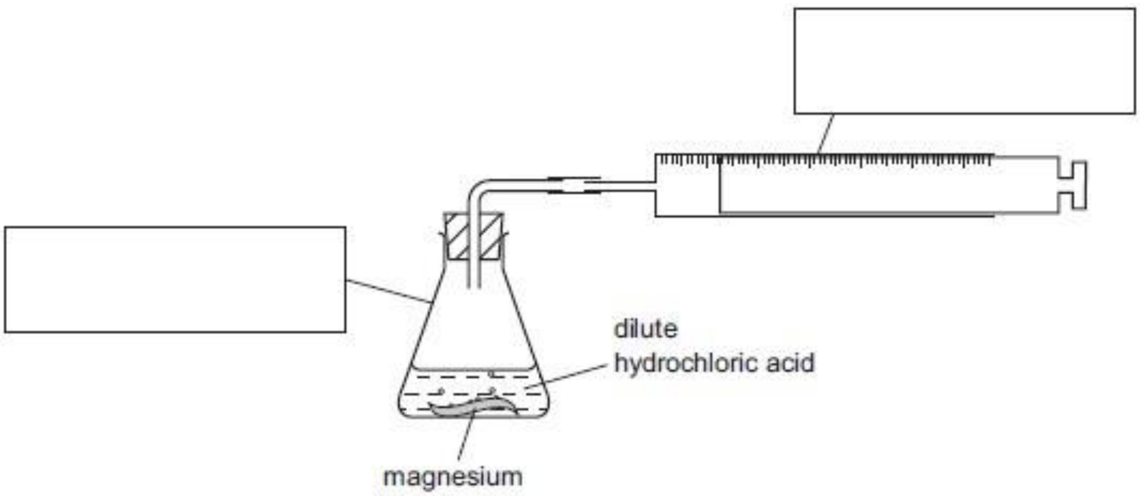
(ii) Determine the percentage yield of sulfur trioxide when the temperature is 500 °C.

.... [1]

When magnesium reacts with hydrochloric acid, the products are aqueous magnesium chloride and hydrogen.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

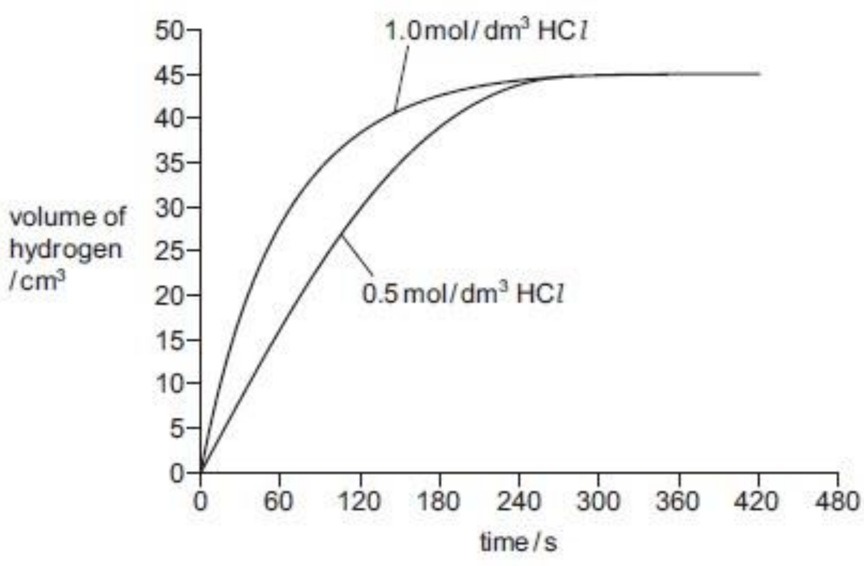
A student used the apparatus shown to follow the progress of this reaction.



(a) Complete the diagram by putting the correct labels in the boxes.

[2]

(b) The student conducted two experiments using the same mass of magnesium in each experiment and two different concentrations of hydrochloric acid. The hydrochloric acid was in excess. All other conditions were kept constant.
The student measured the volume of hydrogen produced over a period of time. The graph shows the results.



(i) Which concentration of hydrochloric acid gave the faster initial rate of reaction?

Use the graph to explain your answer.

....

.... [1]

(ii) Draw a curve **on the graph on page 16** to show how the volume of hydrogen would change if a third experiment was carried out using 1.5 mol / dm^3 hydrochloric acid and the same mass of magnesium.

[2]

0620/33 May/June 2016

Ethanol can be manufactured by reacting ethene with steam.

$$C_2H_4(g) + H_2O(g) \rightleftharpoons C_2H_5OH(g)$$

(a) What is the meaning of the symbol $\rightleftharpoons$?

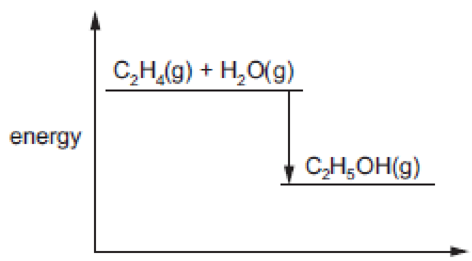
.... [1]

(b) State two conditions needed for this reaction.

....

.... [2]

(c) The energy level diagram for this reaction is shown.



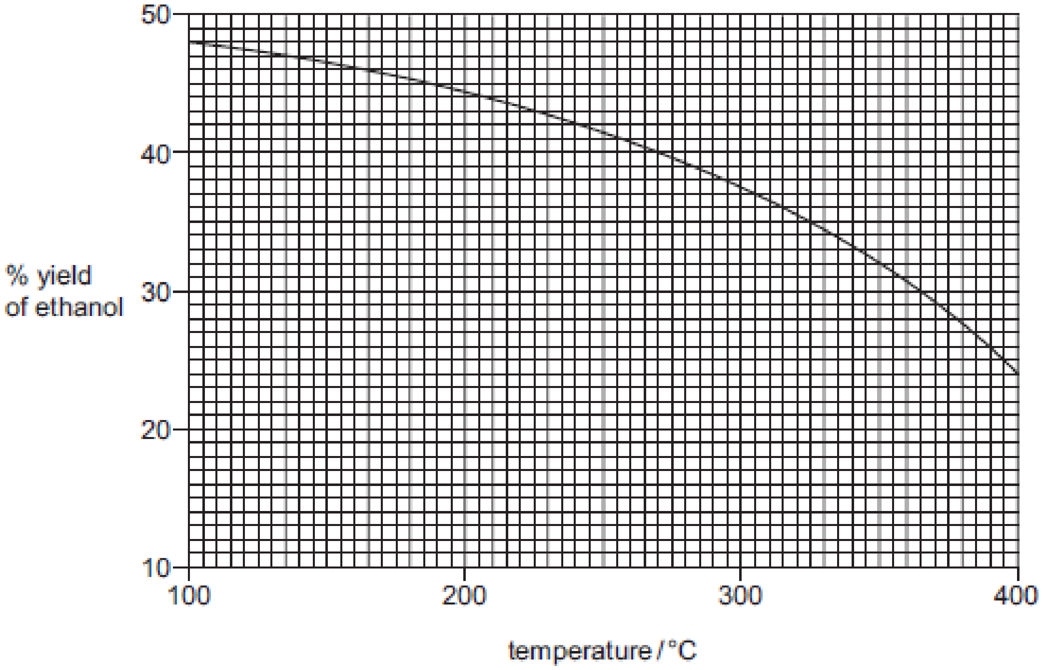
Is this reaction exothermic or endothermic?

Give a reason for your answer.

....

.... [2]

(d) The graph below shows how the percentage yield of ethanol changes with temperature when the pressure is kept constant.



(i) Describe how the percentage yield changes with temperature.

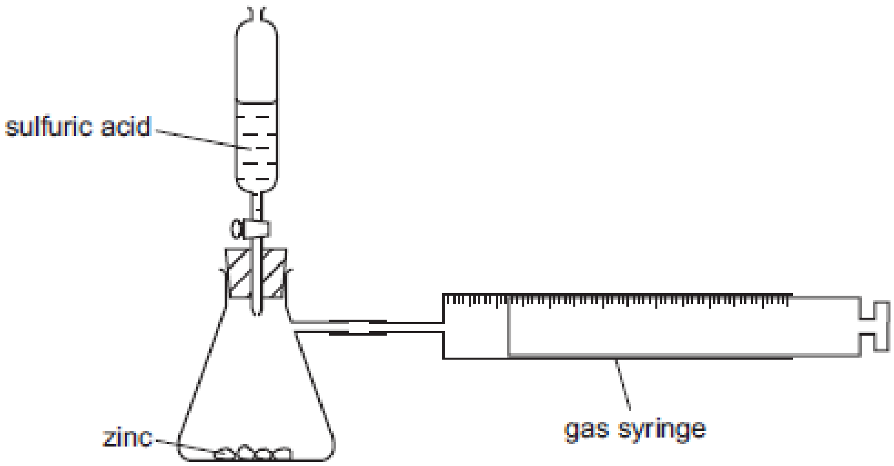
....

.... [1]

(ii) Determine the percentage yield when the temperature is 350 °C.

.... [1]

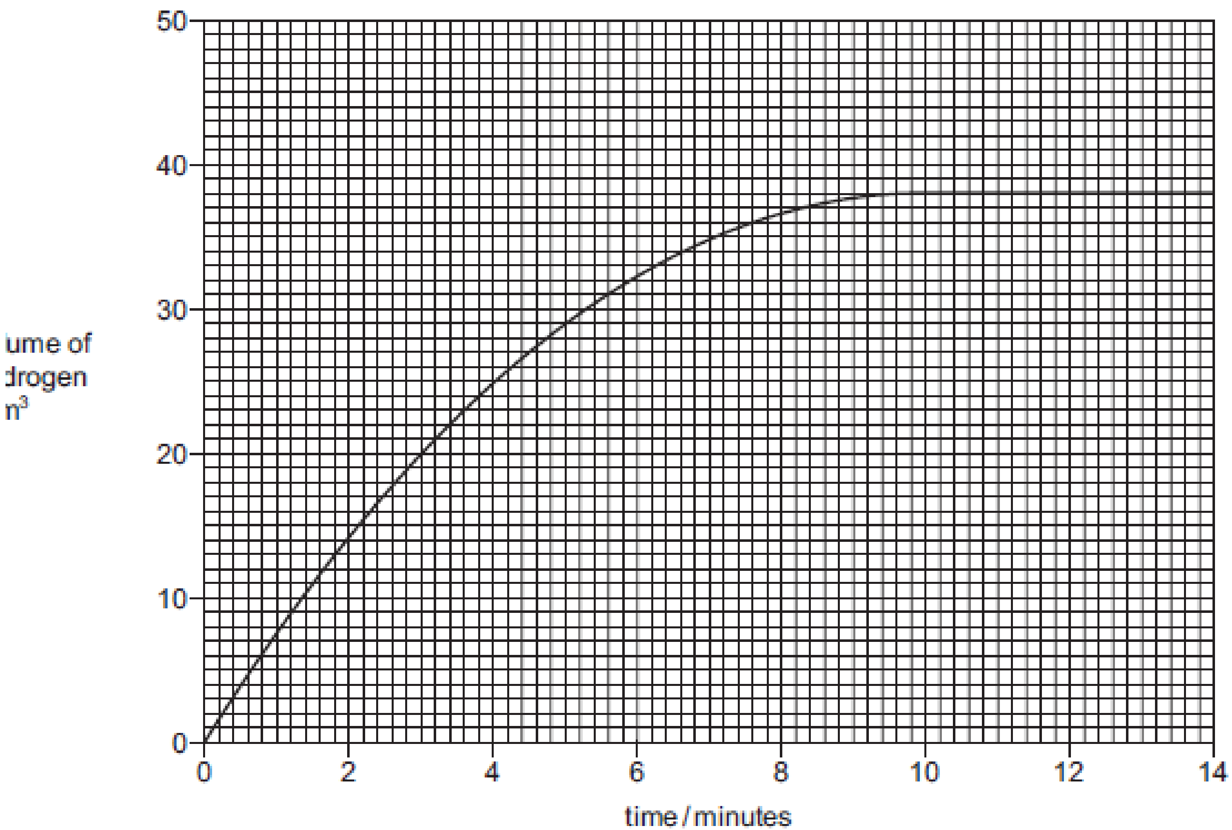
A student investigated the reaction between zinc and sulfuric acid at 20 °C using the apparatus shown. The zinc was in excess.



(a) What should the student do to start the reaction?

.... [1]

(b) The graph shows the volume of hydrogen released as the reaction proceeds.



(i) Explain why the volume of gas stays the same after 10 minutes.

.... [1]

(ii) How long did it take for the first 20 cm^3 of gas to be collected?

.... [1]

(iii) The student repeated the experiment at 30 °C. All other conditions remained the same.

Draw the shape of the line **on the grid on page 16** when the reaction was carried out at 30 °C.

[2]

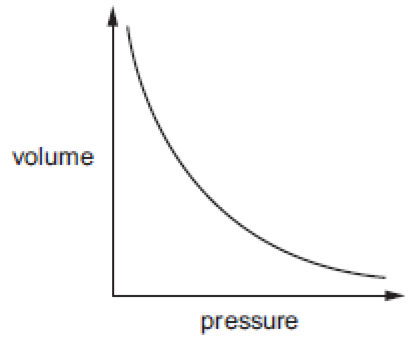
(c) The student repeated the experiment using zinc powder instead of small pieces of zinc.

Describe and explain how the rate of reaction differs when zinc powder is used.

....

.... [2]

The graph shows how increasing the pressure at constant temperature changes the volume of a fixed mass of carbon dioxide gas.



(a) Describe how the volume of gas changes with pressure.

....

....

.... [2]

(b) What happens to the average distance of the molecules from each other when the pressure is decreased?

.... [1]

0620/12 February/March 2015

Limestone can be changed into slaked lime in two chemical reactions.

1 When limestone, $CaCO_3$, is heated it decomposes into lime, CaO.

2 Water is slowly dripped onto the cooled lime. The lime appears to expand and steam is produced. Slaked lime, $Ca(OH)_2$, is formed.

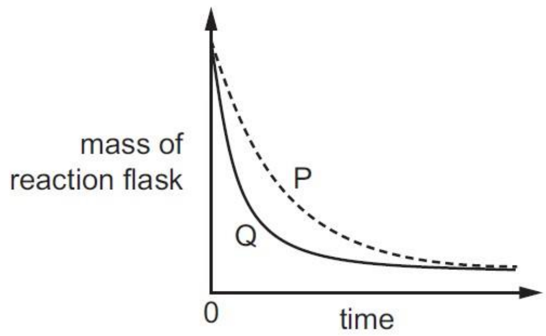
Which row shows the correct description of each of the chemical reactions?

	reaction 1	reaction 2
A	endothermic	endothermic
B	endothermic	exothermic
C	exothermic	endothermic
D	exothermic	exothermic

A student investigates the rate of reaction between marble chips and hydrochloric acid.

The mass of the reaction flask is measured.

The graph shows the results of two experiments, P and Q.



Which change explains the difference between P and Q?

A A catalyst is added in P.

B A higher temperature is used in P.

C Bigger marble chips are used in Q.

D Hydrochloric acid is more concentrated in Q.

Hydrated cobalt(II) chloride decomposes on heating.

The equation for the reaction is

$$CoCl_2.6H_2O \rightleftharpoons CoCl_2 + 6H_2O$$

The reaction is reversed by adding water.

Which row describes the colour change and the type of reaction for the **reverse** reaction?

	colour change	type of reaction
A	blue to pink	endothermic
B	blue to pink	exothermic
C	pink to blue	endothermic
D	pink to blue	exothermic

Ammonia is manufactured by the Haber process. Nitrogen and hydrogen are passed over a catalyst at a temperature of 450 °C and a pressure of 200 atmospheres.

The equation for the reaction is as follows.

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

The forward reaction is exothermic.

(a) State **one** use of ammonia.

.... [1]

(b) What is the meaning of the symbol $\rightleftharpoons$?

.... [1]

(c) What are the sources of nitrogen and hydrogen used in the Haber process?

nitrogen ....

hydrogen ....

[2]

(d) Name the catalyst in the Haber process.

.... [1]

(e) (i) If a temperature higher than 450 °C was used in the Haber process, what would happen to the **rate** of the reaction? Give a reason for your answer.

....

....

.... [2]

(ii) If a temperature higher than 450 °C was used in the Haber process, what would happen to the **yield** of ammonia? Give a reason for your answer.

....

....

.... [2]

(f) **(i)** If a pressure higher than 200 atmospheres was used in the Haber process, what would happen to the **yield** of ammonia? Give a reason for your answer.

....

....

.... [2]

(ii) Explain why the rate of reaction would be faster if the pressure was greater than 200 atmospheres.

....

.... [1]

(iii) Suggest **one** reason why a pressure higher than 200 atmospheres is not used in the Haber process.

....

.... [1]

(h) Ammonia acts as a base when it reacts with sulfuric acid.

(i) What is a base?

.... [1]

(ii) Write a balanced equation for the reaction between ammonia and sulfuric acid.

.... [2]

[Total: 18]

Which statements about exothermic and endothermic reactions are correct?

1 During an exothermic reaction, heat is given out.

2 The temperature of an endothermic reaction goes up because heat is taken in.

3 Burning methane in the air is an exothermic reaction.

A 1, 2 and 3 **B** 1 and 2 only **C** 1 and 3 only **D** 2 and 3 only

The rate of a reaction depends on temperature, concentration, particle size and catalysts.

Which statement is **not** correct?

A Catalysts can be used to increase the rate of reaction.

B Higher concentration decreases the rate of reaction.

C Higher temperature increases the rate of reaction.

D Larger particle size decreases the rate of reaction.

The diagram shows the change from an anhydrous salt to its hydrated form.

$$\boxed{\text{anhydrous salt}} \underset{\text{reverse}}{\overset{\text{forward}}{\rightleftharpoons}} \boxed{\text{hydrated salt}}$$

Which statement is correct?

A forward reaction requires heat and water

B forward reaction requires water only

C reverse reaction requires heat and water

D reverse reaction requires water only

Which changes decrease the rate of reaction between magnesium and air?

1 heating the magnesium to a higher temperature

2 using a higher proportion of oxygen in the air

3 using magnesium ribbon instead of powdered magnesium

A 1, 2 and 3 **B** 1 only **C** 2 only **D** 3 only

Which statements are properties of an acid?

1 reacts with ammonium sulfate to form ammonia

2 turns red litmus blue

	1	2
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

A method used to make copper(II) sulfate crystals is shown.

1 Place dilute sulfuric acid in a beaker.

2 Warm the acid.

3 Add copper(II) oxide until it is in excess.

4 Filter the mixture.

5 Evaporate the filtrate until crystals start to form.

6 Leave the filtrate to cool.

What are the purposes of step 3 and step 4?

	step 3	step 4
A	to ensure all of the acid has reacted	to obtain solid copper(II) sulfate
B	to ensure all of the acid has reacted	to remove excess copper(II) oxide
C	to speed up the reaction	to obtain solid copper(II) sulfate
D	to speed up the reaction	to remove excess copper(II) oxide

The results of two tests on solid X are shown.

test	observation
aqueous sodium hydroxide added	green precipitate formed
acidified silver nitrate added	yellow precipitate formed

What is X?

A copper(II) chloride

B copper(II) iodide

C iron(II) chloride

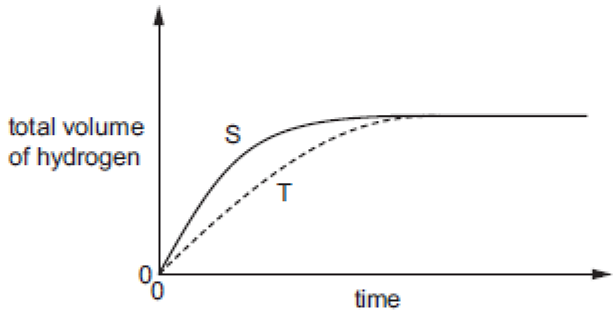
D iron(II) iodide

0620/11 Paper 1 Multiple Choice (Core) May/June 2016

An experiment, S, is carried out to measure the volume of hydrogen produced when excess dilute sulfuric acid is added to zinc.

A second experiment, T, is carried out using the same mass of zinc but under different conditions.

The results of the two experiments are shown.



Which changes in the conditions between experiments S and T give curve T?

	addition of a catalyst	the zinc is in large pieces not powdered
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Some chemical properties of three metals W, X and Y and their oxides are shown.

metal	reaction with steam	reaction with dilute hydrochloric acid	reaction of metal oxide with carbon
W	reacts	reacts	reacts
X	no reaction	no reaction	reacts
Y	reacts	reacts	no reaction

What is the order of reactivity of these metals, most reactive first?

A W → Y → X

B X → Y → W

C Y → W → X

D Y → X → W

Which statements are properties of an acid?

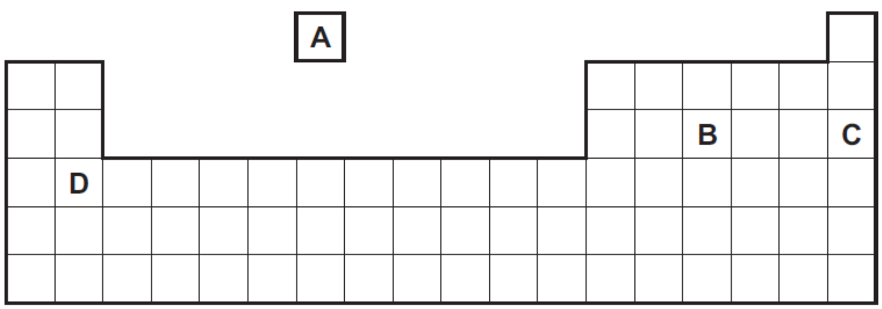
1 reacts with ammonium sulfate to form ammonia

2 turns red litmus blue

	1	2
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Part of the Periodic Table is shown.

Which element forms an acidic oxide?



Salts can be made by adding different substances to dilute hydrochloric acid.

For which substance could any excess **not** be removed by filtration?

A copper(II) oxide

B magnesium

C sodium hydroxide

D zinc hydroxide

A solution containing substance X was tested. The table shows the results.

test	result
flame test	lilac colour
acidified silver nitrate solution added	yellow precipitate

What is X?

A lithium bromide

B lithium iodide

C potassium bromide

D potassium iodide

0620/12 May/June 2016

Which statements are properties of an acid?

1 reacts with ammonium sulfate to form ammonia

2 turns red litmus blue

	1	2
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Which row describes whether an amphoteric oxide reacts with acids and bases?

	reacts with acids	reacts with bases
A	no	no
B	no	yes
C	yes	no
D	yes	yes

Which substance reacts with dilute sulfuric acid to form a salt that can be removed from the resulting mixture by filtration?

A aqueous barium chloride

B aqueous sodium hydroxide

C copper

D copper(II) carbonate

0620/22 May/June 2016

Which statements are properties of an acid?

1 reacts with ammonium sulfate to form ammonia

2 turns red litmus blue

	1	2
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Which row describes whether an amphoteric oxide reacts with acids and bases?

	reacts with acids	reacts with bases
A	no	no
B	no	yes
C	yes	no
D	yes	yes

Silver chloride is insoluble in water and is prepared by precipitation.

Which two substances can be used to make silver chloride?

A barium chloride and silver nitrate

B hydrochloric acid and silver

C hydrochloric acid and silver bromide

D sodium chloride and silver iodide

0620/23

May/June 2016

Which statements are properties of an acid?

1 reacts with ammonium sulfate to form ammonia

2 turns red litmus blue

	1	2
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

Which row describes whether an amphoteric oxide reacts with acids and bases?

	reacts with acids	reacts with bases
A	no	no
B	no	yes
C	yes	no
D	yes	yes

Barium sulfate is an insoluble salt.

It can be made by reacting copper(II) sulfate solution with barium nitrate solution.

$$CuSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Cu(NO_3)_2(aq) + BaSO_4(s)$$

What is the correct order of steps to obtain a pure, dry sample of barium sulfate from the reaction mixture?

	step 1	step 2	step 3
A	filter	evaporate the filtrate to dryness	leave the solid formed to cool
B	filter	evaporate the filtrate to the point of crystallisation	leave the filtrate to cool
C	filter	leave the residue in a warm place to dry	wash the residue with water
D	filter	wash the residue with water	leave the residue in a warm place to dry

0620/31 May/June 2016

Iodine reacts with hot concentrated nitric acid.

$$I_2 + 10HNO_3 \rightarrow 2HIO_3 + 4H_2O + 10NO_2$$

(i) Explain why this reaction could have an adverse effect on health if not carried out in a fume cupboard.

....

.... [2]

(ii) Nitric acid is strongly acidic.

Which one of the following pH values represents a strongly acidic solution?

Put a ring around the correct answer.

pH 1 pH 7 pH 9 pH 13

[1]

(iii) Nitric acid reacts with zinc oxide.

State the names of the products of this reaction.

.... and .... [2]

Ammonia is a weak base.

Describe how you would measure the pH of an aqueous solution of a weak base using Universal Indicator.

....

.... [2]

Complete the chemical equation for the reaction of ammonia with chlorine.

$$\text{........}NH_3 + 3Cl_2 \rightarrow N_2 + \text{......}HCl$$ [2]

When calcium carbonate is heated strongly, calcium oxide is formed.

(i) Give **one** use of calcium oxide.

.... [1]

(ii) What type of oxide is calcium oxide?

Explain your answer.

....

.... [2]

0620/32 May/June 2016

(a) The table describes the ease of reduction of some metal oxides with carbon.

metal oxide	ease of reduction on heating
lead oxide	moderate heating at 200 °C needed
nickel oxide	high temperature at 750 °C needed
titanium oxide	very high temperatures above 1700 °C needed
zinc oxide	very high temperature at 900 °C needed

Put the metals in order of their reactivity. Put the least reactive metal first.

least reactive → most reactive

[2]

Sulfur dioxide reacts with water to form sulfurous acid, H_2SO_3. Sulfurous acid reacts with hydrogen sulfide to form water and sulfur.

Complete the chemical equation for this reaction.

$$H_2SO_3 + 2H_2S \rightarrow .....H_2O + 3S$$

[1]

Three chemicals, P, Q and R, were each dissolved in water. The table shows some of the reactions of these solutions.

solution	reaction when solid sodium carbonate is added	reaction when heated with solid ammonium chloride
P	gas evolved	no reaction
Q	no reaction	gas evolved
R	no reaction	no reaction

The pH of the three solutions was also measured.

What are the correct pH values of these solutions?

	P	Q	R
A	2	7	13
B	2	13	7
C	7	2	13
D	13	7	2

The oxide of element X forms a solution with pH 4.

The oxide of element Y forms a solution that turns Universal Indicator blue.

Which row correctly classifies elements X and Y?

	element X	element Y
A	metal	metal
B	metal	non-metal
C	non-metal	metal
D	non-metal	non-metal

Which two processes are involved in the preparation of magnesium sulfate from dilute sulfuric acid and an excess of magnesium oxide?

A neutralisation and filtration

B neutralisation and oxidation

C thermal decomposition and filtration

D thermal decomposition and oxidation

Which statement about aqueous sodium hydroxide is correct?

A When it is added to a solution containing sulfate ions, a white precipitate is formed.

B When it is added to a solution of copper(II) ions, a blue precipitate is formed which dissolves in excess to give deep blue solution.

C When it is added to a solution of iron(II) ions, a green precipitate is formed which does not dissolve in excess.

D When it is added to ammonium chloride, a gas is produced which turns blue litmus red.

The list gives four experiments carried out with calcium carbonate.

1 acid added

2 alkali added

3 heated strongly

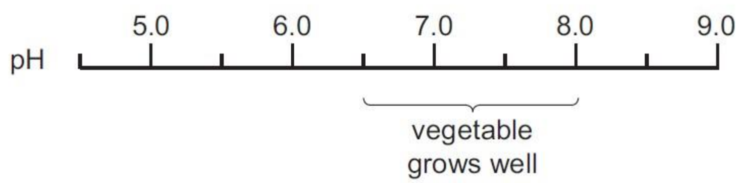
4 water added

Which experiments produce carbon dioxide?

A 1 and 2 **B** 1 and 3 **C** 2 and 3 **D** 2 and 4

The diagram shows the soil pH range over which a vegetable grows well.

The pH of the soil to be used is 5.5.



Why is lime added to the soil before planting the vegetable?

A The lime acts as a catalyst.

B The lime changes the soil acidity.

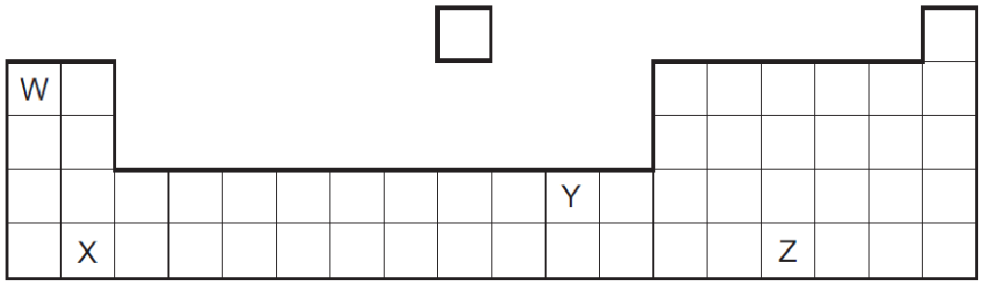
C The lime is an indicator.

D The lime supplies nitrogen.

Which substance is the most acidic?

	substance	pH
A	calcium hydroxide	12
B	lemon juice	4
C	milk	6
D	washing up liquid	8

The positions of elements W, X, Y and Z in the Periodic Table are shown.



Which elements form basic oxides?

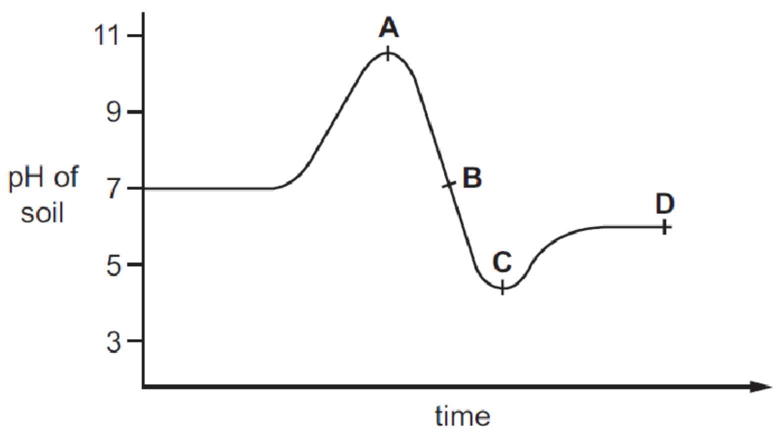
A W, X and Y **B** W and X only **C** Y only **D** Z only

How many different salts could be made from a supply of dilute sulfuric acid, dilute hydrochloric acid, copper, magnesium oxide and zinc carbonate?

A 3 **B** 4 **C** 5 **D** 6

The graph shows how the pH of soil in a field changes over time.

At which point was the soil neutral?



NOTES FOR USE IN QUALITATIVE ANALYSIS
Test for anions

anion	test	test result
carbonate (CO_3^{2-})	add dilute acid	effervescence, carbon dioxide produced
chloride (Cl^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	white ppt.
bromide (Br^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	cream ppt.
iodide (I^-) [in solution]	acidify with dilute nitric acid, then add aqueous silver nitrate	yellow ppt.
nitrate (NO_3^-) [in solution]	add aqueous sodium hydroxide then aluminium foil; warm carefully	ammonia produced
sulfate (SO_4^{2-}) [in solution]	acidify, then add aqueous barium nitrate	white ppt.
sulfite (SO_3^{2-})	add dilute hydrochloric acid, warm gently and test for the presence of sulfur dioxide	sulfur dioxide produced will turn acidified aqueous potassium manganate(VII) from purple to colourless

Test for aqueous cations

cation	effect of aqueous sodium hydroxide	effect of aqueous ammonia
aluminium (Al^{3+})	white ppt., soluble in excess giving a colourless solution	white ppt., insoluble in excess
ammonium (NH_4^+)	ammonia produced on warming	–
calcium (Ca^{2+})	white ppt., insoluble in excess	no ppt. or very slight white ppt.
chromium(III) (Cr^{3+})	green ppt., soluble in excess	grey-green ppt., insoluble in excess
copper (Cu^{2+})	light blue ppt., insoluble in excess	light blue ppt., soluble in excess giving a dark blue solution
iron(II) (Fe^{2+})	green ppt., insoluble in excess	green ppt., insoluble in excess
iron(III) (Fe^{3+})	red-brown ppt., insoluble in excess	red-brown ppt., insoluble in excess
zinc (Zn^{2+})	white ppt., soluble in excess giving a colourless solution	white ppt., soluble in excess, giving a colourless solution

Test for gases

gas	test and test results
ammonia (NH_3)	turns damp, red litmus paper blue
carbon dioxide (CO_2)	turns limewater milky
chlorine (Cl_2)	bleaches damp litmus paper
hydrogen (H_2)	'pops' with a lighted splint
oxygen (O_2)	relights a glowing splint
sulfur dioxide (SO_2)	turns acidified aqueous potassium manganate(VII) from purple to colourless

Flame tests for metal ions

metal ion	flame colour
lithium (Li^+)	red
sodium (Na^+)	yellow
potassium (K^+)	lilac
copper(II) (Cu^{2+})	blue-green

0620/53 October/November 2016

1 You are going to investigate what happens when two different metals, iron and magnesium, react with aqueous copper(II) sulfate.

Read all the instructions carefully before starting the experiments.

Instructions
You are going to carry out two experiments.

(a) *Experiment 1*

Use a measuring cylinder to pour 25 cm^3 of aqueous copper(II) sulfate into the polystyrene cup provided. Put the polystyrene cup into a 250 cm^3 beaker for support. Measure the initial temperature of the solution and then the temperature after 30 seconds and 60 seconds. Record your results in the table.

At 60 seconds add all of the iron to the aqueous copper(II) sulfate and stir the mixture continuously with the thermometer.

Measure the temperature of the mixture every 30 seconds for 300 seconds (5 minutes). Record your results in the table.

time / s	0	30	60	90	120	150	180	210	240	270	300
temperature / °C											

[2]

(b) *Experiment 2*

Empty the polystyrene cup and rinse it with water.

Use a measuring cylinder to pour 25 cm^3 of aqueous copper(II) sulfate into the polystyrene cup. Put the polystyrene cup into a 250 cm^3 beaker for support. Measure the initial temperature of the solution and then the temperature after 30 seconds and 60 seconds. Record your results in the table.

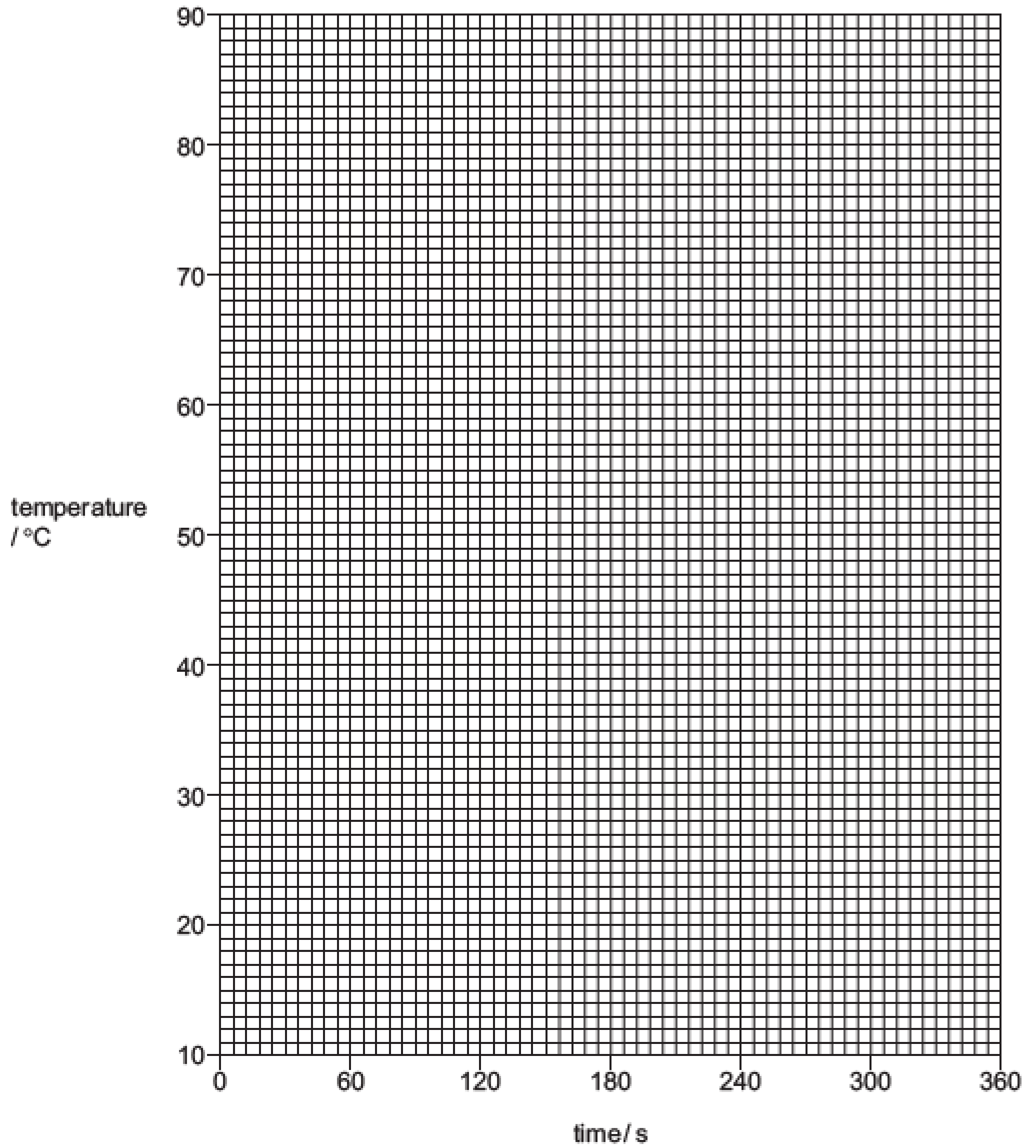
At 60 seconds add all of the magnesium to the aqueous copper(II) sulfate and stir the mixture continuously with the thermometer.

Measure the temperature of the mixture every 30 seconds for 300 seconds (5 minutes). Record your results in the table.

time / s	0	30	60	90	120	150	180	210	240	270	300
temperature / °C											

[2]

(c) Plot the results for Experiments 1 and 2 on the grid and draw **two** smooth line graphs.
Clearly label the graphs.



[4]

(d) (i) **From your graph**, deduce the temperature of the mixture in Experiment 1 after 135 seconds.
Show clearly **on the grid** how you worked out your answer.

.... °C [2]

(ii) **From your graph**, deduce the time taken for the temperature of the mixture in Experiment 2 to change by 30 °C **after the magnesium was added**.
Show clearly **on the grid** how you worked out your answer.

.... s [2]

(e) Predict the temperature of the mixture in Experiment 2 after one hour. Explain your answer.

....

.... [2]

(f) Suggest an advantage of taking the temperature readings every 15 seconds.

....

.... [2]

(g) Explain why a polystyrene cup is used in the experiments and **not** a copper can.

....

.... [2]

[Total: 18]

You are provided with two solutions, solution **Q** and solution **R**.
Carry out the following tests on solution **Q** and solution **R**, recording all of your observations at each stage.

tests on solution Q

(a) Divide solution **Q** into four equal portions in four test-tubes. Carry out the following tests.

(i) Use pH indicator paper to measure the pH of the first portion of solution **Q**.

pH .... [1]

(ii) Add a 2 cm strip of magnesium ribbon to the second portion of solution **Q**. Test the gas given off.
Record your observations.

....

.... [2]

(iii) Add a spatula measure of sodium carbonate to the third portion of solution **Q**. Test the gas given off.
Record your observations.

....

.... [2]

(iv) Add a few drops of dilute nitric acid and about 1 cm^3 of aqueous barium nitrate to the fourth portion of solution **Q**.
Record your observations.

.... [1]

tests on solution R

(b) Divide solution **R** into four equal portions in four test-tubes. Carry out the following tests.

(i) Measure the pH of the first portion of solution **R**.

pH .... [1]

(ii) Add several drops of aqueous sodium hydroxide to the second portion of solution **R** and shake the test-tube.
Then add excess aqueous sodium hydroxide to the test-tube.
Record your observations.

....

.... [2]

(iii) Add aqueous silver nitrate to the third portion of solution **R** and leave to stand for about 5 minutes.
Record your observations.

....

.... [2]

(iv) Add a spatula measure of iron(II) sulfate crystals to the fourth portion of solution **R** and shake the mixture.
Record your observations.

.... [1]

(c) Identify solution **Q**.

.... [2]

(d) Identify solution **R**.

.... [2]

[Total: 16]

A liquid cleaner is a mixture of three substances. These substances are shown in the table.

name of substance	properties of substance
water	liquid, boiling point 100 °C
sodium carbonate	solid, soluble in water
silica	solid, insoluble in water

Plan experiments to obtain separate pure samples of each substance from the mixture in the liquid cleaner. You are provided with common laboratory apparatus.

....

....

....

....

....

....

....

....

....

.... [6]

[Total: 6]

0620/61 October/November 2016

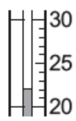
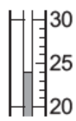
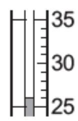
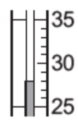
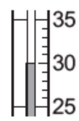
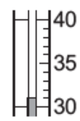
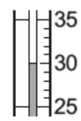
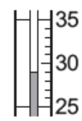
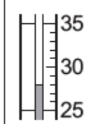
2 A student investigated what happened when dilute nitric acid reacted with aqueous solutions of two different alkalis, solution **N** and solution **O**.

Two experiments were carried out.

(a) *Experiment 1*

A measuring cylinder was used to pour 50 cm^3 of solution **N** into a polystyrene cup. The initial temperature of the solution was measured.
A burette was filled with nitric acid to the 0.0 cm^3 mark.
5.0 cm^3 of nitric acid were added to solution **N** in the polystyrene cup and the solution stirred.
The maximum temperature of the solution was measured.
A further 5.0 cm^3 of nitric acid were added to the polystyrene cup and the solution stirred. The maximum temperature of the solution was measured.
The student continued to add 5.0 cm^3 portions of nitric acid to the polystyrene cup, until a total volume of 40 cm^3 of nitric acid had been added. After each addition, the solution was stirred and the maximum temperature measured.

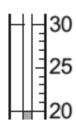
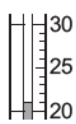
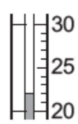
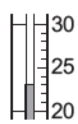
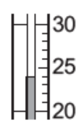
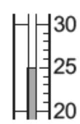
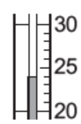
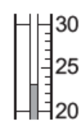
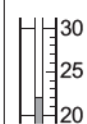
Use the thermometer diagrams to record the maximum temperatures in the table.

volume of nitric acid added / cm^3	0.0	5.0	10.0	15.0	20.0	25.0	30.0	35.0	40.0
thermometer diagram									
maximum temperature of the solution in the polystyrene cup / °C									

[2]

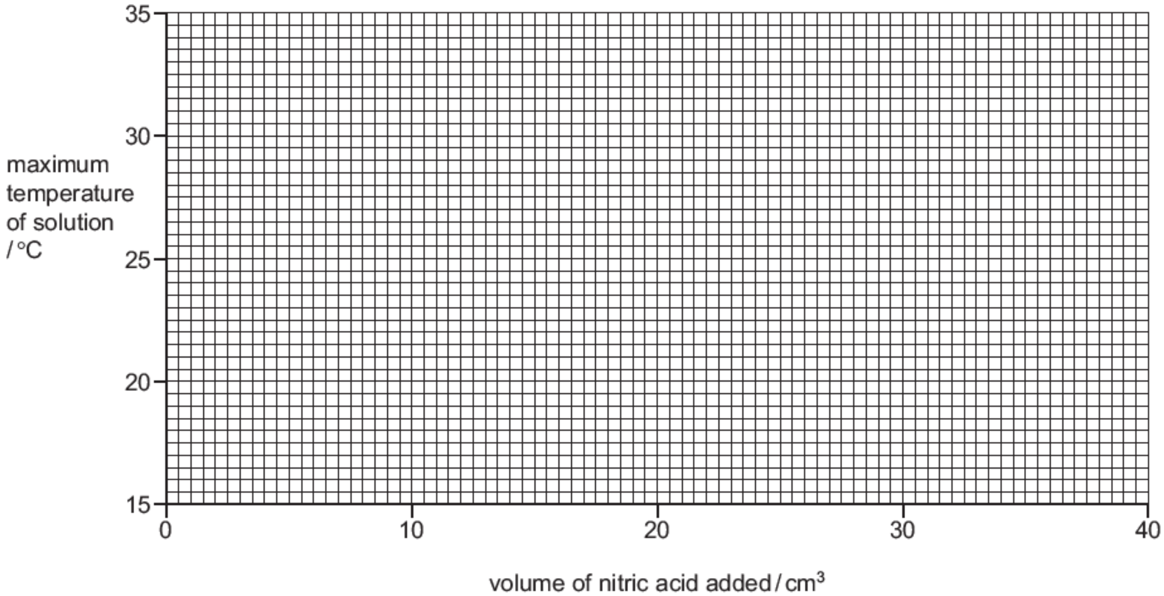
(b) *Experiment 2*

Experiment 1 was repeated using solution **O** instead of solution **N**.
Use the thermometer diagrams to record the maximum temperatures in the table.

volume of nitric acid added / cm^3	0.0	5.0	10.0	15.0	20.0	25.0	30.0	35.0	40.0
thermometer diagram									
maximum temperature of the solution in the polystyrene cup / °C									

[2]

(c) Plot the results for Experiments 1 and 2 on the grid and draw **two** smooth line graphs.
Clearly label your graphs.



[4]

(d) **Use your graph** to estimate the maximum temperature of the solution when 13 cm^3 of nitric acid were added to 50 cm^3 of solution **N** in Experiment 1.
Show clearly **on the grid** how you worked out your answer.

.... °C [2]

(e) Name a suitable indicator that could be used in Experiment 1.

.... [1]

(f) Solution **N** and solution **O** were the same concentration.

In which experiment is the temperature change greater? Suggest why the temperature change is greater in this experiment.

....

.... [2]

(g) How would the results differ in Experiment 1 if 100 cm^3 of solution **N** were used?

....

.... [1]

(h) Suggest why a polystyrene cup was used in these experiments and **not** a copper can.

.... [1]

(i) State **one** source of error in the experiments. Suggest an improvement to reduce this source of error.

source of error ....

improvement ....

[2]

[Total: 17]

3 Solid **P**, which is an aluminium salt, was analysed.
The tests on solid **P**, and some of the observations, are shown.

tests on solid P

(a) **test 1**

Solid **P** was divided into three portions. The first portion of solid **P** was heated.

observations *condensation formed on the sides of the test-tube*

Any gases given off were tested with cobalt(II) chloride paper.

observations *cobalt(II) chloride paper turned from blue to pink*

What does **test 1** tell you about solid **P**?

.... [1]

(b) **test 2**

A flame test was carried out on the second portion of solid **P**.

observations .... [1]

tests on a solution of P

Distilled water was added to the rest of solid **P** in a test-tube and shaken to dissolve.

(c) The solution was divided into four equal portions in four test-tubes. The following tests were carried out.

(i) **test 3**

Several drops of aqueous sodium hydroxide were added to the first portion of the solution.

Excess aqueous sodium hydroxide was then added to the mixture.

observations ....

....

.... [3]

(ii) **test 4**

Several drops of aqueous ammonia were added to the second portion of the solution.

Excess aqueous ammonia was then added to the mixture.

observations ....

.... [2]

Two further tests were carried out and the following observations made.

tests on a solution of **P**	observations
test 5 Dilute nitric acid and aqueous silver nitrate were added to the third portion of the solution.	no visible reaction
test 6 Dilute nitric acid and aqueous barium nitrate were added to the fourth portion of the solution.	white precipitate formed

(d) What does **test 5** tell you about solid **P**?

.... [1]

(e) Identify solid **P**.

.... [1]

(f) Describe the appearance of solid **P**.

.... [1]

[Total: 10]

Agri Limes are mixtures of calcium carbonate and calcium oxide. Farmers use Agri Limes on fields to neutralise acidity.

Plan an investigation to find out which of **two** different Agri Limes, **Q** or **R**, will neutralise more acid. You are provided with common laboratory apparatus and chemicals, including dilute nitric acid.

....

....

....

....

....

....

....

.... [6]

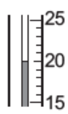
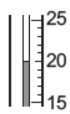
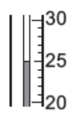
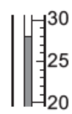
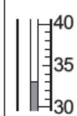
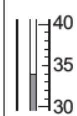
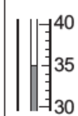
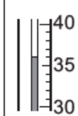
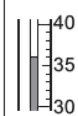
[Total: 6]

0620/63 October/November 2016

2 A student investigated what happened when two different metals, iron and magnesium, reacted with aqueous copper(II) sulfate.
Two experiments were carried out.

(a) *Experiment 1*

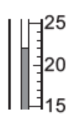
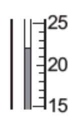
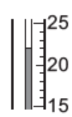
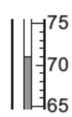
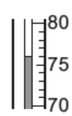
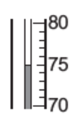
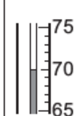
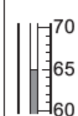
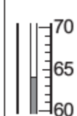
A measuring cylinder was used to pour 25 cm^3 of aqueous copper(II) sulfate into a polystyrene cup. The initial temperature of the solution was measured, then again at 30 seconds and at 60 seconds.
At 60 seconds, the iron was added to the aqueous copper(II) sulfate and the mixture stirred continuously with a thermometer.
The temperature of the mixture was measured every 30 seconds for 300 seconds (5 minutes).
Use the thermometer diagrams to record the results in the table.

time / s	0	30	60	90	120	150	180	210	240	270	300
thermometer diagram											
temperature / °C											

[2]

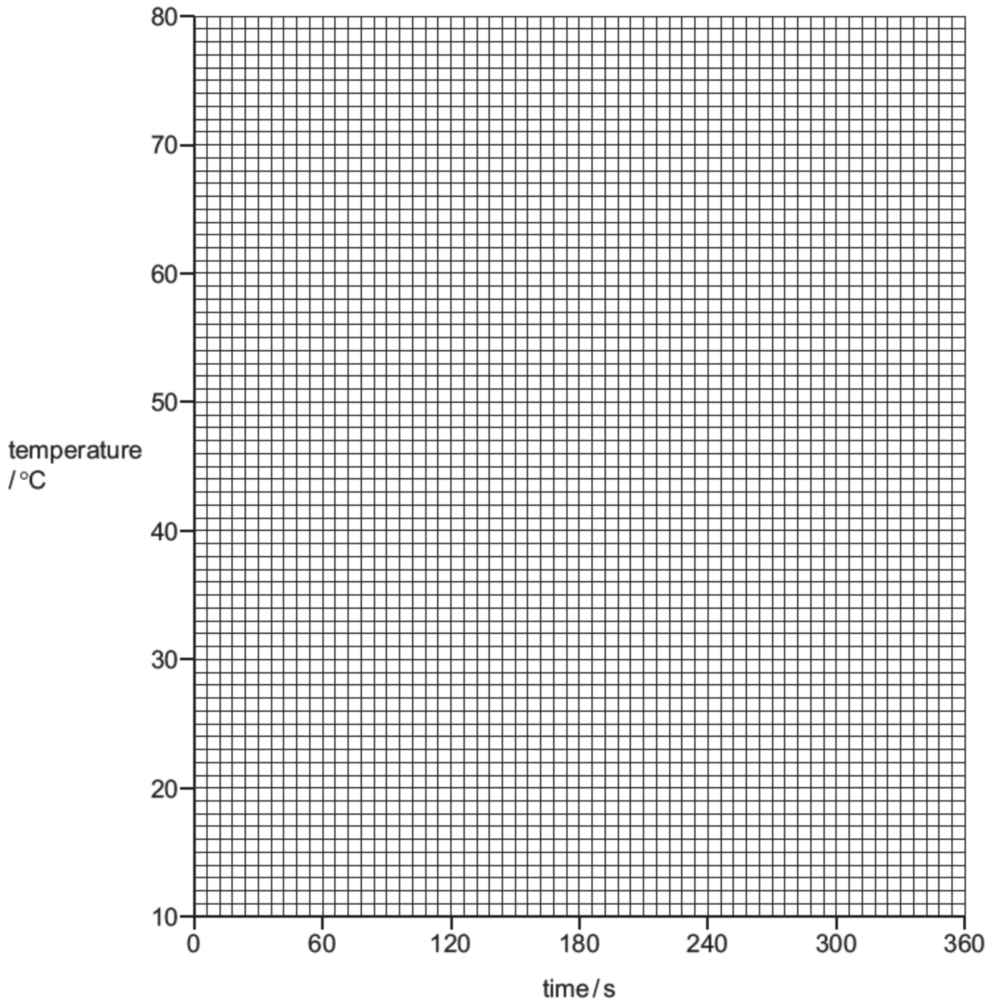
(b) *Experiment 2*

Experiment 1 was repeated using magnesium instead of iron.
Use the thermometer diagrams to record the results in the table.

time / s	0	30	60	90	120	150	180	210	240	270	300
thermometer diagram											
temperature / °C											

[2]

(c) Plot the results for Experiments 1 and 2 on the grid and draw **two** smooth line graphs. Clearly label the graphs.



[4]

(d) (i) **From your graph**, deduce the temperature of the mixture in Experiment 1 after 135 seconds.
Show clearly **on the grid** how you worked out your answer.

.... °C [2]

(ii) **From your graph**, deduce the time taken for the temperature of the mixture in Experiment 2 to change by 30 °C **after the magnesium was added**.
Show clearly **on the grid** how you worked out your answer.

.... s [2]

(e) Predict the temperature of the mixture in Experiment 2 after one hour. Explain your answer.

....

.... [2]

(f) Suggest an advantage of taking the temperature readings every 15 seconds.

....

.... [2]

(g) Explain why a polystyrene cup is used in the experiments and **not** a copper can.

....

.... [2]

[Total: 18]

3 Two solutions, solution **Q** and solution **R**, were analysed. Solution **Q** was aqueous sulfuric acid.

tests on solution Q

(a) Solution **Q** was divided into four equal portions in four test-tubes. The following tests were carried out.

Complete the observations for **tests 1–4**.

(i) **test 1**

The pH of the first portion of solution **Q** was measured.

pH .... [1]

(ii) **test 2**

Magnesium ribbon was added to the second portion of solution **Q**. The gas given off was tested.

observations ....

.... [3]

(iii) **test 3**

Sodium carbonate was added to the third portion of solution **Q**. The gas given off was tested.

observations ....

.... [3]

(iv) **test 4**

Dilute nitric acid and aqueous barium nitrate were added to the fourth portion of solution **Q**.

observations .... [1]

tests on solution R

Solution **R** was divided into three equal portions in three test-tubes.
The following tests were carried out.

tests	observations
test 5 The pH of the first portion of solution **R** was measured.	pH = 10
test 6 Drops of aqueous sodium hydroxide were added to the second portion of solution **R** and the test-tube shaken. Excess aqueous sodium hydroxide was then added to the test-tube.	white precipitate no visible change
test 7 Aqueous iron(II) sulfate was added to the third portion of solution **R** and the mixture shaken.	green precipitate formed

(b) Identify solution **R**.

....

.... [2]

[Total: 10]

4 A liquid cleaner is a mixture of three substances. These substances are shown in the table.

name of substance	properties of substance
water	liquid, boiling point 100 °C
sodium carbonate	solid, soluble in water
silica	solid, insoluble in water

Plan an experiment to obtain separate pure samples of each substance from the mixture in the liquid cleaner. You are provided with common laboratory apparatus.

....

....

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....

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....

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....

....

.... [6]

[Total: 6]

A liquid cleaner is a mixture of three substances. These substances are shown in the table.

name of substance	properties of substance
water	liquid, boiling point 100 °C
sodium carbonate	solid, soluble in water
silica	solid, insoluble in water

Plan an experiment to obtain separate pure samples of each substance from the mixture in the liquid cleaner. You are provided with common laboratory apparatus.

....

....

....

....

....

....

....

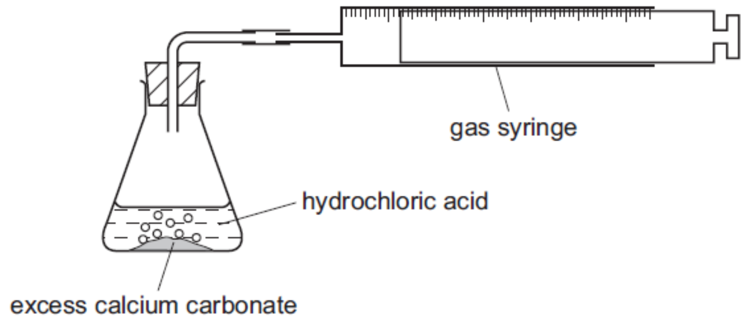
....

....

.... [6]

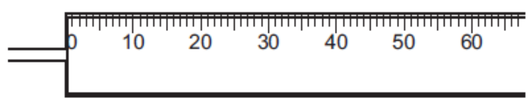
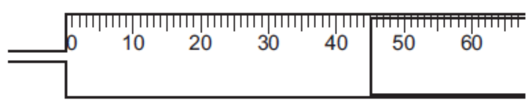
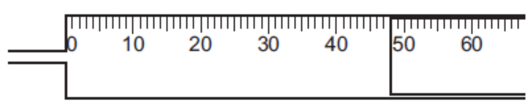
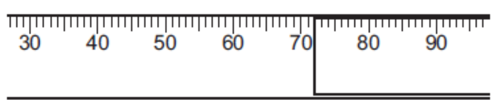
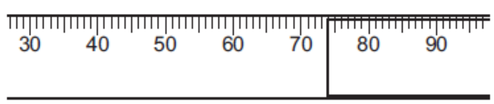
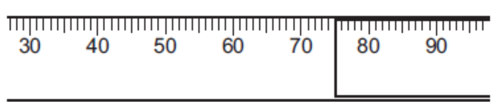
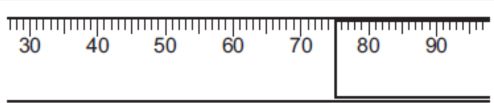
[Total: 6]

The rate of reaction between excess calcium carbonate and dilute hydrochloric acid was investigated using the apparatus shown below. The temperature of the hydrochloric acid was 25 °C.



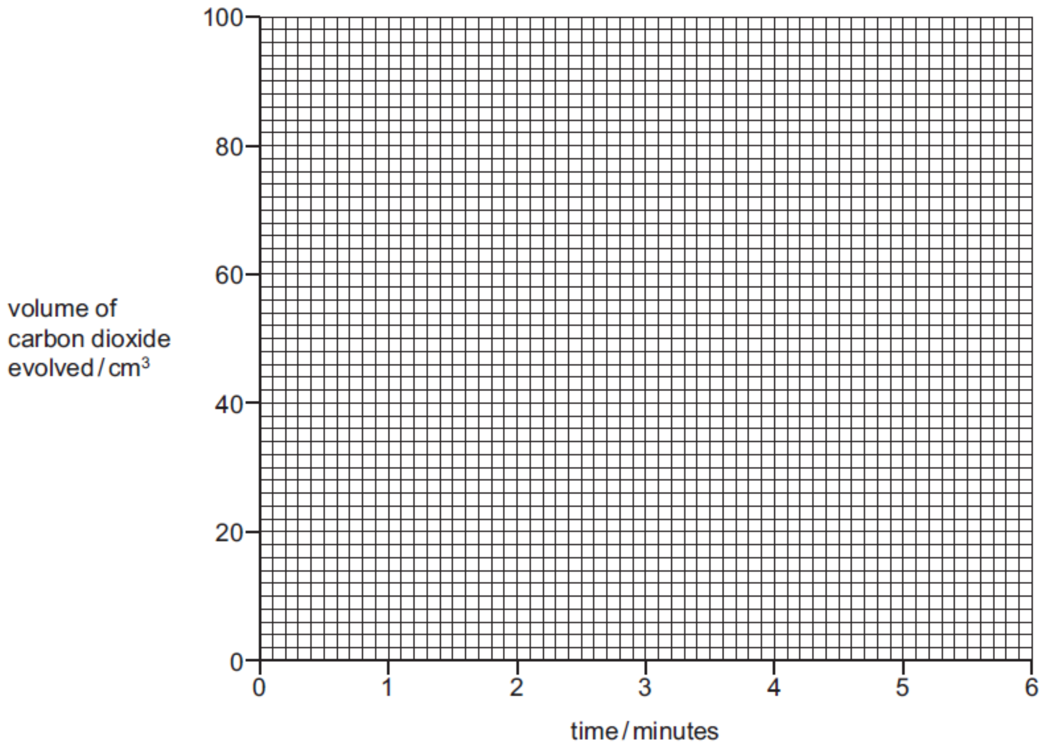
The volume of carbon dioxide evolved was measured every minute for six minutes.

(a) Use the gas syringe diagrams to complete the table of results.

time / minutes	gas syringe diagram	total volume of carbon dioxide evolved / cm^3
0		
1		
2		
3		
4		
5		
6		

[3]

(b) Plot the results on the grid below and draw a smooth line graph.



[4]

(c) **(i)** Which point appears to be inaccurate? Explain why.

....

.... [2]

(ii) **Use your graph** to work out the volume of gas expected at that time. Show clearly **on the grid** how you worked out your answer.

.... [2]

(d) Sketch, **on the grid**, the graph you would expect if the experiment was repeated using hydrochloric acid at a temperature of 50 °C. [2]

[Total: 13]

A student investigated the solubility of salt **D** in water at various temperatures.

Four experiments were carried out.

(a) *Experiment 1*

4 g of salt **D** was added to a boiling tube. A burette was filled with distilled water and 10.0 cm^3 of water added to the boiling tube. The mixture of salt **D** and water was heated carefully until all of the solid had dissolved. The boiling tube was removed from the heat and the solution allowed to cool. The solution was stirred gently with a thermometer.

The temperature at which crystals first appeared was noted.

The boiling tube and its contents were kept for the remaining three experiments.

(b) *Experiment 2*

From the burette another 2.0 cm^3 of water was added to the boiling tube and contents from Experiment 1.

The mixture was heated to dissolve the crystals and allowed to cool as in Experiment 1. The temperature at which crystals first appeared was noted.
Record, in the table, the total volume of water in the boiling tube.

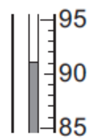
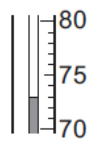
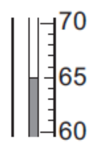
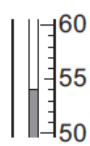
(c) *Experiment 3*

From the burette another 2.0 cm^3 of water was added to the boiling tube and contents from Experiment 2. The experiment was repeated exactly as before.
Record, in the table, the total volume of water in the boiling tube.

(d) *Experiment 4*

From the burette another 4.0 cm³ of water was added to the boiling tube and contents from Experiment 3. The experiment was repeated exactly as before.
Record in the table the total volume of water in the boiling tube.

Use the thermometer diagrams in the table to record the temperatures at which crystals first appeared in the four experiments.

Experiment number	total volume of water / cm³	thermometer diagram	temperature at which crystals first appeared / °C
1	10.0		
2			
3			
4			

[3]

(h) The solubility of salt **D** at 100 °C is 57 g in 100 cm^3 of water.
Suggest, with a reason, the effect of using 8 g of salt **D** instead of 4 g in these experiments.

....

.... [2]

(i) Salt **C** is less soluble in water than salt **D**.

Sketch on the grid the graph you would expect for salt **C**. Label this graph. [2]

(j) Describe and explain **one** improvement that could be made to the experimental method to obtain more reliable results in this investigation.

improvement ....

explanation ....

....

.... [2]

[Total: 18]

Two metal salt solutions, **E** and **F**, were analysed.
E was a mixture of iron(II) sulfate and ammonium sulfate.
The tests on the solutions and some of the observations are in the following table.
Complete the observations in the table.

tests	observations
tests on solution **E** **(a)** Appearance of solution **E**.	.... [1]
The solution was divided into three equal portions in separate test-tubes. **(b)** Dilute nitric acid and aqueous barium nitrate were added to the first portion of the solution.	.... [1]
(c) **(i)** Excess aqueous sodium hydroxide was added to the second portion of the solution. **(ii)** The mixture was filtered and the filtrate heated. The gas given off was tested with damp litmus paper.	.... [2] .... [2]
(d) Dilute sulfuric acid and aqueous potassium manganate(VII), an oxidising agent, were added to the third portion of the solution. Aqueous sodium hydroxide was then added to the mixture.	.... [1]
tests on solution **F** **(e)** Appearance of solution **F**.	yellow liquid
(f) Zinc powder was added to solution **F**. The solution was observed for five minutes. The gas given off was tested with a splint.	rapid effervescence turned blue, then green and finally light purple lighted splint popped

(g) Identify the gas given off in test **(f)**.

.... [1]

(h) What conclusions can you draw about solution **F**?

....

.... [2]

[Total: 10]

Rhubarb Leaves

Ethanedioic acid dihydrate, $H_2C_2O_4.2H_2O$, is a white crystalline solid. This acid is water-soluble and is found in rhubarb leaves.

Plan an investigation to obtain crystals of ethanedioic acid dihydrate from some rhubarb leaves. You are provided with common laboratory apparatus, water and sand.

....

....

....

....

....

....

....

....

.... [7]

[Total: 7]

1 (a) Match the following pH values to the solutions given below.

1 3 7 10 13

The solutions all have the same concentration.

solution	**pH**
aqueous ammonia, a weak base	....
dilute hydrochloric acid, a strong acid	....
aqueous sodium hydroxide, a strong base	....
aqueous sodium chloride, a salt	....
dilute ethanoic acid, a weak acid	....

[5]

(b) Explain why solutions of hydrochloric acid and ethanoic acid with the same concentration, in mol / dm^3, have a different pH.

....

....

.... [2]

(c) Measuring pH is one way of distinguishing between a strong acid and a weak acid.
Describe another method.

method ....

....

results ....

....

[2]

[Total: 9]

3 The main use of sulfur dioxide is the manufacture of sulfuric acid.

(a) State **two** other uses of sulfur dioxide.

....

.... [2]

(b) One source of sulfur dioxide is burning sulfur in air.
Describe how sulfur dioxide can be made from the ore zinc sulfide.

....

.... [2]

(c) The Contact process changes sulfur dioxide into sulfur trioxide.

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

the forward reaction is exothermic

temperature 400 to 450 °C

low pressure 1 to 10 atmospheres

catalyst vanadium(V) oxide

(i) What is the formula of vanadium(V) oxide?

.... [1]

(ii) Vanadium(V) oxide is an efficient catalyst at any temperature in the range 400 to 450 °C.
Scientists are looking for an alternative catalyst which is efficient at 300 °C.
What would be the advantage of using a lower temperature?

....

....

.... [2]

(iii) The process does not use a high pressure because of the extra expense.
Suggest **two** advantages of using a high pressure?
Explain your suggestions.

....

....

....

....

.... [4]

(d) Sulfuric acid is made by dissolving sulfur trioxide in concentrated sulfuric acid to form oleum.
Water is reacted with oleum to form more sulfuric acid.
Why is sulfur trioxide not reacted directly with water?

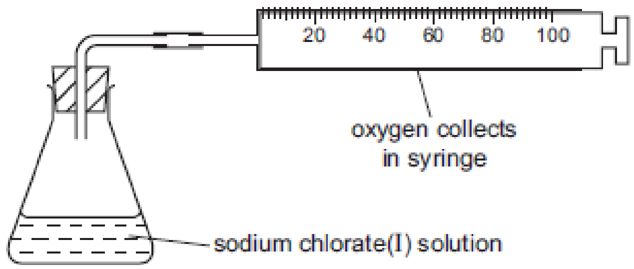
.... [1]

[Total: 12]

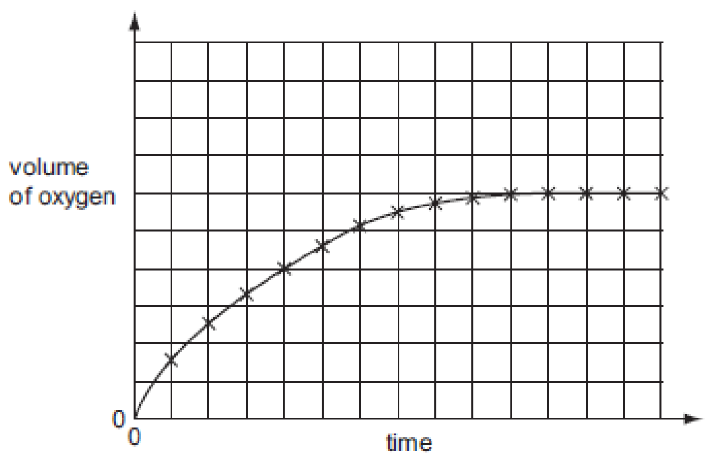
5 **(a)** Sodium chlorate(I) decomposes to form sodium chloride and oxygen. The rate of this reaction is very slow at room temperature provided the sodium chlorate(I) is stored in a dark bottle to prevent exposure to light.

$$2NaClO \rightarrow 2NaCl + O_2$$

The rate of this decomposition can be studied using the following experiment.



Sodium chlorate(I) is placed in the flask and 0.2 g of copper(II) oxide is added. This catalyses the decomposition of the sodium chlorate(I) and the volume of oxygen collected is measured every minute. The results are plotted to give a graph of the type shown below.



(i) Explain why the gradient (slope) of this graph decreases with time.

....

.... [2]

(ii) Cobalt(II) oxide is a more efficient catalyst for this reaction than copper(II) oxide. Sketch, on the grid, the graph for the reaction catalysed by cobalt(II) oxide. All other conditions were kept constant. [2]

(iii) What can you deduce from the comment that sodium chlorate(I) has to be shielded from light?

....

.... [1]

(iv) Explain, in terms of collisions between particles, why the initial gradient would be steeper if the experiment was repeated at a higher temperature.

....

....

....

.... [3]

(b) The ions present in aqueous sodium chloride are Na^+(aq), Cl^-(aq), H^+(aq) and OH^-(aq).

The electrolysis of concentrated aqueous sodium chloride forms three products. They are hydrogen, chlorine and sodium hydroxide.

(i) Explain how these three products are formed. Give ionic equations for the reactions at the electrodes.

....

....

....

....

.... [4]

(ii) If the solution of the electrolyte is stirred, chlorine reacts with sodium hydroxide to form sodium chlorate(I), sodium chloride and water.
Write an equation for this reaction.

Cl_2 + ...NaOH → .... + ... + ...

[2]

[Total: 14]

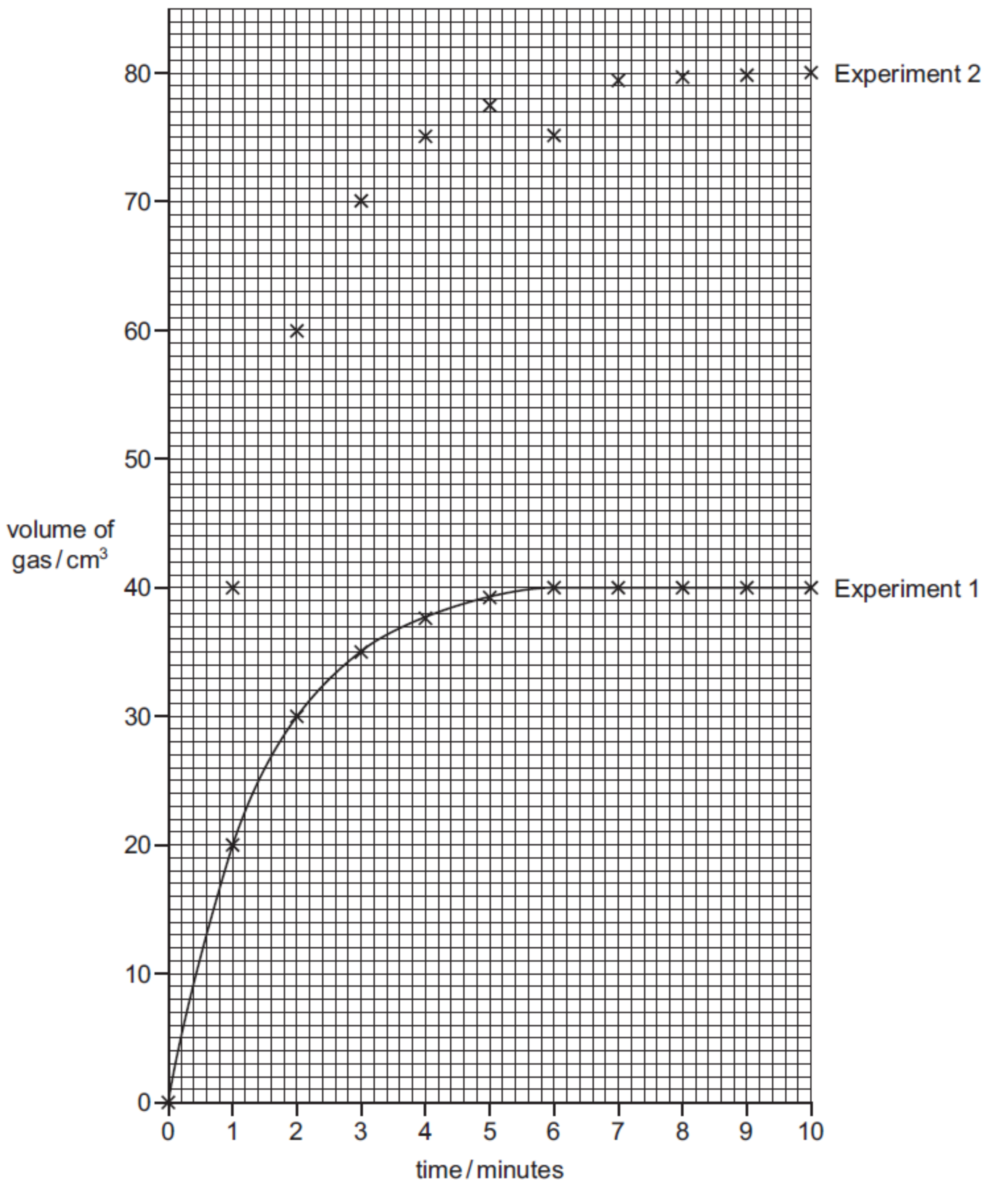
Two experiments were carried out to show what factors affect the rate of decomposition of hydrogen peroxide, H_2O_2.
In each experiment the volume of gas produced was measured every minute for ten minutes.

Experiment 1

The student used a mixture of 50 cm^3 of hydrogen peroxide, 50 cm^3 of water and 1 g of manganese(IV) oxide at a room temperature of 20 °C.
The results were plotted to obtain the graph shown.

Experiment 2

The student repeated Experiment 1 but did not record how much of each substance was used. The points were plotted on the grid.



(a) Complete the graph for Experiment 2. [1]

(b) Suggest the composition of the mixture used in Experiment 2. Explain your suggestion.

composition ....

....

explanation ....

.... [4]

(c) What is the function of the manganese(IV) oxide?

.... [1]

(d) Sketch on the grid the curve that you would expect if Experiment 1 was repeated at 10 °C. [2]

[Total: 8]

5 A solid **D**, which is a soluble metal sulfate, was analysed.
The tests on **D**, and some of the observations, are in the following table.
Complete the observations in the table.

tests	observations
tests on solid **D**	
(a) **(i)** Appearance of solid **D**.	pale green crystals
(ii) Solid **D** was heated in a test-tube gently and then strongly.	condensation formed at the top of the test-tube
tests on the aqueous solution Solid **D** was added to distilled water and shaken to dissolve. The solution was divided into four equal portions in separate test-tubes.	
(b) **(i)** Several drops of aqueous sodium hydroxide were added to the first portion of the solution.	green precipitate
Excess aqueous sodium hydroxide was added to the mixture.	green precipitate remained
(ii) Excess aqueous ammonia was added to the second portion of the solution.	green precipitate
(c) Aqueous silver nitrate and dilute nitric acid were added to the third portion of the solution.	.... [1]
(d) Aqueous barium nitrate and dilute nitric acid were added to the fourth portion of the solution.	.... [2]

(e) What does test **(a)** tell you about solid **D**?

.... [2]

(f) What conclusions can you draw about the identity of solid **D**?

....

.... [3]

[Total: 8]

Paper 2 Questions (7)

1. 2011

The sign $\rightleftharpoons$ is used in some equations to show that a reaction is reversible.

Two incomplete equations are given.

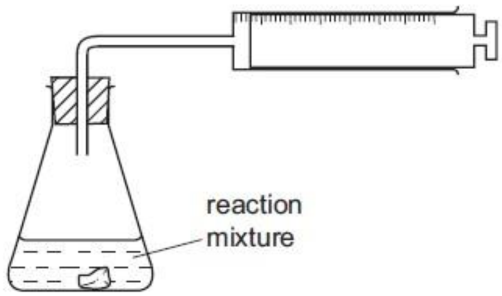
	reactants	products
P	$CoCl_2 + 2H_2O$	$CoCl_2.2H_2O$
Q	$C + O_2$	CO_2

For which of these reactions can a $\rightleftharpoons$ sign be correctly used to complete the equation?

	P	Q
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

2. 2011

An experiment to determine the rate of a chemical reaction could be carried out using the apparatus shown.



Which reaction is being studied?

A $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$

B $Mg + H_2SO_4 \rightarrow MgSO_4 + H_2$

C $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$

D $NaOH + HCl \rightarrow NaCl + H_2O$

3. 2011

Copper(II) carbonate reacts with dilute sulfuric acid.

$$CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + CO_2(g) + H_2O(l)$$

The speed of the reaction can be changed by varying the conditions.

Which conditions would always increase the speed of this chemical reaction?

1 Increase the concentration of the reactants.

2 Increase the size of the pieces of copper(II) carbonate.

3 Increase the temperature.

4 Increase the volume of sulfuric acid.

A 1, 3 and 4 **B** 1 and 3 only **C** 2 and 3 **D** 3 and 4 only

4. 2011 (Also under Unit 6)

Which type of reaction always forms a salt and water?

A exothermic

B neutralisation

C oxidation

D polymerisation

5. 2012

Separate samples of anhydrous and hydrated copper(II) sulfate are heated.

Which shows the correct colour changes?

	anhydrous copper(II) sulfate	hydrated copper(II) sulfate
A	blue to white	white to blue
B	no change	blue to white
C	white to blue	blue to white
D	white to blue	no change

6. 2012

Which change is an oxidation?

A FeO to Fe_2O_3

B Fe_2O_3 to FeO

C H_2O_2 to H_2O

D H_2O to H_2

7. 2012

Which change does **not** increase the speed of reaction between zinc and hydrochloric acid?

A adding a catalyst

B decreasing the particle size of the zinc

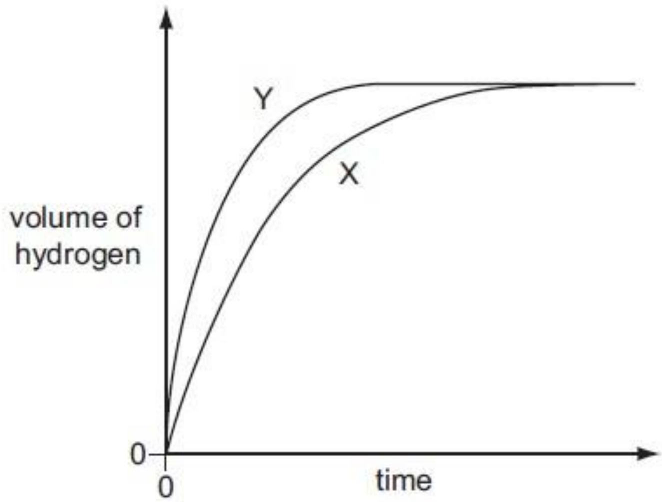
C decreasing the temperature

D using more concentrated acid

8. 2013

A student investigates the rate of reaction between zinc and an excess of sulfuric acid.

The graph shows the results of two experiments, X and Y.



Which change explains the difference between X and Y?

A A catalyst is added in Y.

B A lower temperature is used in Y.

C Larger pieces of zinc are used in Y.

D Less concentrated acid is used in Y.

9. 2013

Anhydrous copper(II) sulfate can be made by heating hydrated copper(II) sulfate.

$$CuSO_4.5H_2O \rightarrow CuSO_4 + 5H_2O$$

What can be added to anhydrous copper(II) sulfate to turn it into hydrated copper(II) sulfate?

A concentrated sulfuric acid

B sodium hydroxide powder

C sulfur dioxide

D water

10. 2013

The reactions shown may occur in the air during a thunder storm.

$$N_2 + O_2 \rightarrow 2NO$$

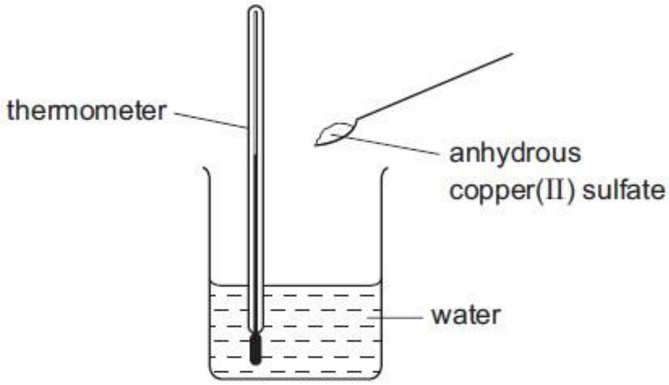
$$2NO + O_2 \rightarrow 2NO_2$$

$$NO + O_3 \rightarrow NO_2 + O_2$$

Which row shows what happens to the reactant molecules in each of these reactions?

	N_2	NO	O_3
A	oxidised	oxidised	oxidised
B	oxidised	oxidised	reduced
C	reduced	reduced	oxidised
D	reduced	reduced	reduced

When anhydrous copper(II) sulfate is added to water a solution is formed and heat is given out.



Which row correctly shows the temperature change and the type of reaction taking place?

	temperature change	type of reaction
A	decreases	endothermic
B	decreases	exothermic
C	increases	endothermic
D	increases	exothermic

Reversible reactions can come to equilibrium. The following are three examples of types of gaseous equilibria.

$A_2(g) + B_2(g) \rightleftharpoons 2AB(g)$ **reaction 1**

$A_2(g) + 3B_2(g) \rightleftharpoons 2AB_3(g)$ **reaction 2**

$2AB_2(g) \rightleftharpoons 2AB(g) + B_2(g)$ **reaction 3**

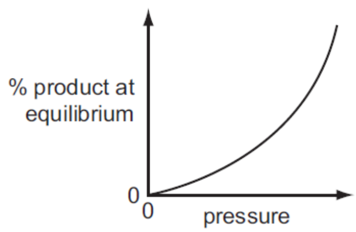
(a) Explain the term *equilibrium*.

....

.... [2]

(b) The following graphs show how the percentage of products of a reversible reaction at equilibrium could vary with pressure.
For each graph, decide whether the percentage of products decreases, increases or stays the same when the pressure is **increased**, then match each graph to one of the above reactions and give a reason for your choice.

(i)



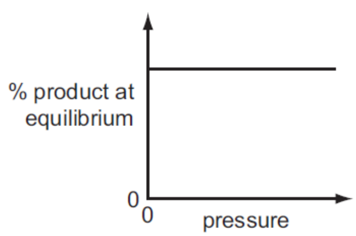
effect on percentage of products ....

reaction ....

reason ....

.... [3]

(ii)



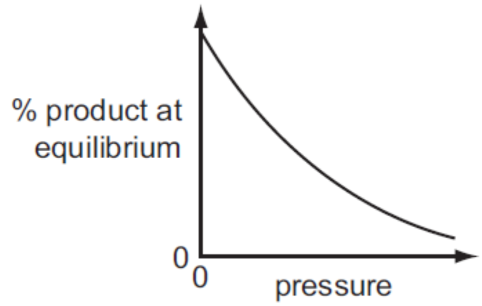
effect on percentage of products ....

reaction ....

reason ....

.... [3]

(iii)



effect on percentage of products ....

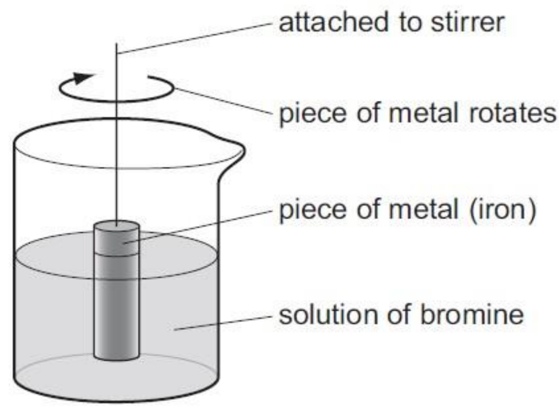
reaction ....

reason ....

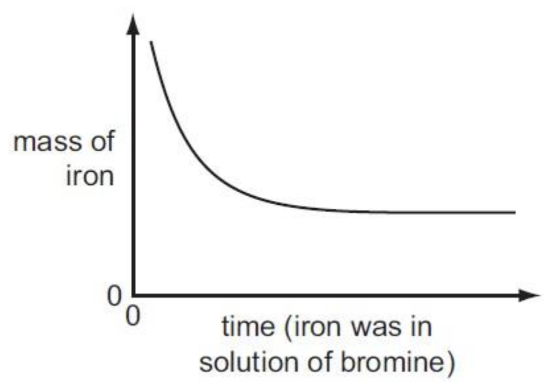
.... [3]

[Total: 11]

The rate of the reaction between iron and aqueous bromine can be investigated using the apparatus shown below.



(a) A piece of iron was weighed and placed in the apparatus. It was removed at regular intervals and the clock was paused. The piece of iron was washed, dried, weighed and replaced. The clock was restarted.
This was continued until the solution was colourless.
The mass of iron was plotted against time. The graph shows the results obtained.



(i) Suggest an explanation for the shape of the graph.

....

....

.... [3]

(ii) Predict the shape of the graph if a similar piece of iron with a much rougher surface had been used.
Explain your answer.

....

.... [2]

(iii) Describe how you could find out if the rate of this reaction depended on the speed of stirring.

....

....

.... [2]

(b) Iron has two oxidation states +2 and +3. There are two possible equations for the redox reaction between iron and bromine.

$$Fe + Br_2 \rightarrow Fe^{2+} + 2Br^-$$

$$2Fe + 3Br_2 \rightarrow 2Fe^{3+} + 6Br^-$$

(i) Indicate, on the first equation, the change which is oxidation. Give a reason for your choice.

....

.... [2]

(ii) Which substance in the first equation is the reductant (reducing agent)?

.... [1]

(c) Describe how you could test the solution to find out which ion, Fe^{2+} or Fe^{3+}, is present.

....

....

.... [3]

The speed (rate) of a chemical reaction depends on a number of factors which include temperature and the presence of a catalyst.

(a) Reaction speed increases as the temperature increases.

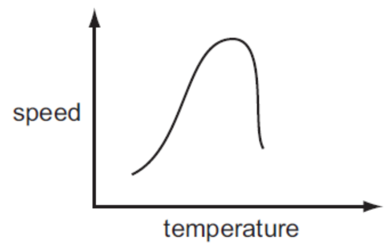
(i) Explain why reaction speed increases with temperature.

....

....

.... [3]

(ii) Reactions involving enzymes do not follow the above pattern.
The following graph shows how the speed of such a reaction varies with temperature.



Suggest an explanation why initially the reaction speed increases then above a certain temperature the speed decreases.

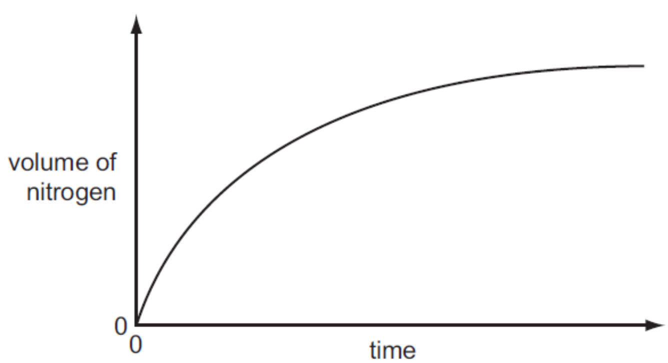
....

.... [2]

(b) An organic compound decomposes to give off nitrogen.

$$C_6H_5N_2Cl(aq) \rightarrow C_6H_5Cl(l) + N_2(g)$$

The speed of this reaction can be determined by measuring the volume of nitrogen formed at regular intervals. Typical results are shown in the graph below.



(i) The reaction is catalysed by copper.
Sketch the graph for the catalysed reaction on the diagram above. [2]

(ii) How does the speed of this reaction vary with time?

.... [1]

(iii) Why does the speed of reaction vary with time?

....

.... [2]

Carbonyl chloride is now made by the reversible reaction given below.

$$CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$$

The forward reaction is exothermic.
The reaction is catalysed by carbon within a temperature range of 50 to 150 °C.

(i) Predict the effect on the yield of carbonyl chloride of increasing the pressure.
Explain your answer.

....

.... [2]

(ii) If the temperature is allowed to increase to above 200 °C, very little carbonyl chloride is formed. Explain why.

....

.... [2]

(iii) Explain why a catalyst is used.

.... [1]

Ammonia is manufactured by the Haber process.

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

The forward reaction is exothermic.

(a) Describe how the reactants are obtained.

(i) Nitrogen

....

.... [2]

(ii) Hydrogen

....

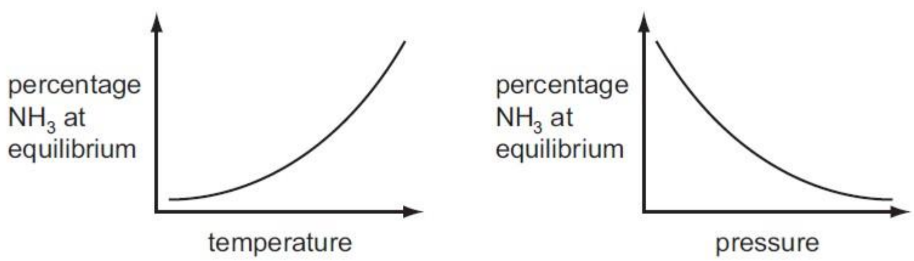
....

.... [3]

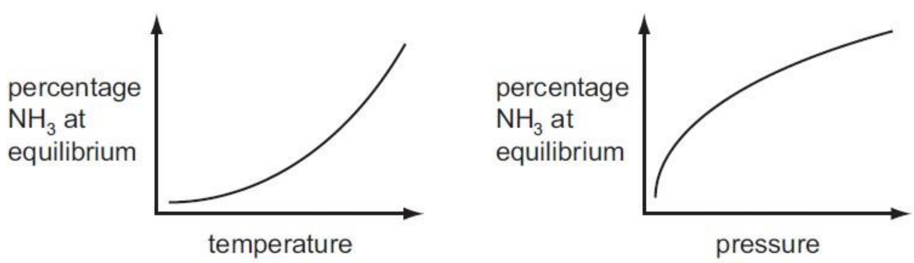
(b) The percentage of ammonia in the equilibrium mixture varies with temperature and pressure.

(i) Which pair of graphs, **A**, **B** or **C**, shows correctly how the percentage of ammonia at equilibrium varies with temperature and pressure?

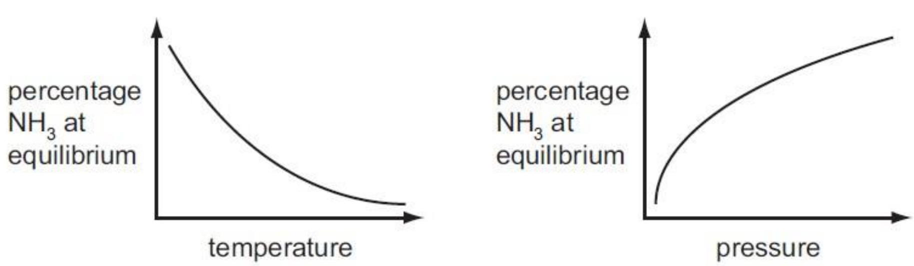
pair **A**



pair **B**



pair **C**



The pair with **both graphs correct** is .... [1]

(ii) Give a full explanation of why the pair of graphs you have chosen in **(i)** is correct.

....

....

....

....

....

....

.... [6]

(iii) Catalysts do not alter the position of equilibrium. Explain why a catalyst is used in this process.

....

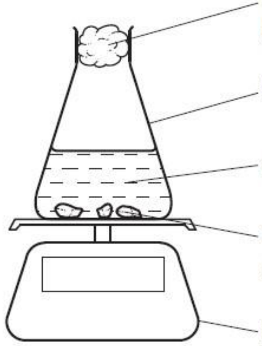
....

....

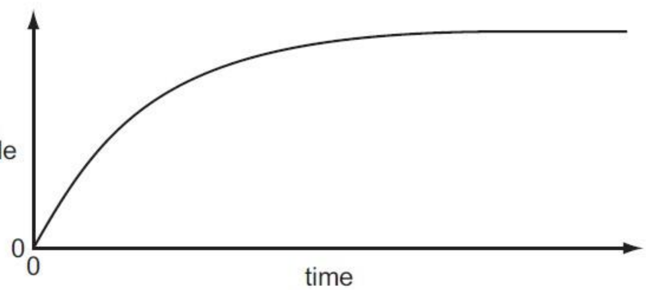
.... [2]

[Total: 14]

20.0 g of small lumps of calcium carbonate and 40 cm^3 of hydrochloric acid, concentration 2.0 mol / dm^3, were placed in a flask on a top pan balance. The mass of the flask and contents was recorded every minute.



The mass of carbon dioxide given off was plotted against time.



$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

In all the experiments mentioned in this question, the calcium carbonate was in excess.

(a) **(i)** Explain how you could determine the mass of carbon dioxide given off in the first five minutes.

.... [1]

(ii) Label the graph **F** where the reaction rate is the fastest, **S** where it is slowing down and **0** where the rate is zero. [2]

(iii) Explain how the shape of the graph shows where the rate is fastest, where it is slowing down and where the rate is zero.

....

....

.... [2]

(b) Sketch on the same graph, the line which would have been obtained if 20.0 g of small lumps of calcium carbonate and 80 cm^3 of hydrochloric acid, concentration 1.0 mol / dm^3, had been used. [2]

(c) Explain in terms of collisions between reacting particles each of the following.

(i) The reaction rate would be slower if 20.0 g of larger lumps of calcium carbonate and 40 cm^3 of hydrochloric acid, concentration 2.0 mol / dm^3, were used.

....

....

.... [2]

(ii) The reaction rate would be faster if the experiment was carried out at a higher temperature.

....

....

.... [2]

(a) The following are two examples of substitution reactions. Only the reaction involving chlorine is a photochemical reaction.

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

$$CH_4 + Br_2 \rightarrow CH_3Br + HBr$$

(i) Explain the phrase *substitution reaction*.

....

.... [1]

(ii) How do photochemical reactions differ from other reactions?

....

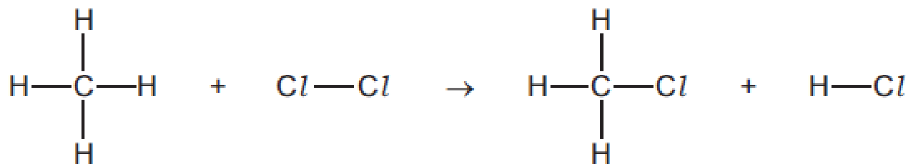
.... [1]

(b) Bond forming is exothermic, bond breaking is endothermic. Explain the difference between an exothermic reaction and an endothermic reaction.

....

.... [2]

(c) Use the bond energies to show that the following reaction is exothermic.
Bond energy is the amount of energy (kJ / mol) which must be supplied to break one mole of the bond.



Bond energies in kJ / mol

C*l*–C*l*	+242
C–C*l*	+338
C–H	+412
H–C*l*	+431

bonds broken .... energy in kJ / mol

....

....

total energy = ...

bonds formed .... energy in kJ / mol

....

....

total energy = ...

....

.... [4]

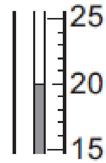
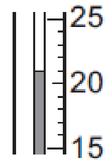
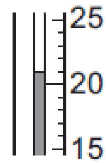
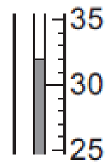
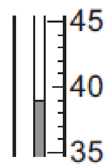
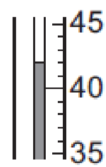
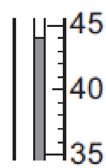
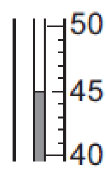
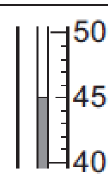
[Total: 8]

A student investigated the reaction between aqueous copper(II) sulfate and two different metals, zinc and iron.
Two experiments were carried out.

Experiment 1
Using a measuring cylinder, 25 cm^3 of aqueous copper(II) sulfate was poured into a polystyrene cup. The temperature of the solution was measured. The timer was started and the temperature was measured every half a minute for one minute.

At 1 minute, 5 g of zinc powder was added to the cup and the mixture stirred with the thermometer. The temperature of the mixture was measured every half minute for an additional three minutes.

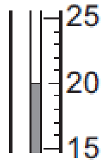
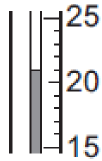
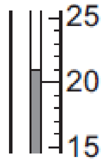
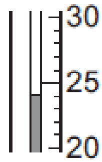
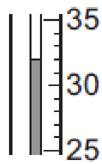
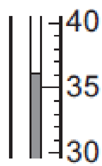
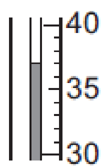
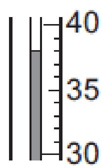
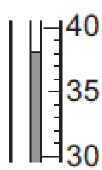
(a) Use the thermometer diagrams in the table to record the temperatures.

time / min	thermometer diagrams	temperature / °C
0.0		
0.5		
1.0		
1.5		
2.0		
2.5		
3.0		
3.5		
4.0		

[3]

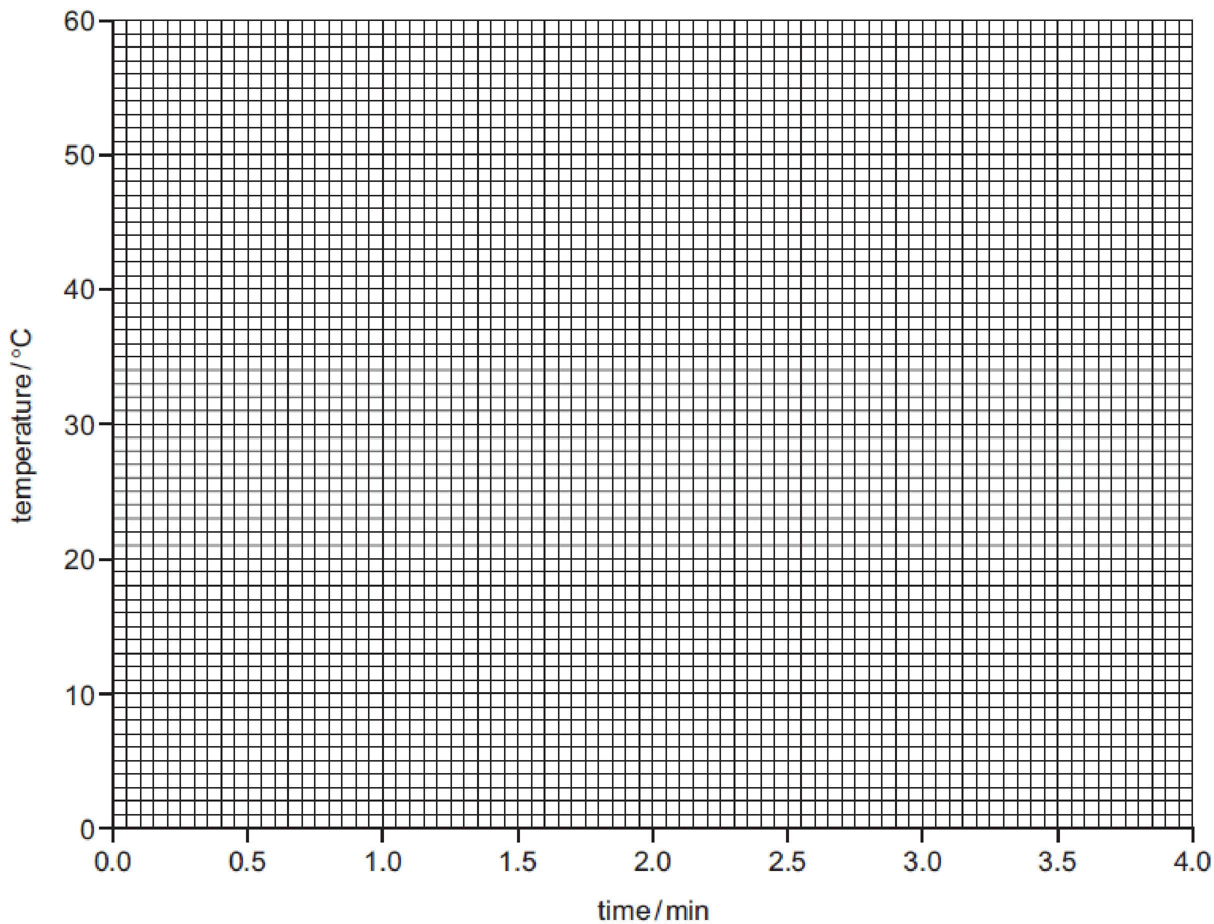
Experiment 2
Experiment 1 was repeated using 5 g of iron powder instead of the zinc powder.

(b) Use the thermometer diagrams in the table to record the temperatures.

time / min	thermometer diagrams	temperature / °C
0.0		
0.5		
1.0		
1.5		
2.0		
2.5		
3.0		
3.5		
4.0		

[3]

(c) Plot the results of both experiments on the grid below. Draw two smooth line graphs. Clearly label your graphs.



[5]

(d) **From your graph**, work out the temperature of the reaction mixture in Experiment 1 after 1 minute 15 seconds. Show clearly **on the graph** how you worked out your answer.

.... [3]

(e) What type of chemical process occurs when zinc and iron react with aqueous copper(II) sulfate?

.... [1]

(f) (i) Compare the temperature changes in Experiments 1 and 2.

.... [1]

(ii) Suggest an explanation for the difference in temperature changes.

....

.... [1]

(g) Explain how the temperature changes would differ in the experiments if 12.5 cm^3 of copper(II) sulfate solution were used.

....

....

.... [2]

(h) Predict the effect of using lumps of zinc in Experiment 1. Explain your answer.

....

.... [2]

[Total: 21]

A student investigated the speed of reaction when iodine was produced by the reaction of solution **L** with potassium iodide at different temperatures.

Five experiments were carried out.

Experiment 1

A burette was filled with the aqueous solution **L** to the 0.0 cm^3 mark.
10.0 cm^3 of solution **L** was added from the burette into a boiling tube and the initial temperature of the solution was measured.

Using a measuring cylinder, 5 cm^3 of aqueous potassium iodide and 3 cm^3 of aqueous sodium thiosulfate were poured into a second boiling tube. Starch solution was added to this boiling tube and the mixture shaken.

The mixture in the second boiling tube was added to the solution **L**, shaken and the clock started. These chemicals reacted to form iodine which reacted with the starch. When a blue colour appeared, the clock was stopped and the time measured and recorded in the table. The final temperature of the mixture was measured.

Experiment 2

Experiment 1 was repeated but solution **L** was heated to about 40 °C. The temperature of the solution was measured before adding the mixture in the second boiling tube.
When a blue colour appeared, the clock was stopped and the time measured and recorded in the table. The final temperature of the mixture was measured.

Experiment 3

Experiment 2 was repeated, heating solution **L** to about 50 °C.

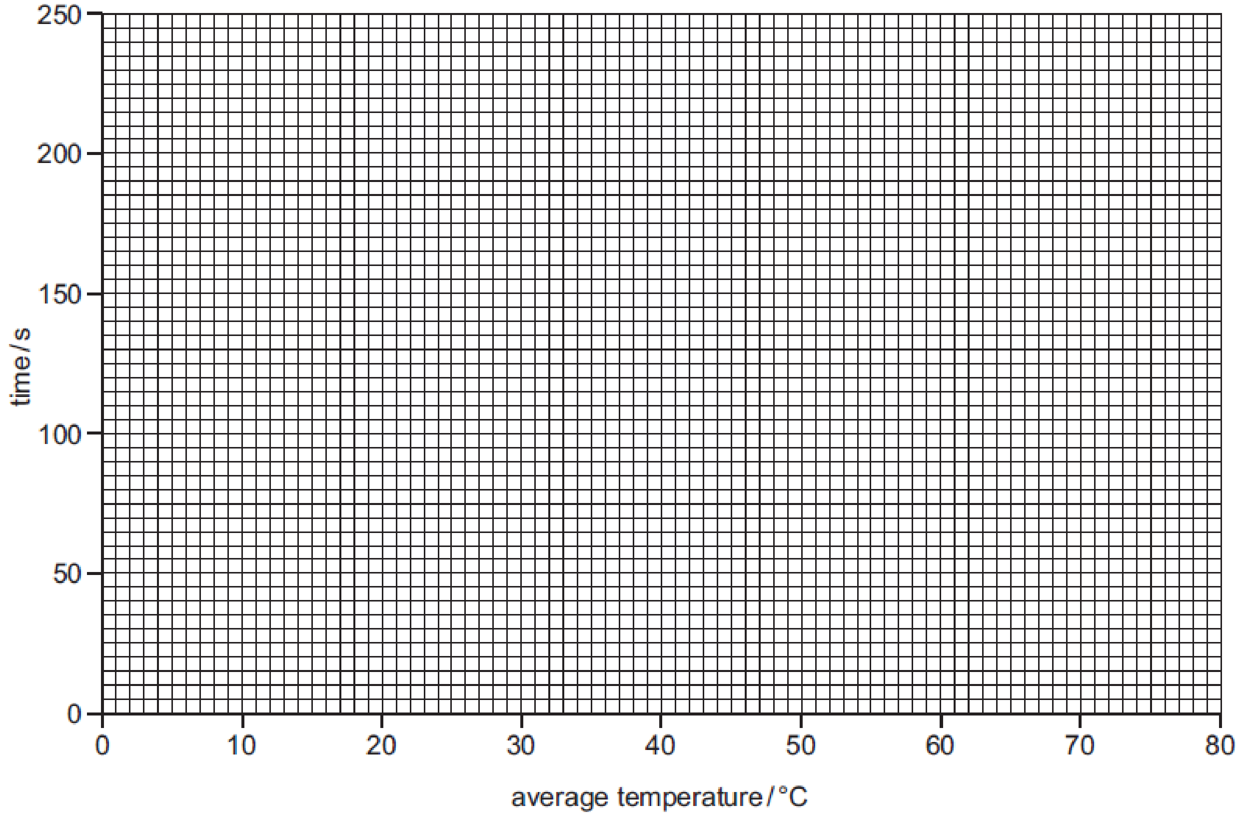
Experiment 4

Experiment 2 was repeated, heating solution **L** to about 60 °C.

Experiment 5

Experiment 2 was repeated, heating solution **L** to about 70 °C.

(b) Plot the results on the grid below and draw a smooth line graph.



[5]

(c) **From your graph**, work out the time taken for the blue colour to appear if solution **L** was heated to 80 °C. The final temperature of the reaction mixture was 64 °C.
Show clearly **on the grid** how you obtained your answer.

.... [2]

(d) Suggest the purpose of the starch solution in the experiments.

.... [1]

(e) **(i)** In which experiment was the reaction speed fastest?

.... [1]

(ii) Explain, using ideas about particles, why this experiment was the fastest.

....

.... [2]

(f) Predict the effect on the time and speed of the reaction in Experiment 5 if it was repeated using a less concentrated solution of **L**.

time ....

speed .... [2]

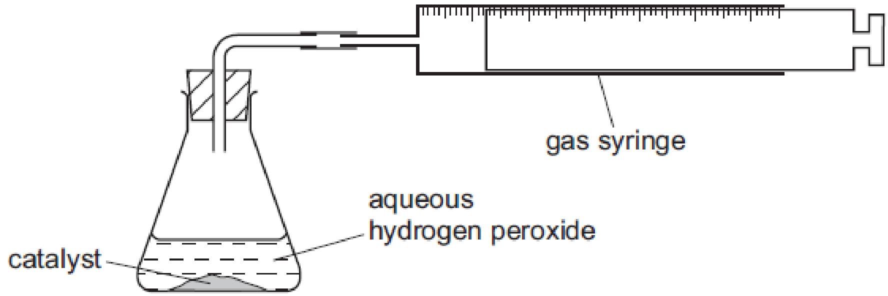
(g) Why was a burette used to measure solution **L** instead of a measuring cylinder?

....

.... [1]

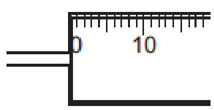
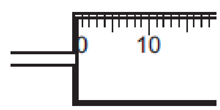
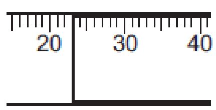
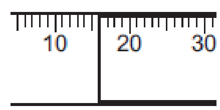
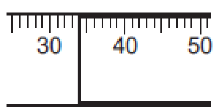
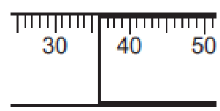
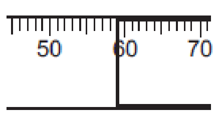
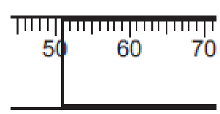
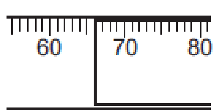
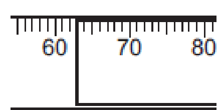
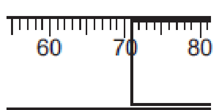
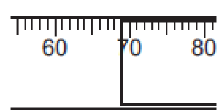
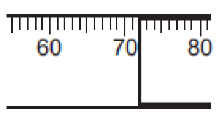
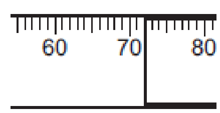
[Total: 19]

Two experiments using catalysts were carried out. Catalysts **R** and **S** were used to break down 50 cm³ of aqueous hydrogen peroxide at a temperature of 20 °C. The volume of oxygen given off was measured using the apparatus shown.



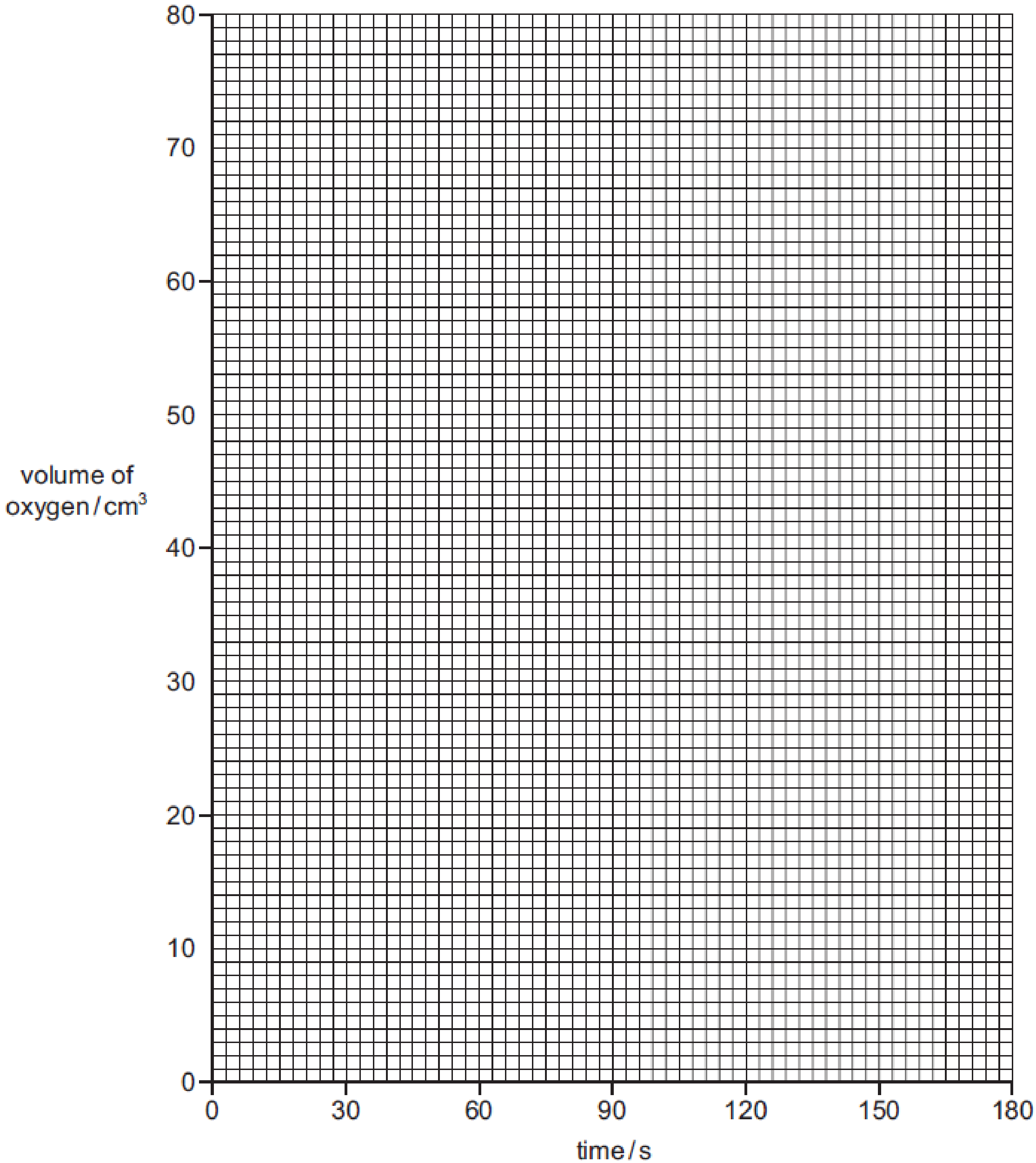
The gas syringe diagrams show the volume of oxygen formed every 30 seconds in each experiment.

(a) Use the syringe diagrams to complete the volumes in the table.

time / s	using catalyst **R**		using catalyst **S**	
	syringe diagram	volume / cm³	syringe diagram	volume / cm³
0				
30				
60				
90				
120				
150				
180				

[4]

(b) Plot a graph to show each set of results. Clearly label the graphs **R** and **S**.



[6]

(c) Which result using catalyst **R** was inaccurate?

.... [1]

(d) Which is the better catalyst in this reaction? Explain your answer.

....

.... [2]

(e) Sketch a line on the grid to show the graph you would expect if the reaction with catalyst **R** was repeated at 50 °C. [2]

[Total: 15]

CHAPTER 8 –

Acids, bases and salts

SUB TOPICS -

The characteristic properties of acids and bases

- Describe the characteristic properties of acids as reactions with metals, bases, carbonates and effect on litmus
- Describe the characteristic properties of bases as reactions with acids and with ammonium salts and effect on litmus
- Describe neutrality and relative acidity and alkalinity in terms of pH (whole numbers only) measured using Universal Indicator paper
- Describe and explain the importance of controlling acidity in soil
- **Define *acids* and *bases* in terms of proton transfer, limited to aqueous solutions**
- **Describe the meaning of weak and strong acids and bases**

Types of oxides

- Classify oxides as either acidic or basic, related to metallic and non-metallic character
- **Further classify other oxides as neutral or amphoteric**

Preparation of salts

- Describe the preparation, separation and purification of salts as examples of some of the techniques specified in section 2.2(b) and the reactions specified in section 8.1
- **Describe the preparation of insoluble salts by precipitation**
- **Suggest a method of making a given salt from suitable starting material, given appropriate information**

Identification of ions and gases

- Describe the following tests to identify:
  aqueous cations:

  aluminium, ammonium, calcium, copper(II), iron(II), iron(III) and zinc (using aqueous sodium hydroxide and aqueous ammonia as appropriate) (Formulae of complex ions are **not** required.)

  anions:

  carbonate (by reaction with dilute acid and then limewater), chloride (by reaction under acidic conditions with aqueous silver nitrate), iodide (by reaction under acidic conditions with aqueous silver nitrate), nitrate (by reduction with aluminium), sulfate (by reaction under acidic conditions with aqueous barium ions)

  gases:

  ammonia (using damp red litmus paper), carbon dioxide (using limewater), chlorine (using damp litmus paper), hydrogen (using lighted splint), oxygen (using a glowing splint).

Acids and Bases

Acid – a substance that acts as a donor of hydrogen ions

Base – a substance that acts as an acceptor of hydrogen ions

Alkali – soluble bases

Acid		Base	
Sour Taste		Bitter Taste	
pH less than 7		pH greater than 7	
In solution, contains hydronium ions (H_3O^+)		In solution, contains hydroxide ions (OH^-)	
Turns blue litmus red		Turns red litmus blue	
Turns phenolphthalein colourless		Turns colourless phenolphthalein pink	
Corrosive		Soapy feel	
Reacts with metals to produce salt and hydrogen		Cannot react with metals	
Examples of Acids		***Examples of Bases***	
Hydrochloric Acid	*HCl*	*Sodium Hydroxide*	*NaOH*
Nitric Acid	HNO_3	*Potassium Hydroxide*	*KOH*
Sulphuric Acid	H_2SO_4	*Calcium Hydroxide*	$Ca(OH)_2$
Ethanoic Acid	CH_3COOH	*Ammonia Solution*	$NH_3\,(aq)$

Hydronium Ion – same as a single proton because when a hydrogen atom loses an electron, only a proton remains. H^+ is irresistibly attractive to water molecules and therefore it would form H_3O^+.

Dissociation – breaking apart

Strong Acids – in aqueous solutions, strong acids donate all their protons to water molecules.

Weak Acids – there is only a slight tendency to donate protons to water molecules, therefore an aqueous solution of a weak acid contains mainly undissociated molecules and a low concentration of H_3O^+.

	Strong Acids	Weak Acids
Dissociation in Aqueous Solution	Completely dissociate	Partially dissociate
Equilibrium	None (forward only)	Equilibrium reaction
Electrolyte	Good	Poor
Electrical conductivity	Good	Poor
$[H_3O^+]$	Higher	Lower
pH value	Lower	Higher
Examples	HCl, HNO_3, H_2SO_4	CH_3COOH, NH_4^+

Amphiprotic – substances can act as both an acid and a base *e.g.* H_2O, HCO_3^-, HSO_4^-

Amphoteric – substances will undergo chemical reactions with both acids and bases

Neutralisation – an alkali or base can neutralise an acid by removing the H^+ ions and converting them to water. Neutralisation always produces a salt.

Concentration – a measure of the amount of acid per dm^3, refers to the proportion or ratio of acid to water in the solution

Concentrated Acids – high proportion of acid to water

Dilute Acids – low proportion of acid to water

Monoprotic – having one transferrable proton

Diprotic – having two transferrable protons

Titration – an indicator shows when the acid properties are just destroyed by the alkali. The salt can then be recovered by evaporating the water away allowing the salt to crystallise. This method is used when the base, acid and salt are all soluble.

Oxides

- Oxides of metals are bases (they will react with acids to form salts)
- Oxides of non-metals are acids (they will react with acids and bases)
- Some metal oxides are amphoteric (they will react with acids and bases)
- Some non-metal oxides are neutral
- Oxide ions immediately react with water and then dissolve to form hydroxide ions. Although potassium hydroxide solution exists, potassium oxide solution does not exist

Metal Oxides – compounds of metal cations and the oxide anion O^{2-}. Few metal oxides react or dissolve in water. The main metal oxides which are considered soluble are potassium and sodium oxides, as well as barium, calcium, and magnesium oxides in decreasing amounts. Metal oxides are either basic or amphoteric. The basic oxides will only react with acids, while the amphoteric oxides will react with both acids and bases.

Non-metal Oxides – covalently bonded compounds of a non-metal with oxygen. They are either acidic or neutral oxides. The acidic oxides react with water immediately and dissociate to form acid solutions while the neutral oxides do nothing when placed in water. The acidic oxides will react only with bases, while the neutral oxides are unreactive with both acids and bases.

Solubility of Ionic Compounds in Water

Precipitate – a solid formed in a solution.

Sparingly Soluble (SpSol) – materials have very low solubilities

Hydrolyse (Hyd) – reacts with water

Always Soluble	Usually Soluble	Usually Insoluble
All NO_3^-	All SO_4^{2-} EXCEPT Ba, Pb, Ag, Ca	All CO_3^{2-} EXCEPT Group 1
All NH_4^+	All Cl^- EXCEPT Ag, Pb	All O^{2-} EXCEPT Group 1 and 2
Group 1	All I^- EXCEPT Ag, Pb	All OH^- EXCEPT Group 1; Ca and Ba are slightly soluble
	All Br^- EXCEPT Ag, Pb	

CHAPTER 9 –

The Periodic Table

SUB TOPICS -

Describe the Periodic Table as a method of classifying elements and its use to predict properties of elements

Periodic trends

- Describe the change from metallic to non-metallic character across a period
- **Describe the relationship between Group number, number of valency electrons and metallic/non-metallic character**

Group properties

- Describe lithium, sodium and potassium in Group I as a collection of relatively soft metals showing a trend in melting point, density and reaction with water
- Predict the properties of other elements in Group I, given data, where appropriate
- Describe chlorine, bromine and iodine in Group VII as a collection of diatomic non-metals showing a trend in colour, and state their reaction with other halide ions
- Predict the properties of other elements in Group VII, given data where appropriate
- **Identify trends in other Groups, given information about the elements concerned**

Transition elements

- Describe the transition elements as a collection of metals having high densities, high melting points and forming coloured compounds, and which, as elements and compounds, often act as catalysts

Noble gases

- Describe the noble gases as being unreactive
- Describe the uses of the noble gases in providing an inert atmosphere, i.e. argon in lamps, helium for filling balloons

Physical Properties of Metals, Non-Metals and Metalloids

Metals	Non-Metals	Metalloids
Lustre (shiny)	No lustre	Can be shiny or dull
Good conductor of heat	Poor conductor of heat	Fair conductor of heat
Good conductor of electricity	Poor conductor of electricity	Fair conductor of electricity
Malleable	Not Malleable	Malleable
Ductile	Not Ductile	Ductile
High Density	Low Density	Solids
High Melting Point	Low Melting Point	

Chemical Properties of Metals and Non-Metals

Metals	Non-Metals
Easily loses electrons	Tends to gain electrons
Oxides generally basic and amphoteric	Oxides generally neutral
Corrodes easily	

Alkali Metals

- Group One Metals
- Very low density and therefore floats on water. The densities increase down the group.
- Silvery and shiny when freshly cut, however they quickly tarnish
- Low melting point
- Low boiling point
- The reactivity increases down the group. Since the valence electron is further from the nucleus, the attractive force holding it is weaker and therefore other stronger forces can easily remove it.

Transition Metals

Physical Properties (compared to Group 1)	Chemical Properties (compared to Group 1)
Much harder	Much less reactive
Higher tensile strength	Many have excellent corrosion resistance
Higher density	Show more than one valency (*e.g.* Fe^{2+} *or* Fe^{3+})
Higher melting point and boiling point	Them and their compounds are useful catalysts
Many of their compounds are coloured	Some are strongly magnetic

Alloys

Alloy	Mixture	Use
Solder	70% Tin, 30% Lead	Joining wires and pipes
Brass	60-95% Copper, 5-40% Zinc	Taps, hose/pipe fittings, zips, screws
Bronze	90% Copper, 10% Tin	Ornaments, bells, bearings
Mild Steel	99.5% Iron, 0.5% Carbon	General structural purposes, cars
Hard Steel	99% Iron, 1% Carbon	Blades
Stainless Steel	74% Iron, 18% Chromium, 8% Nickel	Corrosion resistance
Alnico	Iron/Aluminium/Nickel/Cobalt	Permanent Magnets

Paper 2 Questions (9)

1.2011

Statement 1: Helium is a reactive gas.

Statement 2: Helium can be used to fill balloons.

Which is correct?

A Both statements are correct and statement 2 explains statement 1.

B Both statements are correct but statement 2 does not explain statement 1.

C Statement 1 is correct but statement 2 is incorrect.

D Statement 2 is correct but statement 1 is incorrect.

2. 2011

An element has the following properties.

- It forms coloured compounds.
- It acts as a catalyst.
- It melts at 1539 °C.

In which part of the Periodic Table is the element found?

A Group I

B Group IV

C Group VII

D transition elements

3. 2011

The table shows some properties of two elements in Group VII of the Periodic Table.

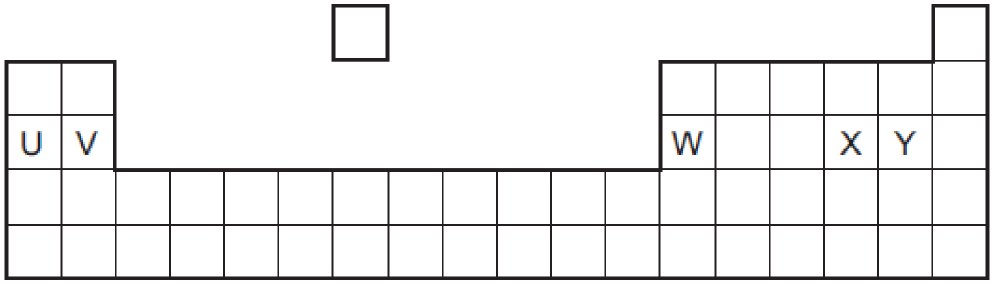
element	state at 20 °C	density / g per cm^3	melting point / °C
chlorine	gas	0.0032	–101
bromine	liquid	3.1	–7

Which properties is fluorine likely to have?

	state at 20 °C	density / g per cm^3	melting point / °C
A	gas	0.0017	–220
B	gas	0.17	–188
C	liquid	0.0017	–220
D	liquid	0.17	–188

4. 2012

The diagram shows an outline of the Periodic Table.



Which of the elements U, V, W, X and Y would react together in the ratio of 1 : 1?

A U and X **B** U and Y **C** V and Y **D** W and X

5. 2012

The element rubidium, Rb, is immediately below potassium in the Periodic Table.

It reacts with bromine to form the compound rubidium bromide.

Which descriptions of this compound are correct?

	type of bond	formula	colour
A	covalent	RbBr	brown
B	covalent	$RbBr_2$	white
C	ionic	RbBr	white
D	ionic	$RbBr_2$	brown

6. 2012

The table gives information about four elements.

Which element is a transition metal?

	colour of element	electrical conductivity of element	colour of oxide
A	black	high	colourless
B	colourless	low	white
C	grey	high	red
D	yellow	low	colourless

7. 2012

Why are weather balloons filled with helium rather than hydrogen?

A Helium is found in air.

B Helium is less dense than hydrogen.

C Helium is more dense than hydrogen.

D Helium is unreactive.

8. 2013

Rubidium is in Group I of the Periodic Table and bromine is in Group VII.

Rubidium reacts with bromine to form an ionic compound.

Which row shows the electron change taking place for rubidium and the correct formula of the rubidium ion?

	electron change	formula of ion formed
A	electron gained	Rb^+
B	electron gained	Rb^-
C	electron lost	Rb^+
D	electron lost	Rb^-

9. 2013

Calcium, on the left of Period 4 of the Periodic Table, is more metallic than bromine on the right of this period.

Why is this?

Calcium has

A fewer electrons.

B fewer protons.

C fewer full shells of electrons.

D fewer outer shell electrons.

10. 2013

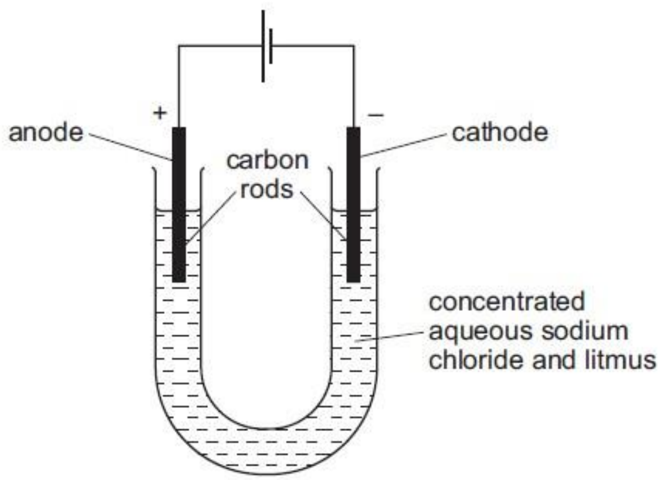
The diagrams show the labels of four bottles.

Which label is **not** correct?

A	B	C	D
Bromine Br_2 Harmful liquid. Do not spill.	Iodine I_2 Danger Avoid breathing vapour from the solid.	Potassium K Danger Store under water.	Sodium Na Danger Store under oil.

11. 2013 (Also under Unit 5)

The diagram shows the electrolysis of concentrated aqueous sodium chloride.



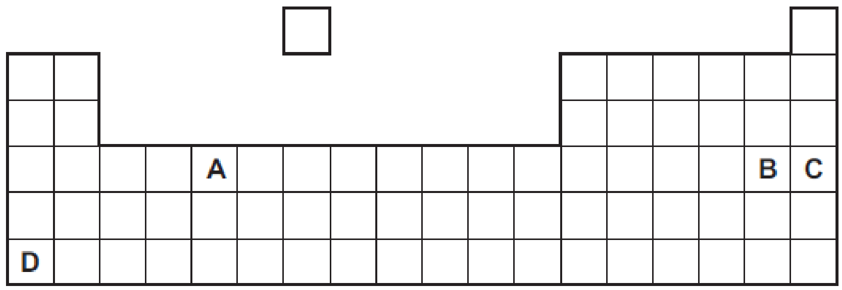
What is the colour of the litmus at each electrode after five minutes?

	colour at anode	colour at cathode
A	blue	red
B	red	blue
C	red	colourless
D	colourless	blue

12. 2013

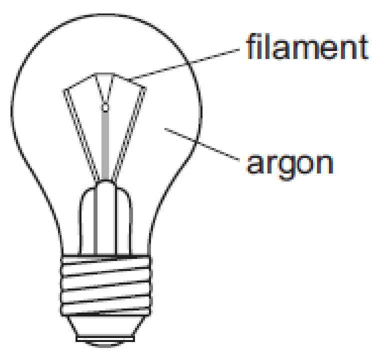
An element has a melting point of 1084 °C and a density of 8.93 g / cm^3. It's oxide can be used as a catalyst.

In which position in the Periodic Table is the element found?



13. 2013

The diagram shows a light bulb.



Why is argon used instead of air in the light bulb?

A Argon is a good conductor of electricity.

B Argon is more reactive than air.

C The filament glows more brightly.

D The filament does not react with the argon.

14. 2013

Some properties of four elements W, X, Y and Z are listed.

1 W melts at 1410 °C and forms an acidic oxide.

2 X has a high density and is easily drawn into wires.

3 Y acts as a catalyst and its oxide reacts with acids.

4 Z is a red-brown solid used to make alloys.

Which of the elements are metals?

A 1 and 3 **B** 2, 3 and 4 **C** 2 and 3 only **D** 2 and 4 only

CHAPTER 9 –

The Periodic Table

Paper 4 Questions (9)

1. 2012 Question 2 (a)

Three of the halogens in Group VII are listed below.

chlorine
bromine
iodine

(a) **(i)** How does their colour change down the Group?

.... [1]

(ii) How do their melting points and boiling points change down the Group?

.... [1]

(iii) Predict the colour and physical state (solid, liquid or gas) of astatine, At.

colour ....

physical state .... [2]

For each of the following, name an element which matches the description.

(a) It is used as a fuel in nuclear reactors.

.... [1]

(b) It is the only non-metal which is a good conductor of electricity.

.... [1]

(c) Inert electrodes are made from this metal.

.... [1]

(d) This gaseous element is used to fill balloons in preference to hydrogen.

.... [1]

(e) An element which can form an ion of the type X^{3-}.

.... [1]

(f) It has the same electron distribution as the calcium ion, Ca^{2+}.

.... [1]

(g) The element is in Period 5 and Group VI.

.... [1]

[Total: 7]

Give **three** differences in physical properties between the Group I metal, potassium, and the transition element, iron.

1. ....

2. ....

3. .... [3]

CHAPTER 10 -

Metals

SUB TOPICS -

Properties of metals

- Describe the general physical and chemical properties of metals
- Explain why metals are often used in the form of alloys
- Identify representations of alloys from diagrams of structure

Reactivity series

- Place in order of reactivity: potassium, sodium, calcium, magnesium, zinc, iron, (hydrogen) and copper, by reference to the reactions, if any, of the metals with
    - water or steam
    - dilute hydrochloric acid

  and the reduction of their oxides with carbon
- Deduce an order of reactivity from a given set of experimental results
- **Describe the reactivity series as related to the tendency of a metal to form its positive ion, illustrated by its reaction, if any, with**
    - **the aqueous ions**
    - **the oxides**

  of the other listed metals
- **Describe the action of heat on the hydroxides and nitrates of the listed metals**
- **Account for the apparent unreactivity of aluminium in terms of the oxide layer which adheres to the metal**

Extraction of metals

- Describe the ease in obtaining metals from their ores by relating the elements to the reactivity series
- Describe the essential reactions in the extraction of iron from hematite
- Describe the conversion of iron into steel using basic oxides and oxygen
- **Describe in outline, the extraction of zinc from zinc blende**
- **Name the main ore of aluminium as bauxite**

Uses of metals

- Name the uses of aluminium:
    - in the manufacture of aircraft because of its strength and low density
    - in food containers because of its resistance to corrosion
- Describe the idea of changing the properties of iron by the controlled use of additives to form steel alloys
- Name the uses of mild steel (car bodies and machinery) and stainless steel (chemical plant and cutlery)
- **Name the uses of zinc for galvanising and for making brass**
- **Name the uses of copper related to its properties (electrical wiring and in cooking utensils)**

Metals Uses

Metal	Property	Uses
Aluminium	Does not corrode	Food containers
	Low density, unreactive	Containers and packaging buildings
	Low density, strong, conducts	Long distance wiring
	Low density, strong, cheap	Transport vehicles
	Low density, conducts heat	Car Engines
Zinc	Reactive	Dry cells ("batteries")
	More reactive than iron	Galvanising Iron
Iron	Similar expansivity	Reinforcing concrete
	Strong, cheap	Nails
	Strong and abundant	Ship building
Copper	Good conductor of electricity	Electrical wiring
	Unreactive, workable	Alloys – brass and bronze
	Unreactive	Coinage (with Nickel)
	Unreactive	Hot water piping

Reactivity Series

Most Reactive

K	Potassium
Na	Sodium
Ca	Calcium
Mg	Magnesium
Al	Aluminium
C	Carbon
Zn	Zinc
Fe	Iron
Sn	Tin
Pb	Lead
H	Hydrogen
Cu	Copper
Ag	Silver
Au	Gold
Pt	Platinum

Least Reactive

Any metal higher on the reactivity series will displace another lower metal's ions from solution.

e.g. $Ca\,(s) + Cu^{2+}\,(aq) \rightarrow Ca^{2+}\,(aq) + Cu\,(s)$

BUT $Cu\,(s) + Ca^{2+}\,(aq) \rightarrow$ No Reaction

The more reactive metals are difficult to extract from their ores in compound form as they are stable.

The less reactive metals have the greater tendency to form atoms and therefore their compounds are less stable.

Corrosion

Corrosion – when metals react with water and oxygen. The metal ions lose electrons to form ions.

Rusting – the corrosion of iron metal to form a red-brown compound *(hydrated iron (III) oxide)*

Covering with Protective Coat	Preventing Oxidation of Metal
Painting	Galvanising – zinc atoms react before the iron
Greasing of metal parts	Sacrificial protection – a more reactive metal reacts before the metal that it is protecting
Oiling of bike chains	
Tin Plating – In cans	Carthodic protection – an electric power source pushes electrons into the metal to prevent the loss of electrons
Plastic covering on electric wires	
Galvanising – zinc coating for galvanised steel	
Chromium plating of car parts	

Reduction of Metal Oxides

By Hydrogen: only the metals below Hydrogen in the reactivity series are reduced by using this method (mainly only CuO)

$$CuO\ (s) + H_2\ (g) \rightarrow Cu\ (s) + H_2O\ (l)$$

By Carbon: only the metals below Carbon in the reactivity series are reduced by using this method

$$2\ PbO\ (s) + C\ (s) \rightarrow 2\ Pb\ (l) + CO_2\ (g)$$

By Carbon Monoxide: only metals below Carbon in the reactivity series are reduced by using this method

$$CuO\ (s) + CO\ (g) \rightarrow Cu(s) + CO_2\ (g)$$

Blast Furnace

- Iron is extracted from Haematite or Ironsand in a Blast Furnace
- A charge is a mixture of limestone, coke (carbon) and iron oxide (as well as it's impurities, mainly consisting of SiO_2)
- The charge is placed in the top of the blast furnace and hot air is blasted through at the bottom, making the charge glow white hot.
- The following reactions take place:

$$C\ (s) + O_2\ (g) \rightarrow CO_2\ (g) \qquad CaCO_3\ (s) \rightarrow CO_2\ (g) + CaO\ (s)$$

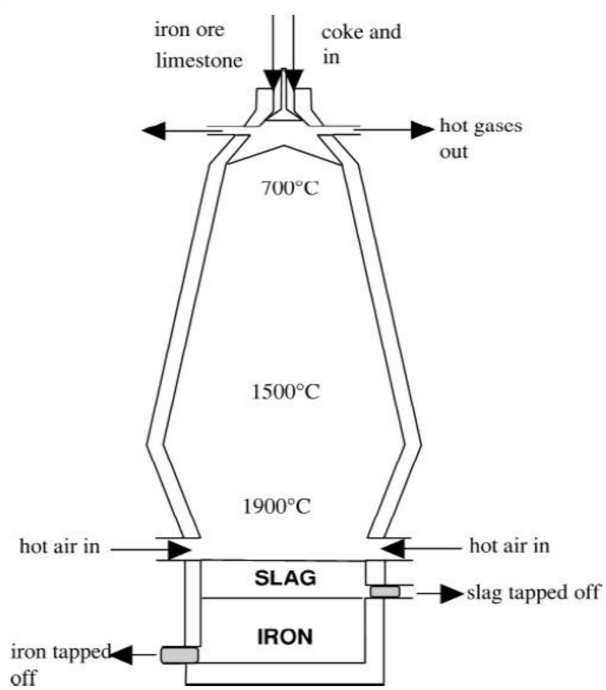
$$CO_2\ (g) + C\ (s) \rightarrow 2\ CO\ (g) \qquad CaO\ (s) + SiO_2\ (s) \rightarrow CaSiO_3\ (l)$$

This is known as slag.

$$3\ CO\ (g) + Fe_2O_3\ (s) \rightarrow 2\ Fe\ (l) + 3\ CO_2\ (g)$$

If the iron from the blast furnace solidifies, it is called cast iron and is mostly turned into steel. Steel is manufactured the following way:

- Unwanted impurities are removed in an oxygen furnace where the molten metal is poured into a furnace along with some scrap iron (to recycle it). Calcium oxide is added and a jet of oxygen is blasted into it. The calcium oxide reacts with the impurities forming slag that can be skimmed off.
- Oxygen reacts with the excess carbon, burning most of it away as CO_2, leaving some to mix with the iron to make the metal hard but not brittle.
- Other elements are then added to gain the desired steel properties.



Zinc from Zincblende

1. The ore zincblende (made mostly from Zinc Sulphide) is crushed and put into water through which air is blown. Rock particles sink and the zinc sulphide floats in a froth which is skimmed off and dried. The product of this stage is 55-75% Zinc Sulphide.
2. The Zinc Sulphide is converted to Zinc Oxide by strong heating in a furnace:

$$ZnS + 3\ O_2 \rightarrow 2\ ZnO + 2\ SO_2$$

3. Zinc Oxide is mixed with coke in a furnace and heated to 1400 °C where it is reduced to zinc:

$$ZnO + C \rightarrow Zn + CO$$

4. The zinc metal produced cools and the carbon monoxide is burnt, with the heat given out to help reduce costs of the furnace.

Heating Metal Compounds

<table>
<tr><th></th><th>Hydroxide</th><th>Nitrate</th><th>Carbonate</th></tr>
<tr><td>Potassium</td><td rowspan="2">Stable
No Reaction</td><td rowspan="2">Decomposes
$2NaNO_3 \rightarrow 2NaNO_2 + O_2$</td><td rowspan="2">Stable
No Reaction</td></tr>
<tr><td>Sodium</td></tr>
<tr><td>Calcium</td><td rowspan="11">Decomposes
$Cu(OH)_2 \rightarrow CuO + H_2O$</td><td rowspan="11">Decomposes
$2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$</td><td rowspan="11">Decomposes
$MgCO_3 \rightarrow MgO + CO_2$</td></tr>
<tr><td>Magnesium</td></tr>
<tr><td>Aluminium</td></tr>
<tr><td>Carbon</td></tr>
<tr><td>Zinc</td></tr>
<tr><td>Iron</td></tr>
<tr><td>Lead</td></tr>
<tr><td>Hydrogen</td></tr>
<tr><td>Copper</td></tr>
</table>

Halogens

- Group VII elements are known as Halogens
- They are non-metals
- They are poisonous
- Melting point and Boiling point will increase as it goes down the group because the size increases, meaning an increase in the strength of the Van der Wall forces holding them together, causing a higher temperature to be needed to break them
- Colour goes darker as it goes down the group
- Less reactive as it goes down the group because the bigger the atom, the smaller attraction between the nucleus and incoming electron
- All have similar properties because they all have seven electrons in the outer shell
- Reacts with metals to form ionic compounds, containing halide ions
- A more reactive halogen will displace a less reactive one from solution

Paper 2 Questions (10)

1. 2011

Which statements are correct?

1 Metals are often used in the form of alloys.

2 Stainless steel is an alloy of iron.

3 Alloys always contain more than two metals.

A 1 and 2 only **B** 1 and 3 only **C** 2 and 3 only **D** 1, 2 and 3

2. 2011

Which statement is true about **all** metals?

A They are attracted to a magnet.

B They are weak and brittle.

C They may be used to form alloys.

D They react with water.

3. 2011

A chemical engineer plans to produce hydrochloric acid.

Which metal is best for the reaction container?

A copper

B iron

C magnesium

D zinc

4. 2011

Alloy X is strong and has a low density.

Alloy Y is heavy but is resistant to corrosion.

Which could be uses of X and Y?

	bridge supports	aircraft	overhead cables
A	X	X	Y
B	X	Y	Y
C	Y	X	X
D	Y	Y	X

\

5. 2011

A metal is extracted from hematite, its oxide ore.

What is the metal and how is the oxide reduced?

	metal	method of reduction
A	A*l*	electrolysis
B	A*l*	heating with carbon
C	Fe	electrolysis
D	Fe	heating with carbon

6. 2012

Some properties of aluminium are listed.

1 It has mechanical strength.

2 It conducts heat.

3 It is resistant to corrosion.

4 It has a low density.

Which properties make aluminium useful for making the bodies of aircraft?

A 1, 2 and 3 **B** 1, 2 and 4 **C** 1, 3 and 4 **D** 2, 3 and 4

7. 2012

Brass is used in electrical equipment.

It contains two ……1…… elements. Together they form ……2…… .

Which words correctly complete gaps 1 and 2?

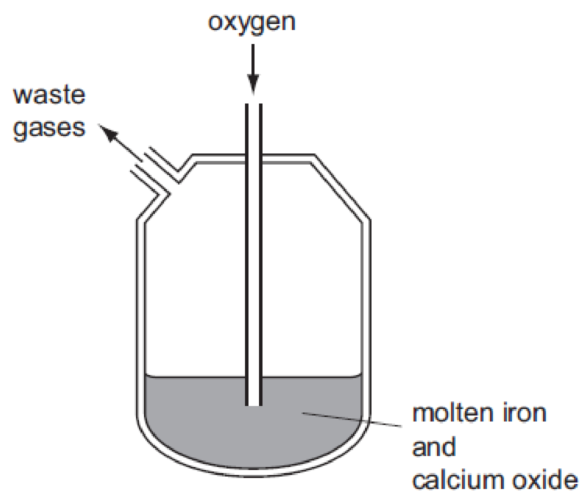
	1	2
A	metallic	a covalent compound
B	metallic	an alloy
C	non-metallic	a covalent compound
D	non-metallic	an alloy

8. 2012

The Basic Oxygen Process converts iron into steel.

In step 1, oxygen is blown into impure molten iron.

In step 2, oxides are removed by reaction with calcium oxide.



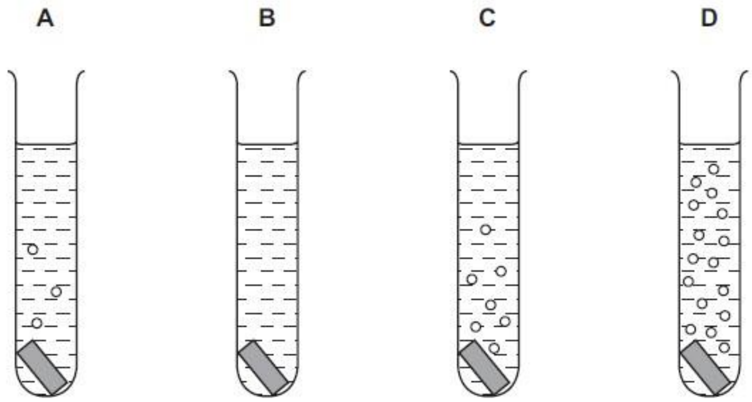
Which chemical reaction takes place in step 1 and which type of oxides are removed in step 2?

	chemical reaction in step 1	type of oxides removed in step 2
A	carbon is converted to carbon dioxide	acidic
B	carbon is converted to carbon dioxide	basic
C	iron is converted to iron(III) oxide	acidic
D	iron is converted to iron(III) oxide	basic

9. 2012

Pieces of copper, iron, magnesium and zinc are added to separate test-tubes containing dilute hydrochloric acid.

Which test-tube contains iron and dilute hydrochloric acid?



10. 2013

M is a shiny silver metal. It has a melting point of 1455 °C. Many of its compounds are green.

What is metal M?

A aluminium

B copper

C mercury

D nickel

11. 2013

Reactions of three metals and their oxides are listed in the table.

metal	reacts with cold water	metal oxide reacts with carbon
W	no	no
X	no	yes
Y	yes	no

What is the order of reactivity of the metals?

	least reactive	$\longrightarrow$	most reactive
A	W	X	Y
B	X	W	Y
C	X	Y	W
D	Y	W	X

12. 2013

Equations P and Q represent two reactions which occur inside a blast furnace.

P $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Q $CaCO_3 \rightarrow CaO + CO_2$

Which type of reactions are P and Q?

	P	Q
A	redox	redox
B	redox	thermal decomposition
C	thermal decomposition	redox
D	thermal decomposition	thermal decomposition

13. 2013

Which row describes the uses of mild steel and stainless steel?

	mild steel	stainless steel
A	car bodies, cutlery	chemical plant, machinery
B	car bodies, machinery	chemical plant, cutlery
C	chemical plant, cutlery	car bodies, machinery
D	chemical plant, machinery	car bodies, cutlery

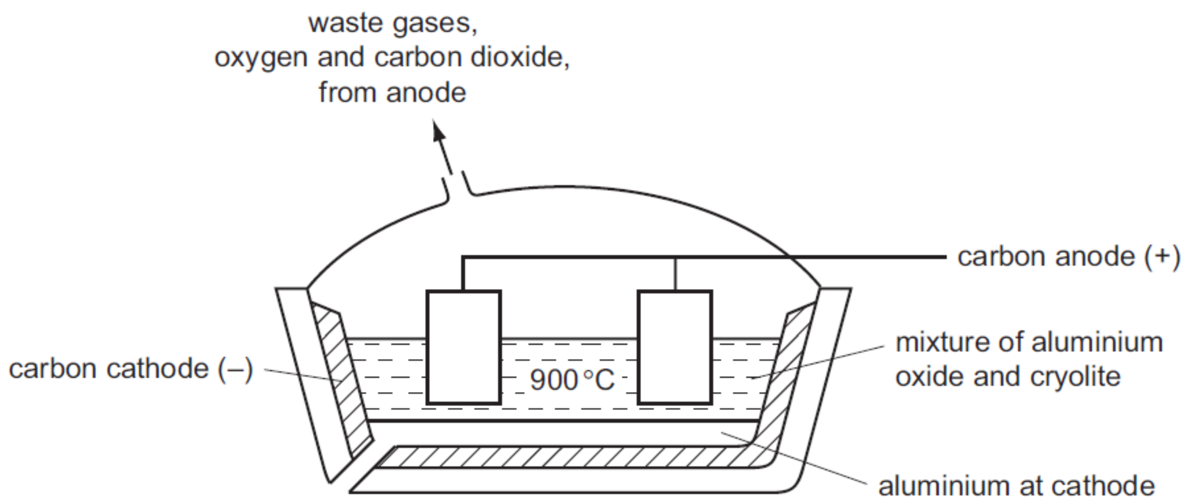
CHAPTER 10 -

Metals

Paper 4 Questions (10)

1. 2011 Question 3 (a)

Aluminium is extracted by the electrolysis of a molten mixture of alumina, which is aluminium oxide, and cryolite.



(a) **(i)** Alumina is obtained from the main ore of aluminium.
Name this ore.

.... [1]

(ii) Explain why it is necessary to use a mixture, alumina and cryolite, rather than just alumina.

....

.... [2]

(iii) Copper can be extracted by the electrolysis of an aqueous solution.
Suggest why the electrolysis of an aqueous solution cannot be used to extract aluminium.

....

....

.... [2]

2. 2011 Question 7 (a & b)

Some hydroxides, nitrates and carbonates decompose when heated.

(a) **(i)** Name a metal hydroxide which does not decompose when heated.

.... [1]

(ii) Write the equation for the thermal decomposition of copper(II) hydroxide.

.... [2]

(iii) Suggest why these two hydroxides behave differently.

.... [1]

(b) **(i)** Metal nitrates, except those of the Group 1 metals, form three products when heated. Name the products formed when zinc nitrate is heated.

....

.... [2]

(ii) Write the equation for the thermal decomposition of potassium nitrate.

.... [2]

The following metals are in order of reactivity.

potassium
zinc
copper

For those metals which react with water or steam, name the products of the reaction, otherwise write 'no reaction'.

potassium ....

....

zinc ....

....

copper ....

.... [5]

Lead is an excellent roofing material. It is malleable and resistant to corrosion. Lead rapidly becomes coated with basic lead carbonate which protects it from further corrosion.

(a) Lead has a typical metallic structure which is a lattice of lead ions surrounded by a 'sea' of mobile electrons. This structure is held together by attractive forces called a metallic bond.

(i) Explain why there are attractive forces in a metallic structure.

....

.... [2]

(ii) Explain why a metal, such as lead, is malleable.

....

.... [2]

CHAPTER 10 -

Metals

Paper 6 Questions (10)

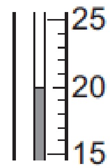
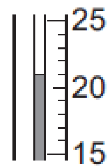
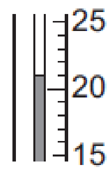
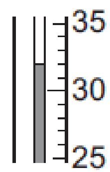
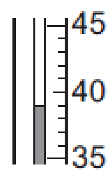
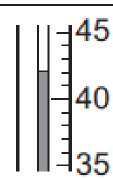
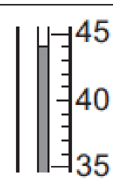
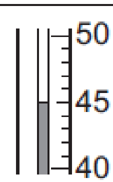
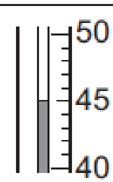
1. 2011 (Also under Unit 7)

A student investigated the reaction between aqueous copper(II) sulfate and two different metals, zinc and iron.
Two experiments were carried out.

Experiment 1
Using a measuring cylinder, 25 cm^3 of aqueous copper(II) sulfate was poured into a polystyrene cup. The temperature of the solution was measured. The timer was started and the temperature was measured every half a minute for one minute.

At 1 minute, 5 g of zinc powder was added to the cup and the mixture stirred with the thermometer. The temperature of the mixture was measured every half minute for an additional three minutes.

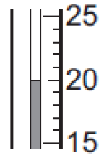
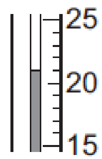
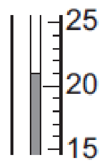
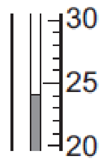
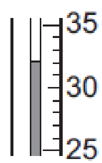
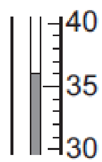
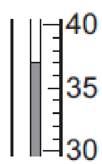
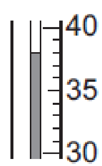
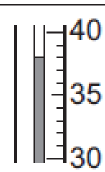
Use the thermometer diagrams in the table to record the temperatures.

time / min	thermometer diagrams	temperature / °C
0.0		
0.5		
1.0		
1.5		
2.0		
2.5		
3.0		
3.5		
4.0		

[3]

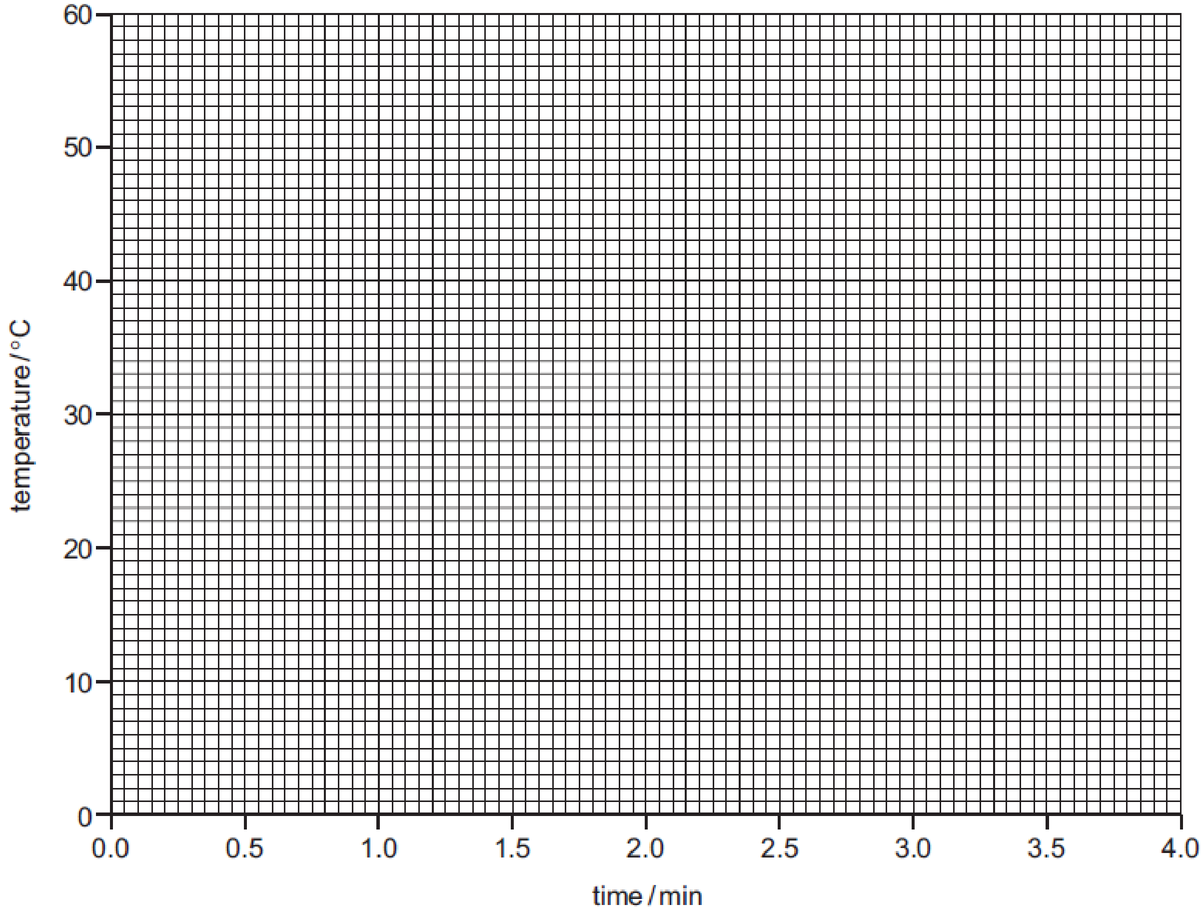
Experiment 2
Experiment 1 was repeated using 5 g of iron powder instead of the zinc powder.

(b) Use the thermometer diagrams in the table to record the temperatures.

time / min	thermometer diagrams	temperature / °C
0.0		
0.5		
1.0		
1.5		
2.0		
2.5		
3.0		
3.5		
4.0		

[3]

(c) Plot the results of both experiments on the grid below. Draw two smooth line graphs. Clearly label your graphs.



[5]

(d) **From your graph**, work out the temperature of the reaction mixture in Experiment 1 after 1 minute 15 seconds. Show clearly **on the graph** how you worked out your answer.

.... [3]

(e) What type of chemical process occurs when zinc and iron react with aqueous copper(II) sulfate?

.... [1]

(f) **(i)** Compare the temperature changes in Experiments 1 and 2.

.... [1]

(ii) Suggest an explanation for the difference in temperature changes.

....

.... [1]

(g) Explain how the temperature changes would differ in the experiments if 12.5 cm^3 of copper(II) sulfate solution were used.

....

....

.... [2]

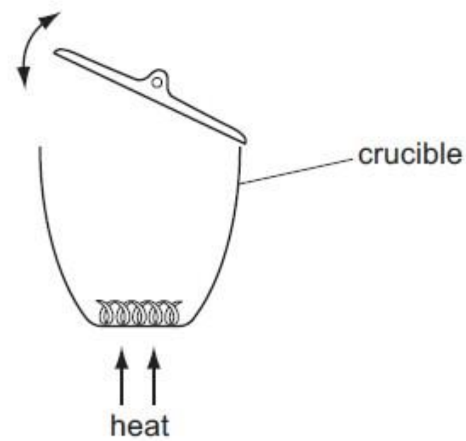
(h) Predict the effect of using lumps of zinc in Experiment 1. Explain your answer.

....

.... [2]

[Total: 21]

A student carried out an experiment to find the mass of magnesium oxide formed when magnesium burns in air.
A strip of magnesium ribbon was loosely coiled and placed in a weighed crucible, which was then reweighed.
The crucible was heated strongly for several minutes. During the heating, the crucible lid was lifted and replaced several times as in the diagram below.



The magnesium was converted into magnesium oxide. After cooling, the crucible and contents were reweighed.

(a) Describe the appearance of the

(i) magnesium .... [1]

(ii) magnesium oxide. .... [1]

(b) Name the element that reacted with the magnesium.

.... [1]

(c) Why was the lid lifted during heating?

.... [1]

(d) Suggest why the mass of the magnesium oxide was found to be **lower** than expected.

....

.... [2]

[Total: 6]

Paper 2 Questions (11)

1. 2011

A liquid turns white anhydrous copper sulfate blue and has a boiling point of 103°C.

Which could be the identity of the liquid?

A alcohol

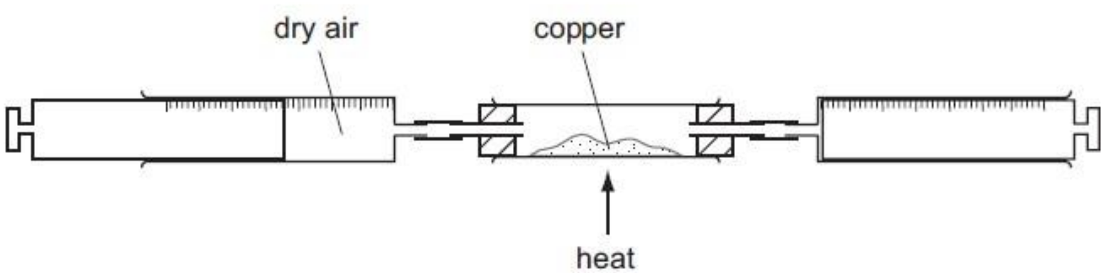
B petrol

C salt solution

D pure water

2. 2011

Dry air is passed over hot copper until all the oxygen has reacted.



The volume of gas at the end of the reaction is $120\,cm^3$.

What is the starting volume of dry air?

A $132\,cm^3$ **B** $150\,cm^3$ **C** $180\,cm^3$ **D** $600\,cm^3$

3. 2011

In which row is the air pollutant **not** correctly matched with its source?

	air pollutant	source
A	carbon monoxide	incomplete combustion of fuels
B	lead compounds	burning petrol in cars
C	nitrogen oxides	decomposing vegetation
D	sulfur dioxide	burning coal and other fossil fuels

4. 2011

Iron is a metal that rusts in the presence of oxygen and water.

Mild steel is used for ……1…… and is prevented from rusting by ……2…… .

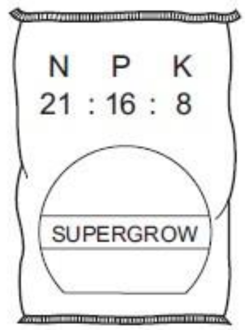
Stainless steel is prevented from rusting by ……3…… it with another metal.

Which words correctly complete gaps 1, 2 and 3?

	1	2	3
A	car bodies	greasing	covering
B	car bodies	painting	mixing
C	cutlery	greasing	covering
D	cutlery	painting	mixing

5. 2011

Which combination of chemical compounds could be used to produce the fertiliser shown?



A NH_4NO_3, $Ca_3(PO_4)_2$

B NH_4NO_3, $CO(NH_2)_2$

C NH_4NO_3, K_2SO_4, $(NH_4)_2SO_4$

D $(NH_4)_3PO_4$, KCl

6. 2011

Which pollutant gas is produced by the decomposition of vegetation?

A carbon monoxide

B methane

C nitrogen oxide

D sulfur dioxide

7. 2012

Which processes are used in the treatment of water?

A filtration and chlorination

B filtration and reduction

C neutralisation and chlorination

D neutralisation and reduction

8. 2012

A factory burns coal with a high sulfur content.

Which pollutant is **most** likely to lead to the death of trees?

A carbon dioxide

B carbon monoxide

C lead compounds

D sulfur dioxide

9. 2012

What is the correct order of abundance of the gases in the air?

A nitrogen → oxygen → argon → carbon dioxide

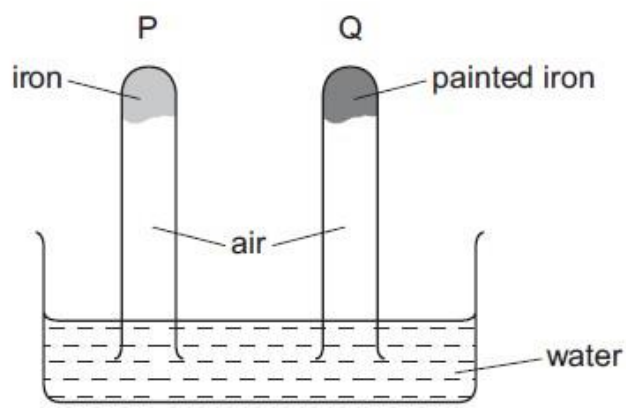
B nitrogen → oxygen → carbon dioxide → argon

C oxygen → nitrogen → argon → carbon dioxide

D oxygen → nitrogen → carbon dioxide → argon

10. 2012

The diagram shows an experiment to investigate how paint affects the rusting of iron.



What happens to the water level in tubes P and Q?

	tube P	tube Q
A	falls	rises
B	no change	rises
C	rises	falls
D	rises	no change

11. 2013

In which process is carbon dioxide **not** formed?

A burning of natural gas

B fermentation

C heating lime

D respiration

12. 2013

Farmers add calcium oxide (lime) and ammonium salts to their fields.

The compounds are not added at the same time because they react with each other.

Which gas is produced in this reaction?

A ammonia

B carbon dioxide

C hydrogen

D nitrogen

13. 2013

The diagrams show four uses of iron.

In which of these uses is the iron most likely to rust?

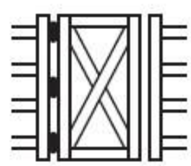
A

iron bucket electroplated with zinc

B

iron cored aluminium electricity cables

C



iron hinges on a gate

D

painted iron fence

14. 2013

Which air pollutant is **not** made when coal burns in a power station?

A carbon monoxide

B lead compounds

C nitrogen oxides

D sulfur dioxide

15. 2013

In many countries river water is used for the washing of clothes.

The same water is not considered to be safe for drinking.

Why is it **not** safe for drinking?

A because river water contains dissolved salts

B because river water may contain harmful bacteria

C because river water may contain small particles of sand

D because river water may contain soap from washing clothes

CHAPTER 11 –

Air and water

Paper 4 Questions (11)

1. 2011 Question 1 (b)

Two of the oxides are responsible for acid rain.
Identify the **two** oxides and explain their presence in the atmosphere.

....

....

....

....

.... [5]

2. 2011 Question 2

Two important greenhouse gases are methane and carbon dioxide.

(a) Methane is twenty times more effective as a greenhouse gas than carbon dioxide. The methane in the atmosphere comes from both natural and industrial sources.

(i) Describe **two** natural sources of methane.

....

.... [2]

(ii) Although methane can persist in the atmosphere for up to 15 years, it is eventually removed by oxidation.
What are the products of this oxidation?

.... [2]

(b) How do the processes of respiration, combustion and photosynthesis determine the percentage of carbon dioxide in the atmosphere?

....

....

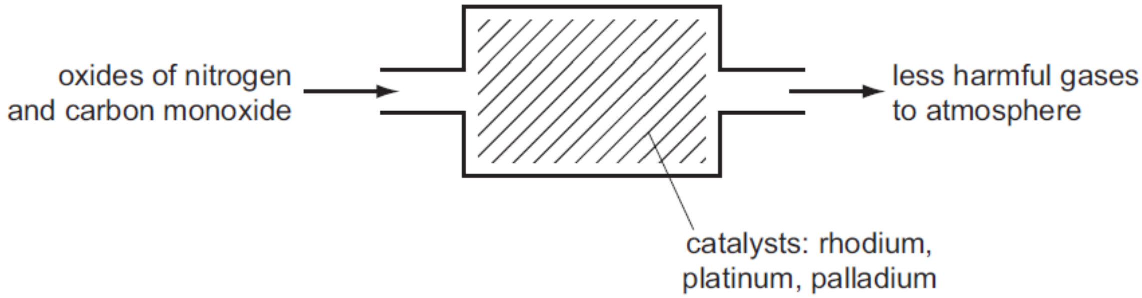
....

.... [4]

[Total: 8]

3. 2012 Question 3 (c)

(c) Catalytic converters reduce the pollution from motor vehicles.



(i) Describe how carbon monoxide and the oxides of nitrogen are formed in car engines.

....

....

....

.... [4]

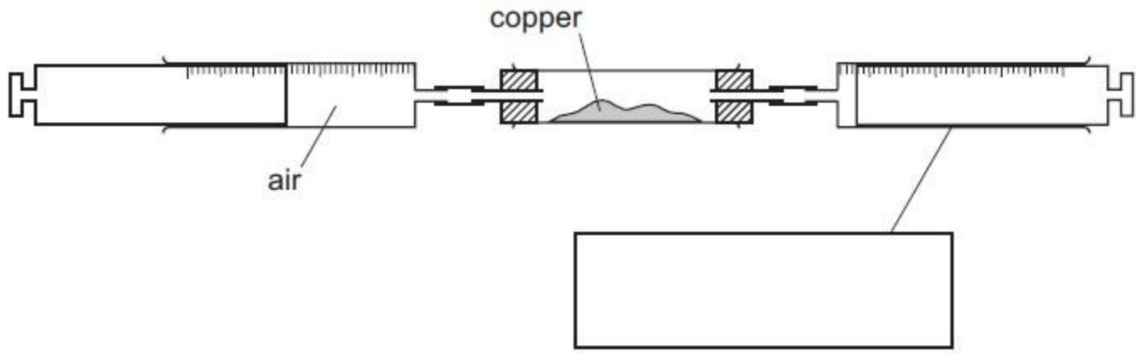
(ii) Describe the reaction(s) inside the catalytic converter which change these pollutants into less harmful gases. Include at least one equation in your description.

....

....

.... [3]

A student investigated the reaction of air with copper. 100 cm^3 of air was passed continuously over heated copper using the apparatus below. When the volume remained constant, the apparatus was left to cool and the volume of gas was measured.



(a) (i) Complete the box to show the apparatus labelled. [1]

(ii) Indicate on the diagram, with an arrow, where heat is applied. [1]

(b) What should be used to transfer the copper from a bottle to the apparatus?

.... [1]

(c) The copper changed colour from brown to .... [1]

(d) Why was the apparatus left to cool before measuring the final volume of gas?

....

.... [2]

[Total: 6]

Seawater contains sodium chloride and other salts.
Plan an experiment to find the mass of salts in 1 dm^3 of seawater.
You will be provided with a small bottle of seawater.
You should include details of the method and any apparatus used.
(1 dm^3 = 1000 cm^3)

....

....

....

....

....

....

....

.... [6]

[Total: 6]

Old documents

Some documents are stored in containers with packets of silica gel crystals. These crystals absorb water from air that enters the container. Water could damage the documents.
Anhydrous cobalt(II) chloride is added to the silica gel. As the crystals absorb water they change colour from blue to pink. Heating the silica gel in an oven removes the water from the crystals so that the crystals can be reused.

Plan an experiment to find the mass of water absorbed by a packet of silica gel crystals.

....

....

....

....

....

....

.... [6]

[Total: 6]

Paper 2 Questions (8)

1. 2011

Which property is **not** characteristic of a base?

A It reacts with a carbonate to form carbon dioxide.

B It reacts with an acid to form a salt.

C It reacts with an ammonium salt to form ammonia.

D It turns universal indicator paper blue.

2. 2011

An alloy contains copper and zinc.

Some of the zinc has become oxidised to zinc oxide.

What is the result of adding an excess of dilute sulfuric acid to the alloy?

A A blue solution and a white solid remains.

B A colourless solution and a pink / brown solid remains.

C The alloy dissolves completely to give a blue solution.

D The alloy dissolves completely to give a colourless solution.

3. 2011

The results of three tests on a solution of compound **X** are shown.

test	result
aqueous sodium hydroxide added	white precipitate formed, soluble in excess
aqueous ammonia added	white precipitate formed, soluble in excess
dilute hydrochloric acid added	bubbles of gas

What is compound **X**?

A aluminium carbonate

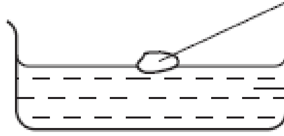
B aluminium chloride

C zinc carbonate

D zinc chloride

4. 2011

When sodium reacts with water, a solution and a gas are produced.



The solution is tested with litmus paper and the gas is tested with a splint.

What happens to the litmus paper and to the splint?

	litmus paper	splint
A	blue to red	glowing splint relights
B	blue to red	lighted splint 'pops'
C	red to blue	glowing splint relights
D	red to blue	lighted splint 'pops'

5. 2012

Which of these pairs of aqueous ions **both** react with dilute sulfuric acid to give a visible result?

A Ba^{2+} and Cl^{-}

B Ba^{2+} and CO_3^{2-}

C NH_4^{+} and Cl^{-}

D NH_4^{+} and CO_3^{2-}

6. 2012

Element X forms an acidic, covalent oxide.

Which row shows how many electrons there could be in the outer shell of an atom of X?

	1	2	6	7
A	✓	✓	✗	✗
B	✓	✗	✓	✗
C	✗	✗	✓	✓
D	✗	✓	✗	✓

7. 2012

Barium hydroxide is an alkali. It reacts with hydrochloric acid.

How does the pH of the hydrochloric acid change as an excess of aqueous barium hydroxide is added?

A The pH decreases from 14 and becomes constant at 7.

B The pH decreases from 14 to about 1.

C The pH increases from 1 and becomes constant at 7.

D The pH increases from 1 to about 14.

8. 2012

A compound is a salt if it

A can neutralise an acid.

B contains more than one element.

C dissolves in water.

D is formed when an acid reacts with a base.

9. 2013

Which are properties of an acid?

1 reacts with ammonium sulfate to form ammonia

2 turns red litmus blue

	1	2
A	✓	✓
B	✓	✗
C	✗	✓
D	✗	✗

10. 2013

Which of the following are properties of the oxides of non-metals?

	property 1	property 2
A	acidic	covalent
B	acidic	ionic
C	basic	covalent
D	basic	ionic

11. 2013

Compound X is tested and the results are shown in the table.

test	result
aqueous sodium hydroxide is added, then heated gently	gas given off which turns damp red litmus paper blue
dilute hydrochloric acid is added	effervescence, gas given off which turns limewater milky

Which ions are present in compound X?

A ammonium ions and carbonate ions

B ammonium ions and chloride ions

C calcium ions and carbonate ions

D calcium ions and chloride ions

Paper 4 Questions (8)

1. 2011 Question 1 (a)

This question is concerned with the following oxides.

sulfur dioxide
carbon monoxide
lithium oxide
aluminium oxide
nitrogen dioxide
strontium oxide

(a) **(i)** Which of the above oxides will react with hydrochloric acid but not with aqueous sodium hydroxide?

.... [1]

(ii) Which of the above oxides will react with aqueous sodium hydroxide but not with hydrochloric acid?

.... [1]

(iii) Which of the above oxides will react with both hydrochloric acid and aqueous sodium hydroxide?

.... [1]

(iv) Which of the above oxides will not react with hydrochloric acid or with aqueous sodium hydroxide?

.... [1]

2. 2012 Question 4 (c)

(i) Name a reagent that reacts with the oxides of both elements.

.... [1]

(ii) Name a reagent that reacts with only one of the oxides.

reagent ....

oxide which reacts .... [2]

3. 2012 Question 7 (b)

Strontium chloride-6-water can be made from the insoluble compound, strontium carbonate, by the following reactions.

$SrCO_3(s) + 2HCl(aq) \rightarrow SrCl_2(aq) + CO_2(g) + H_2O(l)$

$SrCl_2(aq) + 6H_2O(l) \rightarrow SrCl_2.6H_2O(s)$

The following method was used to prepare the crystals.

1 Add excess strontium carbonate to hot hydrochloric acid.
2 Filter the resulting mixture.
3 Partially evaporate the filtrate and allow to cool.
4 Filter off the crystals of $SrCl_2.6H_2O$.
5 Dry the crystals between filter papers.

(i) How would you know when excess strontium carbonate had been added in step 1?

....

.... [1]

(ii) Why is it necessary to filter the mixture in step 2?

.... [1]

(iii) In step 3, why partially evaporate the filtrate rather than evaporate to dryness?

.... [1]

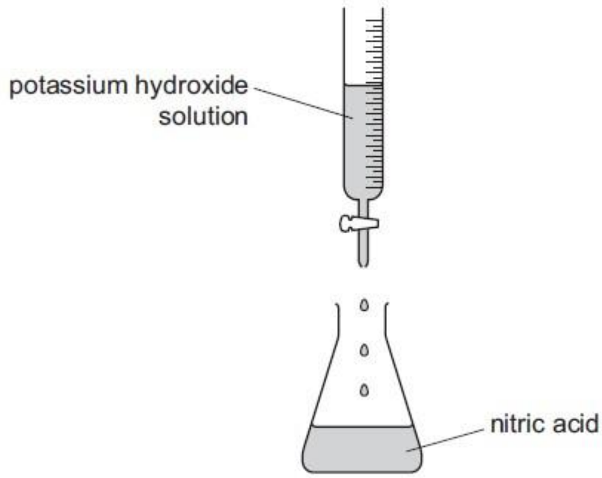
CHAPTER 8 –

Acids, bases and salts

Paper 6 Questions (8)

1. 2011

A student prepared a sample of potassium nitrate by neutralising nitric acid using potassium hydroxide solution.
25.0 cm³ of nitric acid was poured into a conical flask. Potassium hydroxide was added a little at a time from a burette as shown below.

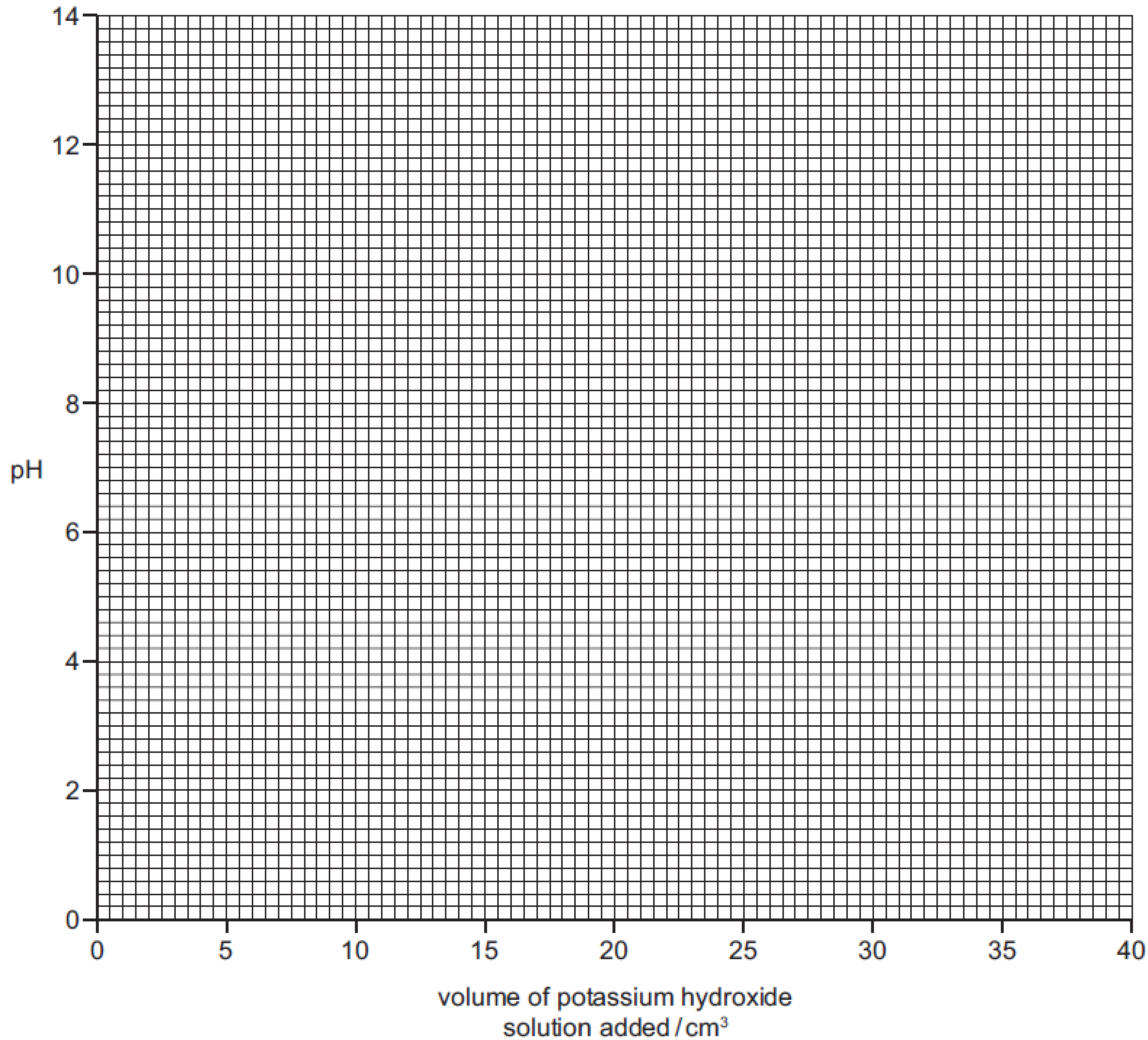


After each addition of potassium hydroxide solution the pH was measured with a pH meter and the values recorded in the table of results.

volume of potassium hydroxide solution added / cm^3	pH value
5.0	1.2
10.0	1.4
15.0	2.6
20.0	2.0
24.0	2.7
24.5	3.0
25.5	11.0
26.0	11.3
30.0	12.0
40.0	13.2

You are going to draw a graph to find the volume of potassium hydroxide solution required to neutralise the 25.0 cm^3 of nitric acid.

(a) Plot the results on the grid below and draw a smooth line graph.



[3]

(b) Which point appears to be inaccurate?

.... [1]

(c) **(i)** **Use your graph** to find the pH of the solution when 35.0 cm^3 of potassium hydroxide was added.

.... [1]

(ii) **Use your graph** to find the pH of 25.0 cm^3 of nitric acid.

Show clearly **on the grid** how you obtained your answer.

.... [2]

(d) (i) What is the pH of the solution when all of the nitric acid has just been neutralised?

.... [1]

(ii) What volume of potassium hydroxide was required to neutralise 25.0 cm^3 of nitric acid?

.... [1]

(e) Describe how the student should modify the experiment to obtain pure crystals of potassium nitrate.

....

....

....

....

.... [3]

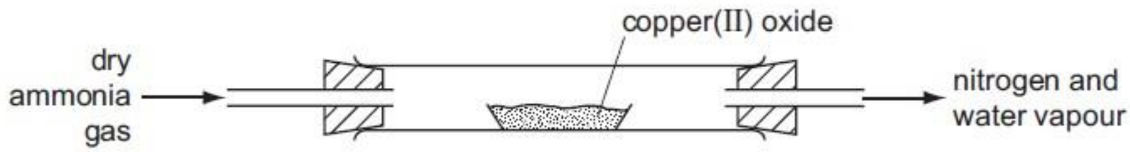
[Total: 12]

Three different liquids **P**, **Q** and **R** were analysed.
P was an aqueous solution of sulfuric acid.
The tests on the liquids and some of the observations are in the following table.
Complete the observations in the table.

tests	observations
(a) **(i)** Appearance of the liquids.	**P** .... [1] **Q** colourless, smell of vinegar **R** colourless, no smell
(ii) The pH of the liquids was tested using Universal Indicator paper.	**P** .... [1] **Q** pH 5 **R** pH 7
(b) A piece of magnesium ribbon was added to a little of each liquid. The gas given off by liquid **P** was tested.	**P** .... [2] **Q** slow effervescence **R** no reaction
(c) To a little of liquid **P**, hydrochloric acid and aqueous barium chloride were added.	.... [2]
(d) Liquid **R** was heated to boiling in a test-tube. A thermometer was used to record the constant temperature of the vapour produced.	temperature = 100 °C

A student reacted dry ammonia gas with hot copper(II) oxide.
The apparatus used is shown below.
The equation for the reaction is

$$2NH_3 + 3CuO \rightarrow 3Cu + N_2 + 3H_2O$$



(a) Indicate with an arrow where the heat is applied. [1]

(b) The colour of the copper(II) oxide would change

from .... to .... [2]

(c) Draw a labelled diagram to show how liquid water could be obtained from the water vapour produced.

[2]

(d) Suggest the effect of nitrogen on a lighted splint.

.... [1]

[Total: 6]

A mixture of two solids, **M** and **N**, was analysed.
Solid **M** was zinc sulfate which is water-soluble and solid **N** was insoluble.
The tests on the mixture, and some of the observations, are in the table.
Complete the observations in the table.

tests	observations
Distilled water was added to the mixture in a boiling tube and shaken. The contents of the tube were filtered and the filtrate and residue kept for the following tests.	
<u>tests on the filtrate</u> The filtrate was divided into four portions.	
(a) **(i)** Drops of aqueous sodium hydroxide were added to the first portion of the filtrate. Excess aqueous sodium hydroxide was then added. **(ii)** Drops of aqueous ammonia were added to the second portion of the filtrate. Excess aqueous ammonia was then added.	.... [3] .... [2]
(b) About 1 cm^3 of dilute nitric acid followed by silver nitrate solution was added to the third portion of the filtrate.	.... [1]
(c) About 1 cm^3 of dilute nitric acid followed by barium nitrate solution was added to the fourth portion of the filtrate.	.... [2]

tests	observations
<u>tests on the residue</u> **(d)** Appearance of the residue.	black solid
(e) Dilute hydrochloric acid was added to a little of the residue. The mixture was heated and the gas given off was tested with damp blue litmus paper.	effervescence pungent gas, bleached litmus paper
(f) Aqueous hydrogen peroxide was added to a little of the residue. The gas given off was tested.	effervescence glowing splint relit

(g) Identify the gas given off in test **(e)**.

.... [1]

(h) Identify the gas given off in test **(f)**.

.... [1]

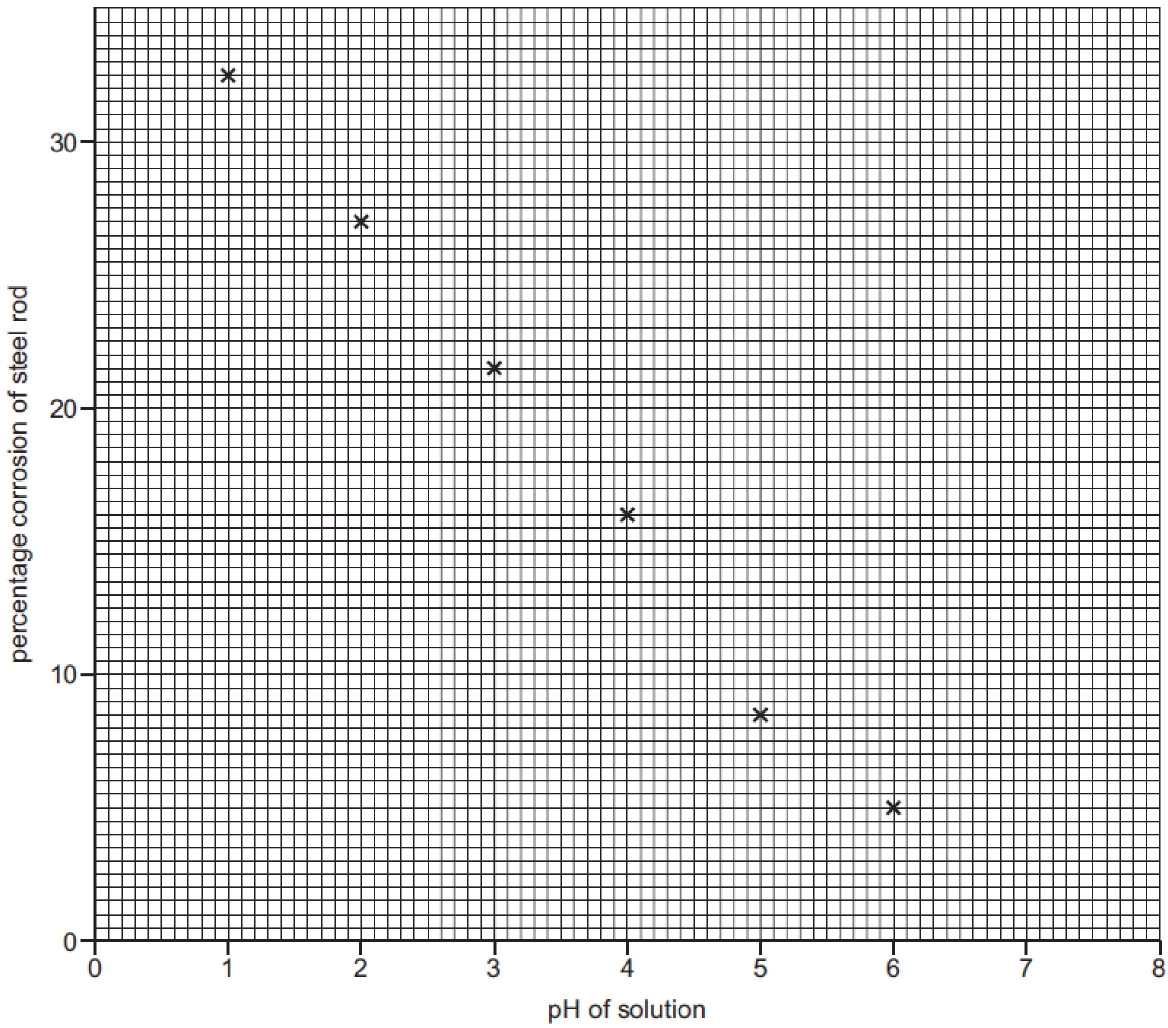
(i) What conclusions can you draw about solid **N**?

....

.... [2]

[Total: 12]

Eight steel rods of the same size were placed in solutions of different pH for one week.
The percentage corrosion of the rods was measured and the results plotted on the grid below.



(a) Draw a best fit straight line through the points. [1]

(b) Why were the steel rods the same size?

.... [1]

(c) State **one** other variable which should have been kept constant.

.... [1]

(d) State one conclusion that could be drawn from the results.

.... [1]

(e) Determine the percentage corrosion of a steel rod in a solution of pH 6.5.

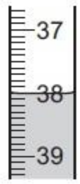
.... [1]

[Total: 5]

A student investigated the reaction between aqueous sodium hydroxide and acid **K**.
Two experiments were carried out.

(a) *Experiment 1*

Using a measuring cylinder, 25 cm^3 of acid **K** was poured into a conical flask. Phenolphthalein indicator was added to the flask. A burette was filled with aqueous sodium hydroxide to the 0.0 cm^3 mark. Aqueous sodium hydroxide was added from the burette to the flask and the mixture shaken until the solution showed a permanent colour change.
The final volume was measured. Use the burette diagram to record the final volume in the table and complete the table.



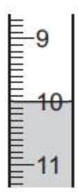
final volume

	burette reading
final volume / cm^3	
initial volume / cm^3	
difference / cm^3	

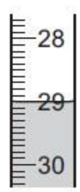
[2]

(b) *Experiment 2*

The solution was poured away and the conical flask rinsed.
Using a measuring cylinder, 50 cm^3 of acid **K** was poured into the conical flask. 0.3 g of powdered calcium carbonate was added to the flask and the flask shaken until no further reaction was observed.
Phenolphthalein was added to the mixture in the flask.
A burette was filled with the same aqueous sodium hydroxide and the initial volume measured. Aqueous sodium hydroxide was added from the burette to the flask and the mixture shaken until the solution showed a permanent colour change.
Use the burette diagrams to record the initial and final volumes in the table and complete the table.



initial volume



final volume

	burette reading
final volume / cm^3	
initial volume / cm^3	
difference / cm^3	

[2]

(c) What colour change was observed after the sodium hydroxide solution was added to the flask?

from .... to .... [2]

(d) What type of chemical reaction occurred when acid **K** reacted with sodium hydroxide?

.... [1]

(e) If Experiment 1 were repeated using 50 cm^3 of acid **K**, what volume of sodium hydroxide would be required to change the colour of the indicator?

.... [2]

(f) **(i)** What were the effects of adding 0.3 g of powdered calcium carbonate to acid **K**?

....

.... [2]

(ii) Use your answer in **(e)** to work out the difference between the volume of sodium hydroxide needed to completely react with 50 cm^3 of acid **K** and the volume of sodium hydroxide used in Experiment 2.

....

.... [2]

(iii) Estimate the mass of calcium carbonate that would be needed to be added to 50 cm^3 of acid **K** to require 0.0 cm^3 of sodium hydroxide.

.... [1]

(g) What would be the effect on the results if the solutions of acid **K** were warmed before adding the sodium hydroxide? Give a reason for your answer.

effect on results ....

reason .... [2]

Suggest the advantage, if any, of

(i) using a pipette to measure the volume of acid **K**.

....

.... [2]

(ii) using a polystyrene cup instead of a flask.

....

.... [2]

[Total: 20]

Unit: Air and Water

WATER – Notes

Uses:

Water is the most useful compound known to man.

At home we need it for drinking, cooking, washing things (including ourselves) and flushing toilet waste away.

On farms it is needed as a drink for animals, and to water crops.

In industry, they use it as a solvent, and to wash things, and to keep hot reaction tanks cool. (Cold-water pipes are coiled around the tanks.)

In power stations it is heated to make steam. The steam then drives the turbines that generate electricity.

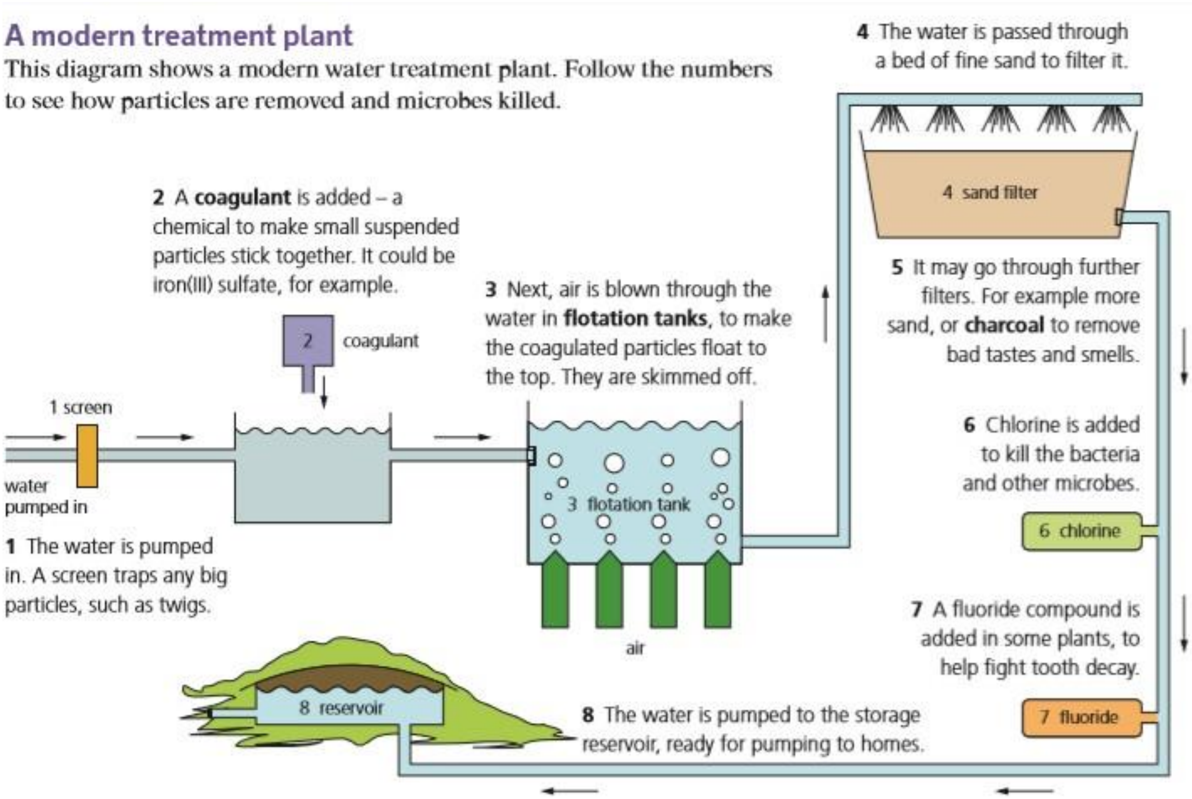
Purification of Water:

Much of the water we use is taken from rivers.

But some is pumped up from below ground, where water that has drained down through the soil lies trapped in rocks. This underground water is called groundwater.

River water is not clean – even if it looks it! It will contain particles of mud, and animal waste, and bits of dead vegetation. But worst of all are the microbes: bacteria and other tiny organisms that can make us ill.

This water has to be treated before sent to the different sources. The basic water purification basically involves stages like filtration, sedimentation and chlorination.



Simple test for water:

If a liquid contains water, it will …

1. Turn white anhydrous copper(II) sulfate blue
2. Turn blue cobalt chloride paper pink.

Both colour changes can be reversed by heating.

The test for pure water:

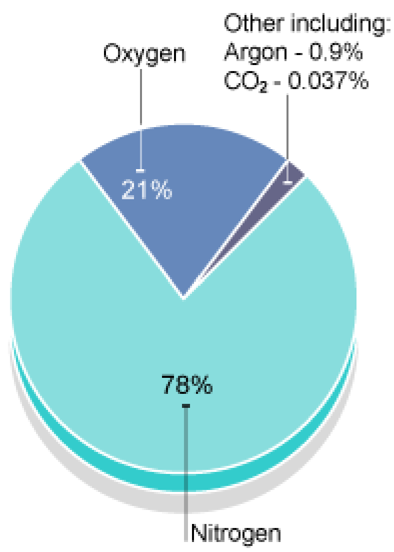
If a liquid is pure water, it will also boil at 100 °C, and freeze at 0 °C.

Air – Notes

Composition of Atmosphere

The atmosphere is the blanket of gas around the Earth. We live in the lowest layer, the troposphere. In the troposphere, we usually call the atmosphere air.

This pie chart shows the gases that make up clean air:



Oxygen: the gas we need most

The gas we depend on most is oxygen. Without it, we would quickly die. We need it for the process called **respiration** that goes on in all our cells. The energy from respiration keeps us warm, and allows us to move, and enables hundreds of different reactions to go on in our bodies.

Planes carry oxygen supplies. So do divers and astronauts.

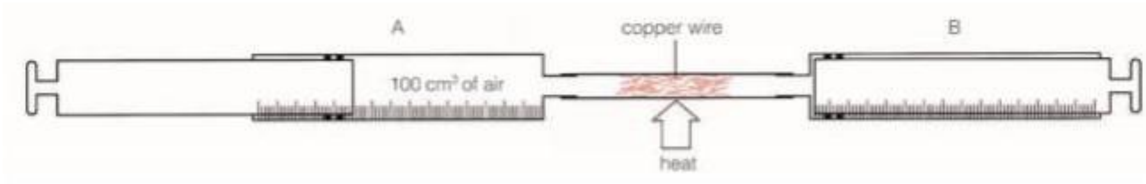
In hospitals, patients with breathing problems are given oxygen through an oxygen mask, or in an oxygen tent.

In steel works, oxygen is used in converting the impure iron from the blast furnace into steels.

A mixture of oxygen and the gas acetylene (C_2H_2) is used as the fuel in oxy-acetylene torches for cutting and welding metal.

Measuring the percentage of oxygen in air

The apparatus: A tube of hard glass is connected to two gas syringes A and B. The tube is packed with small pieces of copper wire. At the start, syringe A contains 100 cm3 of air. B is empty.



The method

These are the steps:

1 Heat the tube containing copper using a Bunsen burner. Then push in A's plunger, as shown above. This forces the air into B. When A is empty, push in B's plunger, forcing the air back to A. Repeat several times. As the air is pushed to and fro, the oxygen in it reacts with the hot copper, turning it black.

2 Stop heating the tube after about 3 minutes, and allow the apparatus to cool. Then push all the gas into one syringe and measure its volume. (It is now less than 100 cm3.)

3 Repeat steps 1 and 2 until the volume of the gas remains steady. This means all the oxygen has been used up. Note the final volume.

The results

Starting volume of air: 100 cm^3.

Final volume of air: 79 cm^3.

So the volume of oxygen in 100 cm^3 air is 21 cm^3.

The percentage of oxygen in air is therefore 21/100 * 100 = 21%.

Separating gases from the air

Air is a mixture of gases. Most of them are useful to us. But first, we must separate them from each other.

How can we separate gases? First the air is cooled until it turns into a liquid. Then the liquid mixture is separated using a method of **fractional distillation**.

This method works because the gases in **air have different boiling points**. So when liquid air is warmed up, the gases boil at different temperatures, and can be collected one by one.

The steps

1. Air is pumped into the plant, and filtered to remove dust particles.
2. Next, water vapour, carbon dioxide, and pollutants are removed (since these would freeze later and block the pipes). Like this: – First the air is cooled until the water vapour condenses to water. Then it is passed over beds of adsorbent beads to trap the carbon dioxide, and any pollutants in it.

3. Now the air is forced into a small space, or compressed. That makes it hot. It is cooled down again by recycling cold air, as the diagram shows.
4. The cold, compressed air is passed through a jet, into a larger space. It expands rapidly, and this makes it very cold. Steps 3 and 4 are repeated several times. The air gets colder each time. By 2200 °C, it is liquid, except for neon and helium. These gases are removed. They can be separated from each other by adsorption on charcoal.
5. The liquid air is pumped into the fractionating column. There it is slowly warmed up. The gases boil off one by one, and are collected in tanks or cylinders. Nitrogen boils off first.

Air Pollution

Everyone likes clean fresh air. But every year we pump billions of tones of harmful gases into the air. Most come from burning fossil fuels.

The main air pollutants

This table shows the main pollutants found in air, and the harm they do:

Pollutant	How is it formed?	What harm does it do?
Carbon monoxide, CO colourless gas, insoluble, no smell	Forms when the carbon compounds in fossil fuels burn in too little air. For example, inside car engines and furnaces.	Poisonous even in low concentrations. It reacts with the **haemoglobin** in blood, and prevents it from carrying oxygen around the body – so you die from oxygen starvation.
Sulfur dioxide, SO_2 an acidic gas with a sharp smell	Forms when sulfur compounds in the fossil fuels burn. Power stations are the main source of this pollutant.	Irritates the eyes and throat, and causes **respiratory** (breathing) problems. Dissolves in rain to form **acid rain**. Acid rain attacks stonework in buildings, especially limestone and marble – they are calcium carbonate. It lowers the pH in rivers and lakes, killing fish and other river life. It also kills trees and insects.
Nitrogen oxides, NO and **NO_2** acidic gases	Form when the nitrogen and oxygen in air react together, inside hot car engines and hot furnaces.	Cause respiratory problems, and dissolve in rain to give acid rain.
Lead compounds	A compound called tetra-ethyl lead used to be added to petrol, to help it burn smoothly in car engines. It is still added in some countries. On burning, it produces particles of other lead compounds.	Lead damages children's brains. It also damages the kidneys and nervous system in adults.

Reducing air pollution

In modern power stations, the waste gas is treated with slaked lime (calcium hydroxide). This removes sulfur dioxide by reacting with it to give calcium sulfate. The process is called **flue gas desulfurisation**.

Most countries have now banned lead in petrol. So lead pollution is much less of a problem. But it can still arise from plants where lead is extracted, and from battery factories. The exhausts of new cars are fitted with **catalytic converters**, in which harmful gases are converted to harmless ones.

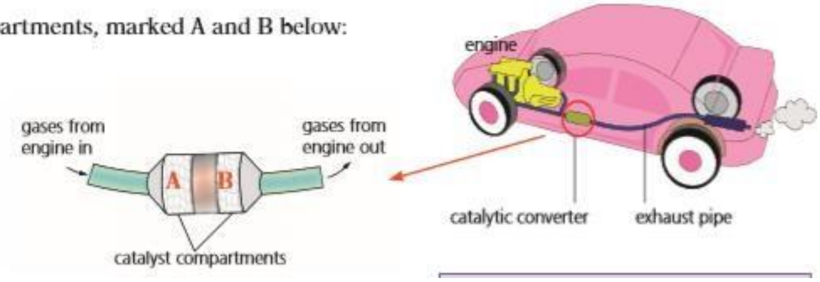
In catalytic converters, the harmful gases are adsorbed onto the surface of catalysts, where they react to form harmless gases. The catalysts speed up the reaction.

The converter usually has two compartments, marked A and B below:

In **A**, harmful compounds are **reduced**. For example:

$2NO\ (g) \longrightarrow N_2\ (g) + O_2\ (g)$

The nitrogen and oxygen from this reaction then flow into **B**.



In B, harmful compounds are oxidised, using the oxygen from A. For example:

$2CO\ (g) + O_2(g) \rightarrow 2CO_2\ (g)$. The harmless products then flow out the exhaust pipe.

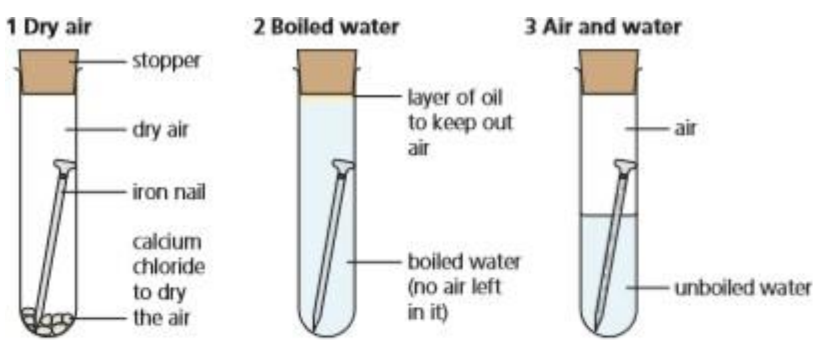
The catalysts are usually the transition elements platinum, palladium, and rhodium. They are coated onto a ceramic honeycomb, or ceramic beads, to give a large surface area for adsorbing the harmful gases. The harmless products flow out the exhaust pipe.

The rusting problem

Rust is an **iron** oxide, a usually red oxide formed by the redox reaction of **iron** and oxygen in the presence of water or air moisture.

An experiment to investigate rusting

Stand three identical nails in three test-tubes.
Now prepare the test-tubes as below, so that: – test-tube 1 contains dry air – test-tube 2 contains water but no air – test-tube 3 has both air and water.
Leave the test-tubes to one side for several days.



Result: After several days, the nails in test-tubes 1 and 2 show no signs of rusting. But the nail in test-tube 3 has rust on it. This is because: Rusting requires oxygen and water. In fact the iron is oxidised, in this reaction:

$$4Fe\ (s) + 3O_2\ (g) + 4H_2O\ (l) \rightarrow 2Fe_2O_3.2H_2O\ (s)$$

How to prevent rusting

There are two approaches.

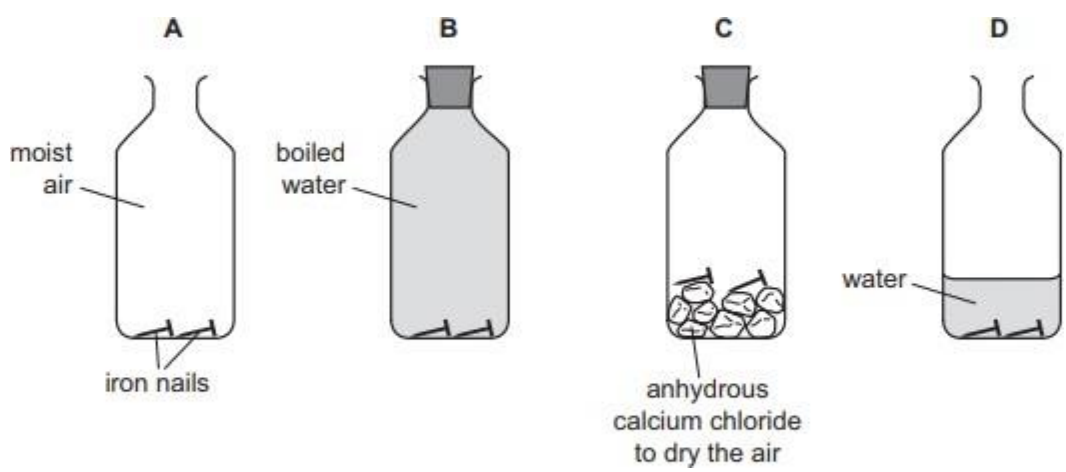
1. Cover the iron: The aim is to keep out oxygen and water. You could use:
    a. Paint. Steel bridges and railings are usually painted.
    b. Grease. Tools and machine parts are coated with grease or oil.
    c. Coat with another metal. Iron is coated with zinc, by dipping it into molten zinc, for roofing. Steel is electroplated with zinc, for car bodies. Coating with zinc has a special name: galvanising. For food tins, steel is coated with tin by electroplating.
2. Let another metal corrode instead during rusting, iron is oxidised: it loses electrons. Magnesium is more reactive than iron, which means it has a stronger drive to lose electrons. So when a bar of magnesium is attached to the side of a steel ship, or the leg of an oil rig, it will corrode instead of the iron. The magnesium dissolves. It has been sacrificed to protect the iron. This is called **sacrificial protection**. The magnesium bar must be replaced before it all dissolves. Note that zinc could also be used for this.

Practice Questions

1. The carbon monoxide in the exhaust gases comes from the incomplete combustion of hydrocarbons.
    a. What is meant by the term hydrocarbon?
    b. Give one adverse effect of carbon monoxide on health
    c. Balance the chemical equation for the complete combustion of pentane.

$$C_5H_{12} + 8O_2 \rightarrow \_\_\_CO_2 + \_\_\_\_\_\_\_\_H_2O$$

2. Some iron nails were placed in bottles under different conditions.



In which bottles will the iron nails not rust? Give reasons for your answer.

3. The table shows the percentage by volume of each of the gases present in the exhaust gases from a petrol engine with a catalytic converter.

name	percentage by volume
carbon monoxide	0.20
carbon dioxide	15.00
hydrocarbons	0.02
hydrogen	0.01
nitrogen	
oxides of nitrogen	0.02
water vapour	12.75
	total 100.00

(a) i. Calculate the percentage by volume of nitrogen in the exhaust gases.

ii. Which gas shown in the table is present in the lowest percentage by volume?

(b) i. Give one adverse effect of oxides of nitrogen on health.

ii. Balance the chemical equation for the reaction of nitrogen dioxide with sodium hydroxide.

$$.....NO_2 + ...NaOH \rightarrow NaNO_3 + NaNO_2 + H_2O$$

iii. State the name of the salt with the formula NaNO3.

4. The table shows the percentage by volume of each of the gases present in the exhaust gases from a petrol engine and a diesel engine.

name	percentage by volume from a petrol engine	percentage by volume from a diesel engine
nitrogen	72.00	67.00
carbon dioxide	14.00	
water vapour	12.00	11.00
carbon monoxide	1.50	0.05
oxides of nitrogen	0.25	0.70
hydrocarbons	0.24	0.22
sulfur dioxide	0.01	0.03
oxygen	0.00	9.00
	total 100.00	total 100.00

Describe three ways in which the composition of the exhaust gases from the petrol engine differ from the composition of the exhaust gases from the diesel engine.

5. Air is a mixture. Nitrogen and oxygen are the two most common gases in air.

a. What is meant by the term mixture?

b. State the percentage of oxygen, to the nearest whole number, in clean dry air.

c. Describe the steps in the industrial process which enables nitrogen and oxygen to be separated from clean dry air. Use scientific terms in your answer.

d. Which physical property of nitrogen and oxygen allows them to be separated?

6. The table shows the composition of the atmosphere of the planet Neptune.

gas present	percentage of gas in Neptune's atmosphere
helium	19.5
hydrogen	78.8
methane	
other gases	0.2

a. Complete the table to calculate the percentage of methane in Neptune's atmosphere.

b. Describe how Neptune's atmosphere differs from the Earth's atmosphere. Give three differences.

7. After chlorination, the water is acidic. A small amount of slaked lime is added to the acidic water.

(i) Explain why slaked lime is added.

(ii) What is the chemical name for slaked lime?

(iii) State one other use of slaked lime.

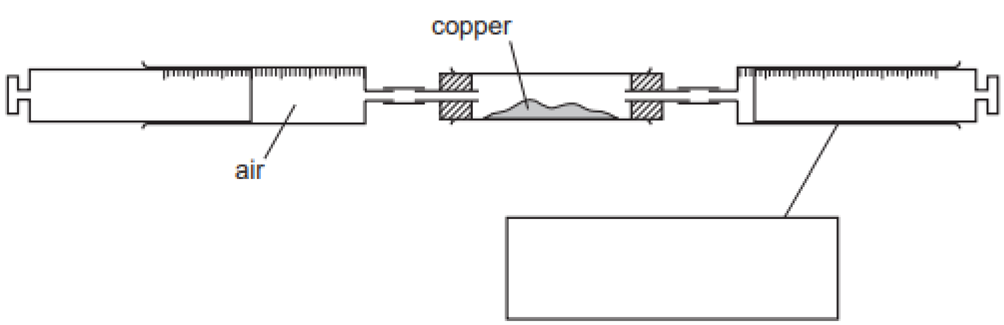
8. Water is needed for industry and in the home.

a. Rain water is collected in reservoirs. How is it treated before entering the water supply.

b. State two industrial uses of water

c. State two uses of water in the home.

9. The student investigated the reaction of air with copper. 100cm3 of air was passed continuously over heated copper using the apparatus below. When the volume remained constant, the apparatus was left to cool and the volume of gas was measured.



(a) i. complete the box to show the apparatus labelled

ii. indicate on the diagram, with an arrow, where heat is applied.

(b) What should be used to transfer the copper from a bottle to the apparatus.

(c) The copper changed color from brown to ______________________

(d) Why was the apparatus left to cool before measuring the final volume of gas?

Sulfur and its compounds– Notes

Sources of Sulphur

Sulfur is a non-metal. It is quite a common element in the Earth's crust.

It is found, as the element, in large underground beds in several countries, including Mexico, Poland and the USA. It is also found around the rims of volcanoes.

It occurs as a compound in many metal ores. For example in the lead ore galena, which is lead(II) sulfide, PbS.

Sulfur compounds also occur naturally in the fossil fuels: coal, petroleum (crude oil) and natural gas

Extracting the sulfur

From oil and gas Most sulfur is now obtained from the sulfur compounds found in petroleum and natural gas. These compounds are removed to help reduce air pollution. For example natural gas is mainly methane. But it can have as much as 30% hydrogen sulfide. This is separated from the methane. Then it is reacted with oxygen, with the help of a catalyst, to give sulfur:

$2H_2S\ (g) + O_2\ (g) \rightarrow 2S\ (s) + 2H_2O\ (l)$

From sulfur beds About 5% of the sulfur we use comes from the underground sulfur beds. Superheated water is pumped down to melt the sulfur and carry it to the surface. (It melts at 115 °C.)

The properties of sulfur

It is a brittle yellow solid.

It has two different forms or allotropes, as shown on the right.

Because it is molecular, it has quite a low melting point.

Like other non-metals, it does not conduct electricity. 6 Like most non-metals, it is insoluble in water. It reacts with metals to form sulfides. For example with iron it forms iron(II) sulphide

It burns in oxygen to form sulfur dioxide: $S\ (s) + O_2\ (g) \rightarrow SO_2\ (g)$

Uses of sulfur

Most sulfur is used to make sulfuric acid.

It is added to rubber, for example for car tyres, to toughen it. This is called vulcanizing the rubber.

It is used in making drugs, pesticides, dyes, matches, and paper.

It is used in making cosmetics, shampoos, and body lotions.

It is added to cement to make sulfur concrete. This is not attacked by acid. So it is used for walls and floors in factories that use acid.

Sulfur dioxide

Sulfur dioxide (SO_2) is a gas. It forms when sulfur burns in air.

It is a colourless gas, heavier than air, with a strong, choking smell.

Like most non-metal oxides, it is an acidic oxide. It dissolves in water, forming sulfurous acid, H_2SO_3: This breaks down easily again to sulfur dioxide and water.

It acts as a bleach when it is damp or in solution. This is because it removes the colour from coloured compounds by reducing them.

It can kill bacteria.

Uses of sulfur dioxide

Its main use is in the manufacture of sulfuric acid.

It is used to bleach wool, silk, and wood pulp for making paper.

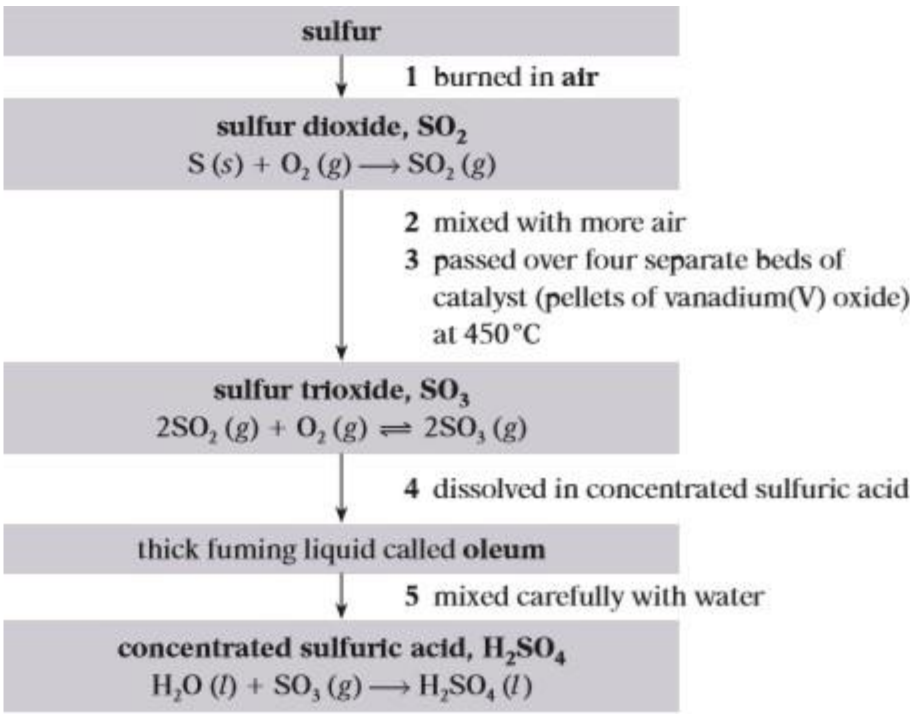
It is used as a sterilizing agent in making soft drinks and jam, and in drying fruit. It stops the growth of bacteria and moulds.

Sulfuric acid

Making sulfuric acid by the Contact process

More sulfuric acid is made than any other chemical! Most of it is made by the Contact process. The raw materials are: sulfur, air, and water or sulfur dioxide, air, and water. The sulfur dioxide is obtained when sulfide ores, such as lead and zinc ores, are roasted in air to extract the metal from them.

Starting with sulfur, the steps in the Contact process are:



Things to note about the Contact process

The reaction in step 3 is reversible. The sulfur trioxide continually breaks down again. So the mixture is passed over four separate beds of catalyst, to give the reactants further chances to react.

Sulfur trioxide is removed between the last two beds of catalyst (using step 4) in order to increase the yield.

The reaction in step 3 is exothermic. So yield rises as temperature falls. But the catalyst will not work below 400 °C, and it works better at higher temperatures. So 450 °C is a compromise.

To keep the temperature down to 450 °C, heat must be removed from the catalyst beds. So pipes of cold water are coiled around them to carry heat away. The heat makes the water boil. The steam is used to generate electricity for the plant, or for heating buildings.

In step 4, the sulfur trioxide is dissolved in concentrated acid instead of water, because with water, a thick, dangerous mist of acid forms.

Uses of sulfuric acid

Sulfuric acid is one of the world's most important chemicals. It has thousands of uses in industry. Its main uses are in making:

Fertilisers such as ammonium sulfate

Paints, pigments, and dyestuffs

Fibres and plastics

Soaps and detergents.

It is also the acid used in car batteries.

Dilute sulfuric acid

In the lab, dilute sulfuric acid is made by adding the concentrated acid to water. And never the other way round – because so much heat is produced that the acid could splash out and burn you.

Dilute sulfuric acid shows the usual reactions of acids:

1. acid + metal → salt + hydrogen 2
2. acid + metal oxide or hydroxide → salt + water 3
3. acid + carbonate→ salt + water + carbon dioxide
   Its salts are called sulfates. And reactions 2 and 3 are neutralisations: water is produced as well as a salt.

Concentrated sulfuric acid

Concentrated sulfuric acid is a **dehydrating agent**. It removes water. It 'likes' water so much that it removes hydrogen and oxygen atoms from other substances.

Practice Questions

Sulfuric acid can be manufactured from the raw materials sulfur, air and water. The process can be divided into four stages.

stage 1 converting sulfur into sulfur dioxide

stage 2 converting sulfur dioxide into sulfur trioxide

stage 3 converting sulfur trioxide into oleum, $H_2S_2O_7$

stage 4 converting oleum into sulfuric acid

stage 1

(a) (i) Describe how sulfur is converted into sulfur dioxide.

.... [1]

(ii) Write a chemical equation for the conversion of sulfur into sulfur dioxide.

.... [1]

stage 2

(b) Sulfur dioxide is converted into sulfur trioxide according to the following equation.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

The reaction is carried out at a temperature of 450°C and a pressure of 1–2 atmospheres using a catalyst. The energy change, ΔH, for the reaction is −196 kJ/mol.

(i) What is the meaning of the symbol $\rightleftharpoons$?

.... [1]

(ii) Name the catalyst used in this reaction.

.... [1]

(iii) Why is a catalyst used?

.... [1]

(iv) If a temperature higher than 450°C were used, what would happen to the amount of sulfur trioxide produced? Give a reason for your answer.

....

.... [2]

(v) Suggest a reason why a temperature lower than 450°C is **not** used.

....

.... [1]

(vi) If a pressure higher than 1–2 atmospheres were used, what would happen to the amount of sulfur trioxide produced? Give a reason for your answer.

....

.... [2]

stage 3

(c) **(i)** What is added to sulfur trioxide to convert it into oleum?

.... [1]

(ii) Write a chemical equation for the conversion of sulfur trioxide into oleum.

.... [1]

stage 4

(d) **(i)** What is added to oleum to convert it into sulfuric acid?

.... [1]

(ii) Write a chemical equation for the conversion of oleum into sulfuric acid.

.... [1]

(e) Give **one** use of sulfuric acid.

.... [1]

(f) Sulfuric acid reacts with a hydrocarbon called benzene to produce benzenesulfonic acid, $C_6H_5SO_3H$. Benzenesulfonic acid is a strong acid which ionises to produce hydrogen ions, H^+, and benzenesulfonate ions, $C_6H_5SO_3^-$.

(i) What is meant by the term *strong acid*?

.... [1]

(ii) Describe how to show that a 1 mol/dm^3 solution of benzenesulfonic acid is a strong acid.

....

.... [2]

(iii) Write a chemical equation for the reaction between benzenesulfonic acid and sodium carbonate, Na_2CO_3.

.... [2]

[Total: 20]

Sulfuric acid is produced by the Contact process. The steps of the Contact process are shown.

$$\text{starting material} \xrightarrow{\textbf{step 1}} \text{sulfur dioxide} \xrightarrow{\textbf{step 2}} \text{sulfur trioxide} \xrightarrow{\textbf{step 3}} \text{oleum} \xrightarrow{\textbf{step 4}} \text{sulfuric acid}$$

(a) Sulfur is a common starting material for the Contact process.

Name a source of sulfur.

.... [1]

(b) Describe **step 2**, giving reaction conditions and a chemical equation. Reference to reaction rate and yield is not required.

....

....

....

....

....

....

.... [5]

(c) **Step 3** involves adding sulfur trioxide to concentrated sulfuric acid to form oleum.

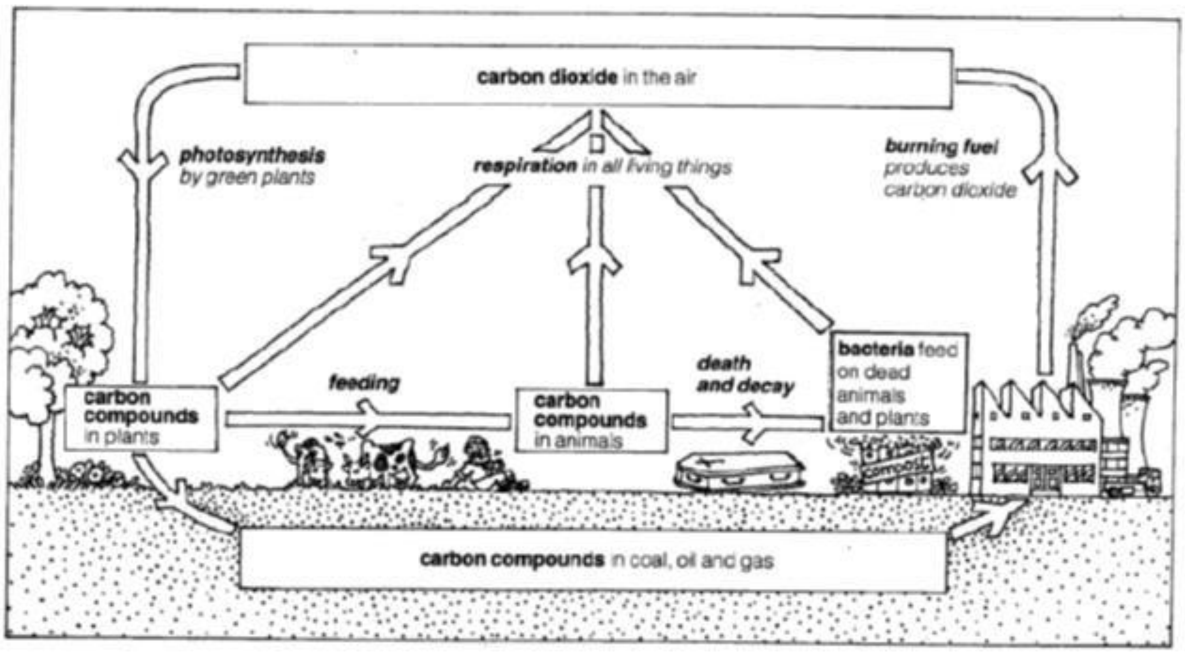
Complete the chemical equation for this reaction.

$$H_2SO_4 + SO_3 \rightarrow \text{......................}$$

[1]

Carbon and its compounds– Notes

The Carbon cycle



Note about the carbon cycle

Carbon moves between the atmosphere, ocean, and living things, in the form of carbon dioxide. Carbon dioxide is …

– removed from the atmosphere by photosynthesis, and dissolving in the ocean

– added to it by respiration, and the combustion (burning) of fuels that contain carbon.

Carbonates

Carbonates are compounds that contain the carbonate ion, $CO_3{}^{-2}$

One example is calcium carbonate, $CaCO_3$, which occurs naturally as limestone, chalk and marble. **These are the main properties of carbonates**:

They are insoluble in water – except for sodium, potassium, and ammonium carbonates, which are soluble.

They react with acids to form a salt, water, and carbon dioxide.

Most of them break down on heating, to an oxide and carbon dioxide:

$CaCO_3$ (s) → CaO (s) + CO2 (g)

But sodium and potassium carbonates do not break down, since the compounds of these reactive metals are more stable.

Uses

It is used to prepare quick lime (CaO) & Slaked lime ($Ca(OH)_2$.)

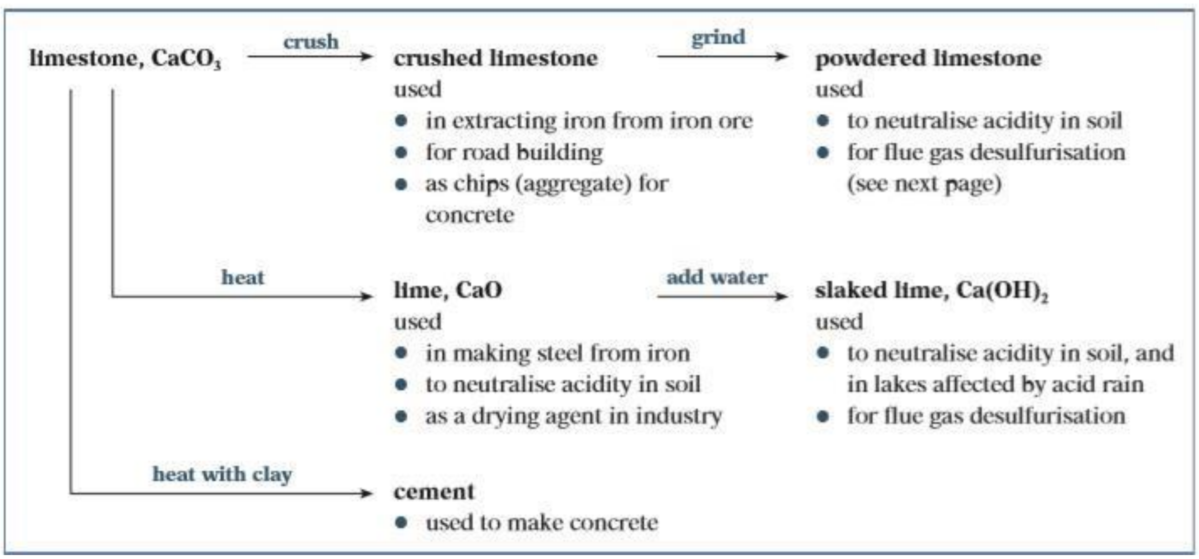
1. **Calcium carbonate ($CaCO_3$), <u>when heated</u> breaks down to form calcium oxide (Quick lime) and carbon dioxide.** This reaction is called **thermal decomposition reaction.**

   $$CaCO_3 \xrightarrow{heat} CaO + CO_2$$

2. **Calcium oxide (Quick lime) reacts with water to produce calcium hydroxide (Slaked lime).**

   $CaO + H_2O \rightarrow Ca(OH)_2$.

Uses of Limestone



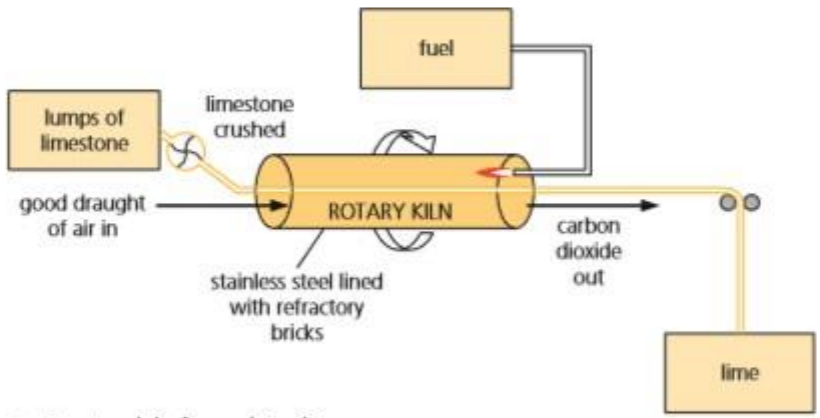
Lime

When limestone is heated, it breaks down to lime (or quicklime):

$CaCO3\ (s) \rightarrow CaO\ (s) + CO2\ (g)$

This is thermal decomposition.

The drawing shows a lime kiln. The kiln is heated. Limestone is fed in at one end. Lime comes out the other. **The reaction is reversible**. So the calcium oxide and carbon dioxide could combine again. But air is blown through the kiln to carry the carbon dioxide away before it has a chance to react.



▲ A rotary kiln for making lime.

<u>Uses of quicklime (CaO):</u> to make Steel from iron, to neutralize acidity in soil, a drying agent in industry (as it can absorb moisture)

Slaked lime

Slaked lime forms when water is added to lime. The reaction is exothermic, so the mixture hisses and steams. Conditions are controlled so that the slaked lime forms as a fine powder:

$CaO (s) + H2O (l) \rightarrow Ca(OH)_2 (s)$

Slaked lime is used to neutralise acidity in soil, and in lakes. In the lab, we use it to test for carbon dioxide. Limewater is a weak solution of calcium hydroxide, which is sparingly soluble in water.

Note: In controlling soil acidity calcium carbonate is better than calcium oxide because –it is less expensive/ less basic/ easier to handle.

Cement

Cement is made by mixing limestone with clay, heating the mixture strongly in a kiln, adding gypsum (hydrated calcium sulfate), and grinding up the final solid to give a powder.

Flue gas desulfurisation

Flue gas desulfurisation means the removal of sulfur dioxide from the waste gases at power stations, before they go out the flue (chimney). It is usually carried out using a runny mixture of powdered limestone, or slaked lime, and water. The mixture is sprayed through the waste gases, or the gases are bubbled through it. When slaked lime is used, the reaction that removes the sulfur dioxide is: $Ca(OH)_2 (s) + SO_2 (g) \rightarrow CaSO_3 (s) + H_2O (l)$

Then the calcium sulfite can be turned into hydrated calcium sulfate:

$2CaSO_3 (s) + O2 (g) + 4H2O (l) \rightarrow 2CaSO_4.2H_2O (s)$

Hydrated calcium sulfate is known as gypsum. It is used in making cement, plaster board, plaster for broken limbs, and other products.

Practice Questions

Quicklime, which is calcium oxide, is made by heating limestone in a furnace.

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$$

The reaction does not come to equilibrium.

(a) Suggest why the conversion to calcium oxide is complete.

.... [1]

(b) Calcium hydroxide, slaked lime, is made from calcium oxide.

Write an equation for this reaction.

.... [2]

(c) Calculate the maximum mass of calcium oxide which could be made from 12.5 tonnes of calcium carbonate. 1 tonne = 1×10^6 g.

....

....

.... [2]

(d) Limestone is used in agriculture to reduce the acidity of soil and for the desulfurisation of flue gases in power stations.

(i) Most crops thrive in soils whose pH is close to 7. Calcium carbonate, which is insoluble in water, and calcium oxide, which is slightly soluble in water, are both used to reduce the acidity of soils.

Suggest **two** advantages of using calcium carbonate for this purpose.

1. ....

2. .... [2]

(ii) Explain the chemistry of desulfurisation of flue gases.

....

....

....

.... [3]

(iii) Give **one** other use of calcium carbonate.

.... [1]

[Total: 11]

Calcium carbonate is an important raw material.

(a) Name a rock which is made up of calcium carbonate.

.... [1]

(b) When calcium carbonate is heated strongly, it decomposes.

$$CaCO_3 \rightarrow CaO + CO_2$$

(i) Calculate the relative formula mass of:

$CaCO_3$ ....

CaO .... [2]

(ii) 7.00 kg of calcium oxide was formed. What mass of calcium carbonate was heated?

....

.... [2]

(c) Calcium carbonate is used to control soil acidity.

(i) Why is it important to control soil acidity?

.... [1]

(ii) Both calcium carbonate, insoluble in water, and calcium oxide, slightly soluble, are used to increase soil pH. Suggest **two** advantages of using calcium carbonate.

....

.... [2]

(iii) Give **one** use of calcium carbonate other than for making calcium oxide and controlling soil pH.

.... [1]

Organic Chemistry

Organic compound, any of a large class of chemical **compounds** in which one or more atoms of carbon are covalently linked to atoms of other elements, most commonly hydrogen, oxygen, or nitrogen.

Naming Organic Compounds

In order to name organic compounds you must first memorize a few basic names. These names are listed within the discussion of naming alkanes. In general, the base part of the name reflects the **number** of carbons in what you have assigned to be the **parent chain**. The **suffix** of the name reflects the type(s) of functional group(s) present on (or within) the parent chain. Other groups which are attached to the parent chain are called **substituents**.

Identify the longest carbon chain. This chain is called the parent chain.

Number of Carbons	Name
1	**meth**ane
2	**eth**ane
3	**prop**ane
4	**but**ane
5	**pent**ane
6	**hex**ane
7	**hept**ane
8	**oct**ane
9	**non**ane
10	**dec**ane
11	**undec**ane
12	**dodec**ane

Identify all of the substituents (groups appending from the parent chain).

Number the carbons of the parent chain from the end that gives the substituents the lowest numbers.

If the same substituent occurs more than once, the location of each point on which the substituent occurs is given. In addition, the number of times the substituent group occurs is indicated by a prefix (di, tri, tetra, etc.).

If there are two or more different substituents they are listed in alphabetical order using the base name (ignore the prefixes).

The halo- substituent is considered of equal rank with an alkyl substituent in the numbering of the parent chain. The halogens are represented as follows:

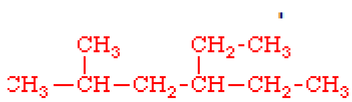
F	fluoro-
Cl	chloro-
Br	bromo-
I	iodo-

Double bonds in hydrocarbons are indicated by replacing the suffix **-ane**with **-ene**. Double bonds have priority over alkyl and halo substituents).

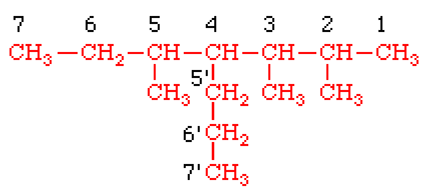
Alcohols are named by replacing the suffix **-ane** with **-anol**. The hydroxyl group takes precedence over alkyl groups and halogen substituents, as well as double bonds, in the numbering of the parent chain.

Carboxylic acids are named by counting the number of carbons in the longest continuous chain including the carboxyl group and by replacing the suffix **-ane** of the corresponding alkane with **-anoic acid**. The carboxyl group takes precedence over alkyl groups and halogen substituents, as well as double bonds, in the numbering of the parent chain.

Examples:



4-ethyl-2-methylhexane



2,3,5-trimethyl-4-propylheptane

Functional groups

A functional group is the part of a molecule that largely dictates how the molecule will react. For example, all the alkenes have similar reactions because they all have the same functional group, the C- C bond.

Homologous series

In a homologous series:

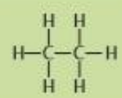
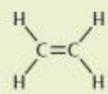
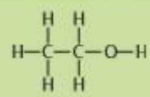
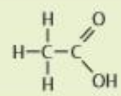
All the compounds fit the same general formula. For the alkanes the general formula is CnH_{2n+2}, where n is a number. For methane n is 1, giving the formula CH_4. For ethane n is 2, giving C_2H_6. For propane n is 3, giving C_3H_8.

The chain length increases by 1 each time.

As the chain gets longer, the compounds show a gradual change in properties. For example, their boiling points rise, and they burn less easily.

Comparing families

This table shows one member from each of the four families. Compare them.

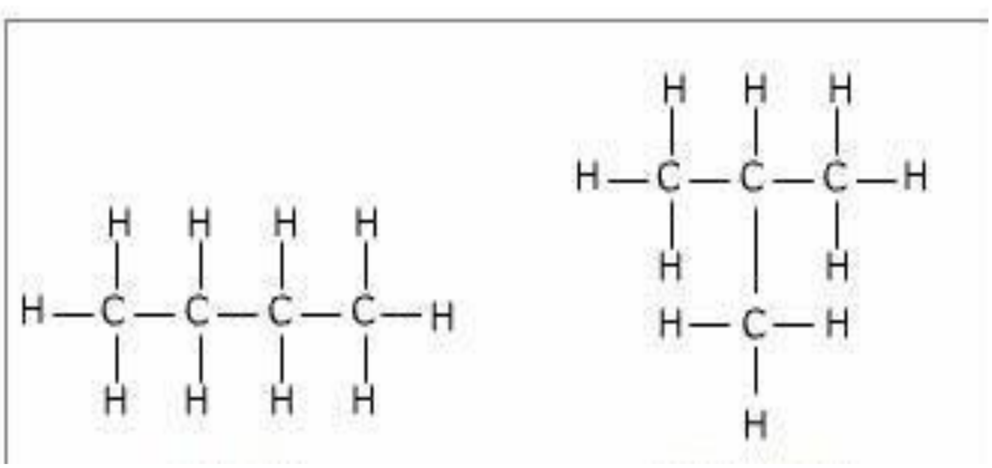
Family	A member	Structural formula	Comments
alkanes	ethane, C_2H_6		• The alkanes contain only carbon and hydrogen, so they are **hydrocarbons**. • The bonds between their carbon atoms are all **single bonds**.
alkenes	ethene, C_2H_4		• The alkenes are **hydrocarbons**. • All alkenes contain **carbon – carbon double bonds**. • The C=C bond is called their **functional group**.
alcohols	ethanol, C_2H_5OH		• The alcohols are not hydrocarbons. • They are like the alkanes, but with an **OH group**. • The OH group is their functional group.
carboxylic acids	ethanoic acid, CH_3COOH		• The carboxylic acids are not hydrocarbons. • All carboxylic acids contain the **COOH group**. • The COOH group is their functional group.

Isomers

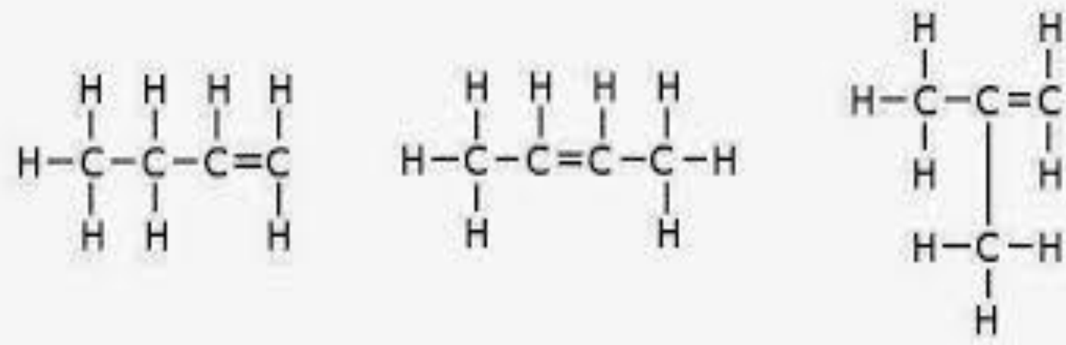
Substances with same molecular formula but different structural formula.

Example:

- **Butane**, or C4H10, has two structural isomers called *normal butane*, or *unbranched butane*, and *branched butane*. These isomers are called simply **butane** and **2-methylpropane.**

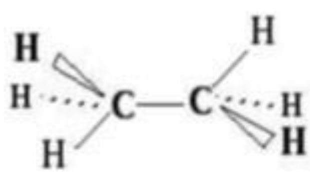
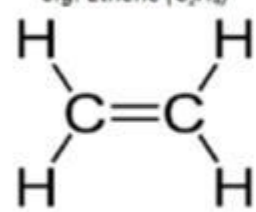
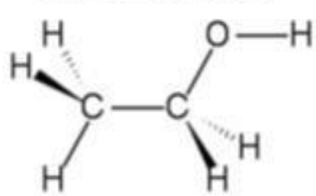
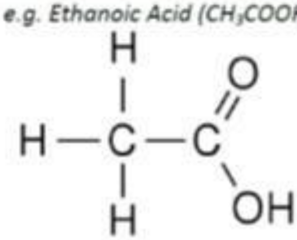


- Butene (C4H8) has following isomers



but-1-ene   but-2-ene   2-methylprop-1-ene

Different families and their Properties

Alkanes	Alkenes
• Non-polar molecules • Weak intermolecular attractions • Low melting point and boiling point (but increases as size increases) • Lower density than water • Saturated Hydrocarbons • Single bonds only • Formula = C_nH_{2n+2} • Gas state between 1-4 C's • Liquid state between 5-17 C's	• Non-polar molecules • Weak intermolecular attractions • Low melting point and boiling point (but increases as size increases) • Unsaturated Hydrocarbons • Contain a double bond • Formula = C_nH_{2n} • Extremely reactive due to double bonds breaking • Turns bromine water from red to colourless
e.g. Ethane (C_2H_6) 	*e.g. Ethene (C_2H_4)* 
Alcohols	**Carboxylic Acids**
• Polar molecules (decreases as more carbons added) • Strong intermolecular attractions • Formula = $C_nH_{2n+1}OH$ • Colourless volatile liquids • Burn cleanly and efficiently, but with less energy from presence of oxygen	• Polar molecules • Formula = $C_nH_{2n+1}COOH$ • Strong intermolecular attractions • Weak acids • Formed through the oxidation of Alcohols • Reacts with Alcohols to form Esters
e.g. Ethanol (C_2H_5OH) 	*e.g. Ethanoic Acid (CH_3COOH)* 

Alkane Reactions

1. Combustion – alkanes burn in oxygen to form carbon dioxide and water as long as sufficient oxygen is present; if insufficient, carbon monoxide or carbon will be produced instead of carbon dioxide
2. Substitution – alkanes will react with halogen molecules in a substitution reaction *e.g.*

$$C_2H_6 + Br_2 \rightarrow C_2H_5Br + HBr$$

Alkene Reactions

1. Hydrogenation - Addition by hydrogen. Alkanes are formed when the H_2 adds to the alkene molecule. A catalyst of nickel or platinum is used at a temperature of about 150 °C *e.g.*

$$CH_2 = CH_2 + H_2 \rightarrow CH_3CH_3$$

2. Halogenation – Addition of bromine or other halogens. Halogen alkenes are formed when halogens attach to the carbons in the double bond by covalent bonds *e.g.*

$$CH_2 = CH_2 + Br_2 \rightarrow CH_2BrCH_2Br$$

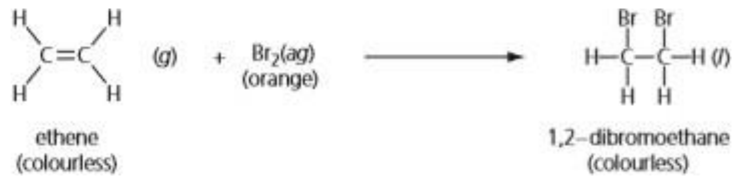
3. Hydration – Addition of water. Alcohols form when water is added to alkene molecules. A catalyst of dilute H_2SO_4 or H_3PO_4 is used.

$$CH_2 = CH_2 + H_2O \rightarrow CH_3–CH_2–OH$$

Number of Carbons	Alkane	Alkene	Alcohol	Carboxylic Acids
1	Methane	---	Methanol	Methanoic Acid
2	Ethane	Ethene	Ethanol	Ethanoic Acid
3	Propane	Propene	Propanol	Propanoic Acid
4	Butane	Butene	Butanol	Butanoic Acid
5	Pentane	Pentene	Pentanol	Pentanoic Acid
6	Hexane	Hexene	Hexanol	Hexanoic Acid
7	Heptane	Heptene	Heptanol	Heptanoic Acid
8	Octane	Octene	Octanol	Octanoic Acid
9	Nonane	Nonene	Nonanol	Nonanoic Acid
10	Decane	Decene	Decanol	Decanoic Acid

A test for unsaturation

You can use bromine water to test whether a hydrocarbon is unsaturated. It is an orange solution of bromine in water. If a C=C bond is present, an addition reaction takes place and the colour disappears. For example:



Alcohol (Ethanol)

Ethanol, an important alcohol Ethanol is the alcohol in alcoholic drinks.

It is a good solvent. It dissolves many substances that do not dissolve in water. It evaporates easily – it is volatile. That makes it a suitable solvent to use in glues, printing inks, perfumes, and aftershave

Two ways to make ethanol

Ethanol is made in two ways, one biological and one chemical.

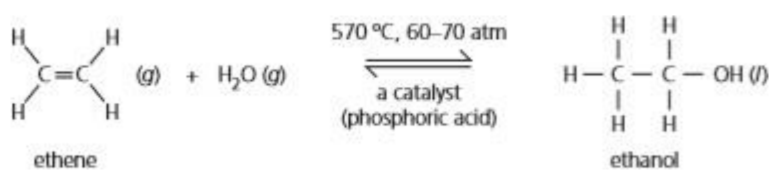
By fermentation – the biological way Ethanol is made from glucose using yeast, in the absence of air:

$C_6H_{12}O_6$ (aq) → $2C_2H_5OH$ (aq) + 2CO2 (g)

Yeast is a mass of living cells. The enzymes in it catalyse the reaction.

The process is called fermentation, and it is exothermic. Ethanol can be made in this way from any substance that contains sugar, starch, or cellulose. (These break down to glucose.) For example it can be made from sugarcane, maize, potatoes, and wood. The yeast stops working when the % of ethanol reaches a certain level, or if the mixture gets too warm. The ethanol is separated from the final mixture by fractional distillation.

By the hydration of ethene – the chemical way Hydration means water is added on. This is an addition reaction.



The reaction is reversible, and exothermic. High pressure and a low temperature would give the best yield. But in practice the reaction is carried out at 570 °C, to give a decent rate of reaction. A catalyst is also used, to speed up the reaction.

Ethanol as a fuel

Ethanol burns well in oxygen, giving out plenty of heat:

C_2H_5OH (l) + $3O_2$ (g) → $2CO_2$ (g) + $3H_2O$ (l) + heat

It is increasingly used as a fuel for car engines because: It can be made quite cheaply from waste plant material. Many countries have no petroleum of their own, and have to buy it from other countries; it costs a lot, so ethanol is an attractive option. Ethanol has less impact on carbon dioxide levels than fossil fuels do.

Carboxylic Acid (Ethanoic Acid)

There are two ways to make ethanoic acid by oxidising ethanol:

By fermentation – the biological way. When ethanol is left standing in air, bacteria bring about its oxidation to ethanoic acid. This method is called acid fermentation. Acid fermentation is used to make vinegar (a dilute solution of ethanoic acid). The vinegar starts as foods such as apples, rice, and honey, which are first fermented to give ethanol.

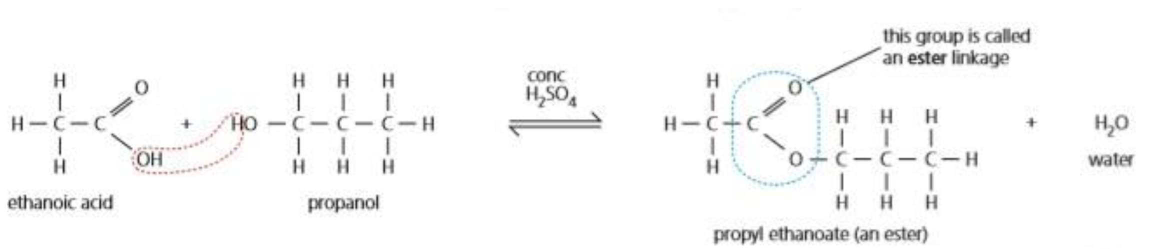
Using oxidising agents – the chemical way. Ethanol is oxidised much faster by warming it with the powerful oxidising agent potassium manganate(VII), in the presence of acid. The manganate(VII) ions are themselves reduced to Mn^{2+} ions, with a colour change. The acid provides the H+ ions for the reaction:

$$\underset{\text{purple}}{MnO_4^-} + 8H^+ + 5e^- \longrightarrow \underset{\text{colourless}}{Mn^{2+}} + 4H_2O$$

Potassium dichromate (VI) could also be used as the oxidising agent.

Esters

Ethanoic acid also reacts with alcohols, to give compounds called esters. The alcohol molecule is reversed below, to help you see what is happening:



Two molecules have joined to make a larger molecule, with the loss of a small molecule, water. So this is called a condensation reaction.

The reaction is reversible, and sulfuric acid acts a catalyst.

The alcohol part comes first in the name – but second in the formula.

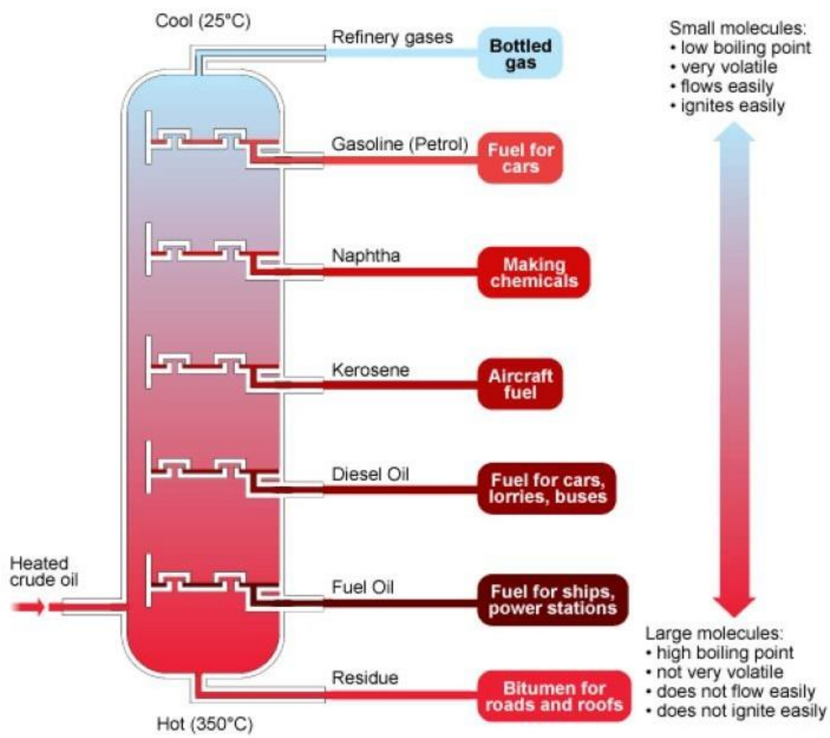
Propyl ethanoate smells of pears. In fact many esters have attractive smells and tastes. So they are added to shampoos and soaps for their smells, and to ice cream and other foods as flavourings.

Crude Oil

Fossil Fuel - organic matter (once living) *e.g. coal (dead plant matter), oil (dead sea creature remains).*

Crude oil - made of hydrocarbons. It is the result of heat and pressure on plant and (sea) animal remains over millions of years in the absence of air. This oil (and gas) rises up through permeable rocks and becomes trapped under impermeable rocks, so they have to be extracted by drilling. The oil is called crude oil because it is unrefined which makes it of little use as it is hard to transport.

Fractional Distillation - process used to separate a mixture of liquids that have different boiling points. When the mixture is heated, liquids with low boiling points evaporate and turn to vapour and can then be separated as liquids. Those with high boiling points remain liquids.



Cracking – allows large hydrocarbon molecules to be broken down into smaller, more useful hydrocarbon molecules. Fractions containing large hydrocarbon molecules are vaporised and passed over a hot catalyst. This breaks chemical bonds in the molecules, and forms smaller hydrocarbon molecules. Cracking is an example of a decomposition reaction. e.g.

$$\underset{\textit{decane}}{CH_3{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_3}$$

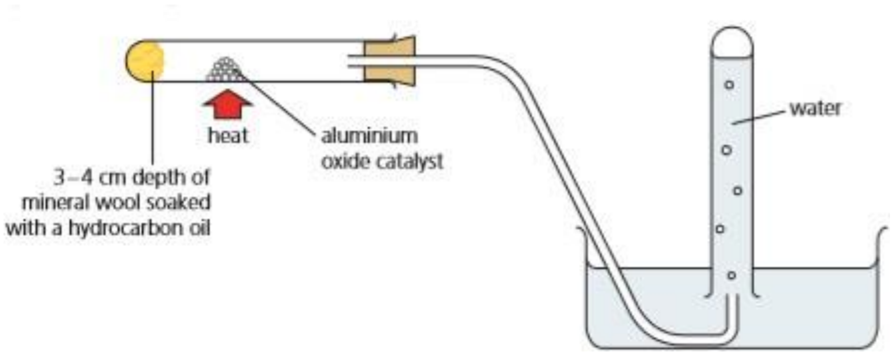
$$\downarrow$$

$$\underset{\textit{propene}}{CH_3{-}CH{=}CH_2} \quad + \quad \underset{\textit{heptane}}{CH_3{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_2{-}CH_3}$$

Note: Cracking of alkane always produces an alkane and an alkene

Cracking a hydrocarbon in the lab

This experiment is carried out using a hydrocarbon oil from petroleum. The product is a gas, collected over water in the inverted test-tube:



The moment heating is stopped, the delivery tube must be lifted out of the water. Otherwise water will get sucked up into the hot test-tube.

Now compare the reactants and products

	The reactant	The product
Appearance	thick colourless liquid	colourless gas
Smell	no smell	pungent smell
Flammability	difficult to burn	burns readily
Reactions	few chemical reactions	many chemical reactions

So the product is quite different from the reactant. Heating has caused the hydrocarbon to break down. A **thermal decomposition** has taken place.

Polymers

Polymers - very large molecules made when hundreds of monomers join together to form long chains. They have no double bonds.

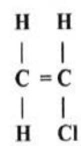
Synthetic Polymers (Plastics) – man-made polymers

Monomers - a molecule that can be bonded to other identical molecules to form a polymer.

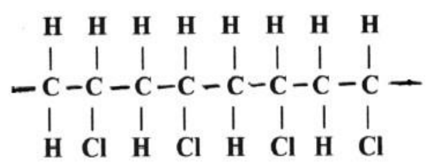
Polymerisation - the combining of monomers to form polymers

Addition Polymers – the monomer is thousands of the same alkene molecules, whose double bond is broken to join the molecules together in one long chain *e.g.*

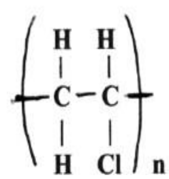
If you had n (number) of this monomer:



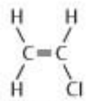
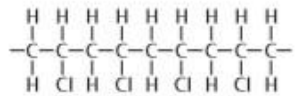
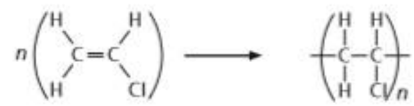
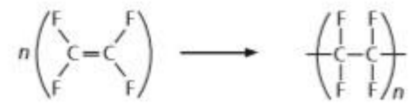
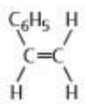
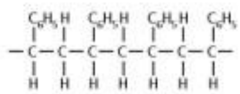
Then n of the monomers would join together, by breaking the double bond and connecting to the other monomers to form a long chain:



This would then be written as a repeated unit:



Making other polymers by addition

The monomer	Part of the polymer molecule	The equation for the reaction
 chloroethene (vinyl chloride)	 poly(chloroethene) or polyvinyl chloride (PVC)	 *n* stands for a large number!
 tetrafluoroethene	 poly(tetrafluoroethene) or Teflon	
 phenylethene (styrene)	 poly(phenylethene) or poly(styrene)	

Condensation polymerisation

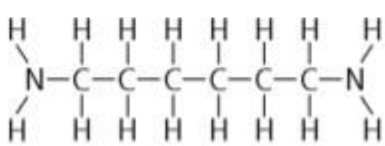
In addition polymerisation, there is only one monomer. Double bonds break, allowing the monomer molecules to join together.

But in condensation polymerisation, no double bonds break. Instead: two different monomers join. Each has two functional groups that take part in the reaction. The monomers join at their functional groups, by getting rid of or eliminating small molecules.

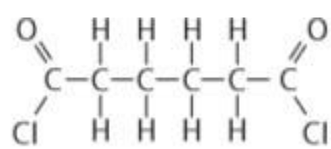
Examples: Nylon, Terylene

Making nylon

Below are the two monomers used in making nylon



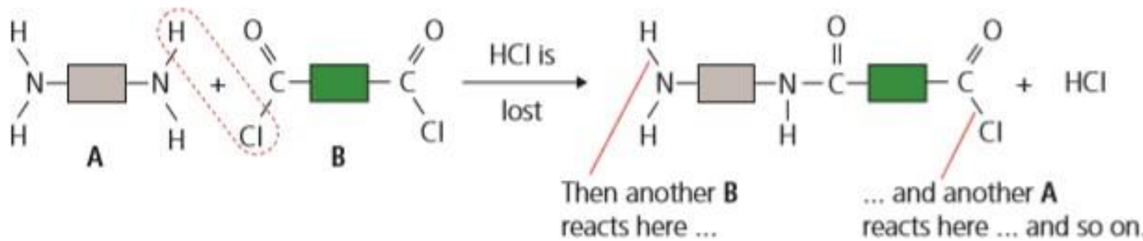
A 1,6-diaminohexane



B hexan-1,6-dioyl chloride

The reaction

This shows the reaction between the two monomer molecules:

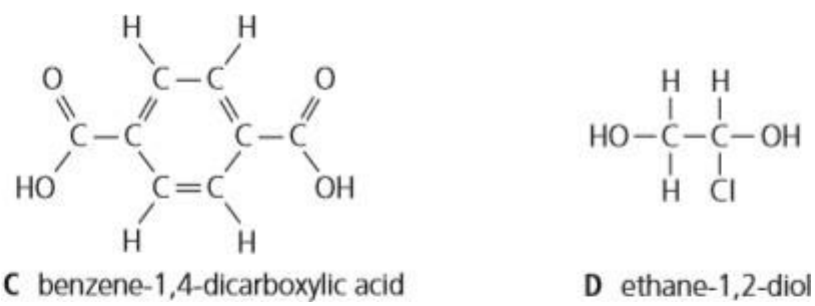


The group where the monomers joined is called the **amide linkage**. So nylon is called a **polyamide**

Nylon can be drawn into tough strong fibres that do not rot away. So it is used for thread, ropes, fishing nets, car seat belts, and carpets.

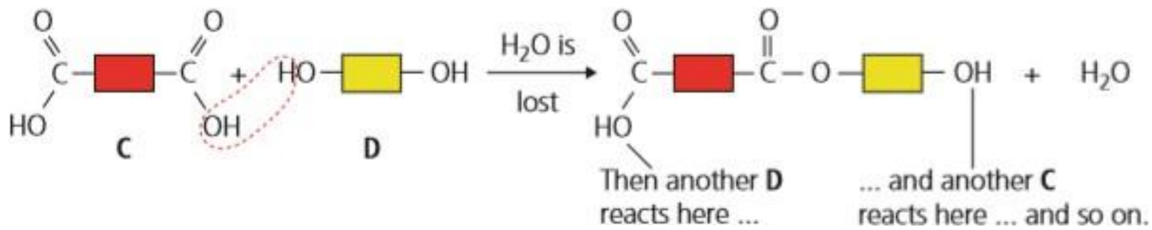
Making Terylene

Terylene is made by condensation polymerisation, using two different monomers.

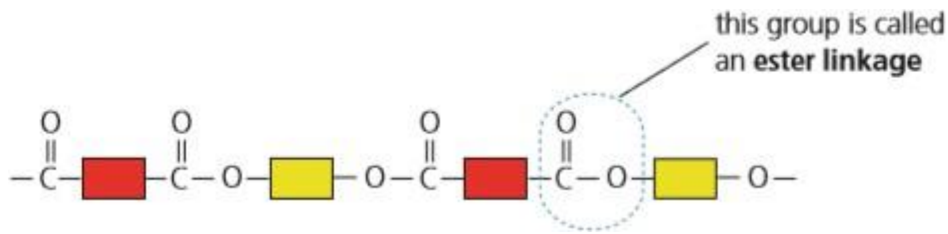


The reaction

This shows the reaction between the two monomer molecules:



In this way thousands of molecules join, giving a macromolecule of Terylene. Here is part of it:



So the group where the monomers have joined is called an **ester linkage**. Terylene is called a **polyester**.

Terylene is used for shirts and other clothing, and for bedlinen. It is usually woven with cotton. The resulting fabric is more hard-wearing than cotton, and does not crease so easily. Terylene is also sold as polyester thread.

Natural Polymers

Polymers produces naturally. Carbohydrates and Proteins

Plants: the polymer factories

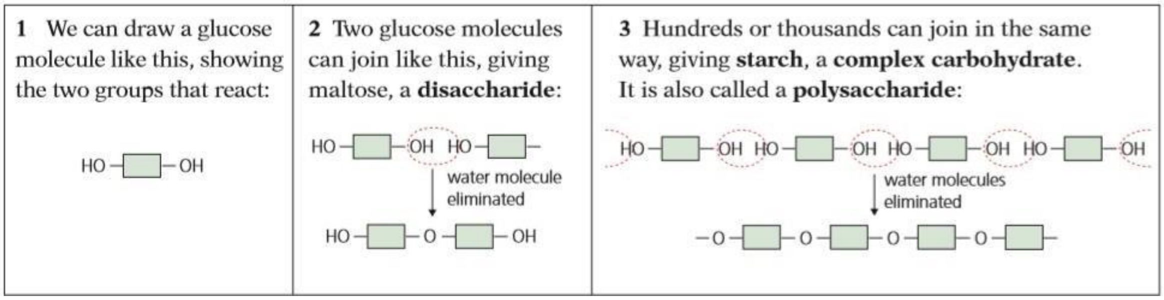
Plants take in carbon dioxide from the air, and water from the soil. 2 Using energy from sunlight, and chlorophyll as a catalyst, they turn them into glucose and oxygen, in a process called photosynthesis:

Then they turn the glucose molecules into macromolecules of starch and cellulose, by polymerisation. These natural polymers are called carbohydrates. Plants use cellulose to build stems and other structures. They use starch as an energy store.

Using glucose, and minerals from the soil, they also produce macromolecules of proteins

Carbohydrates

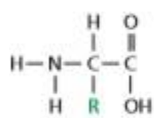
Carbohydrates contain just carbon, hydrogen and oxygen. Glucose is called a simple carbohydrate. It is also called a monosaccharide, which means a single sugar unit.



In reaction 2, two molecules join, eliminating a small molecule (water). So it is a condensation reaction. Reaction 3 is a condensation polymerisation, so starch is a polymer.

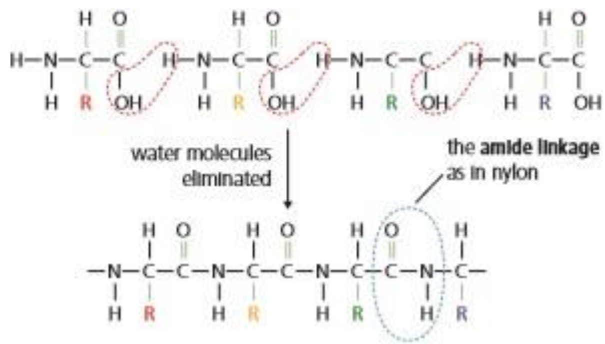
Proteins

Proteins are polymers, built up from molecules of amino acids. Amino acids contain carbon, hydrogen, oxygen, and nitrogen, and some contain sulfur. The general structure of an amino acid molecule is shown on the right. Note the COOH and NH2 functional groups.



▲ An amino acid is a carboxylic acid with an amino (NH_2) group. R stands for the rest of the molecule.

How amino acids join up to make proteins



The reaction is a condensation polymerisation, with loss of water molecules. Note the amide linkage, as in nylon

What happens during digestion?

The natural macromolecules in food were built up by condensation reactions, with the loss of water molecules. The opposite happens when you eat them. In your mouth, stomach and small intestine, the macromolecules are broken down again, by reacting with water. This is called hydrolysis. **Hydrolysis is a reaction in which molecules are broken down by reaction with water.**

Hydrolysis in the digestive system

This is what happens in your body, during digestion:

- **Starch** and any disaccharides get broken down to glucose. Your cells then use the glucose to provide energy, in a process called respiration. It is the reverse of photosynthesis:

$$C_6H_{12}O_6\,(aq) + 6O_2\,(g) \longrightarrow 6CO_2\,(g) + 6H_2O\,(l) + \text{energy}$$

glucose + oxygen ⟶ carbon dioxide + water + energy

- **Proteins** get broken down to **amino acids** which your body then uses to build up the proteins it needs.

Hydrolysis in the lab

You can also carry out hydrolysis of starch, proteins and fats in the lab.
This table shows the conditions, and the results for complete hydrolysis.

Macromolecule	Conditions for the hydrolysis	Complete hydrolysis gives ...
starch	heat with dilute hydrochloric acid	glucose
proteins	boil with 6M hydrochloric acid for 24 hours	amino acids
fats	boil with dilute sodium hydroxide	glycerol plus the sodium salts of the fatty acids ($R-COO^-\ Na^+$)

Practice Questions

(b) Starch is a natural polymer made from glucose.

(i) What type of polymerisation occurs when glucose is converted into starch?

.... [1]

(ii) What type of reaction occurs when starch is converted into glucose?

.... [1]

(iii) Starch can be represented as shown.

Complete the diagram below to represent the structure of the glucose monomer.

[1]

(a) Alkanes and alkenes are two homologous series of hydrocarbons.

(i) What is meant by the term *hydrocarbon*?

.... [1]

(ii) What is the general formula of the homologous series of

alkanes, ...

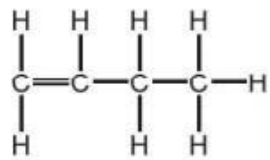
alkenes? ....
[2]

(iii) Other than having a general formula, state **two** characteristics of a homologous series.

1 ....

2 ....
[2]

(iv) The structure of an alkene molecule with the molecular formula C_4H_8 is shown.



Draw the structure of a different alkene molecule with the molecular formula C_4H_8. Show all of the atoms and all of the bonds.

[1]

(v) What term describes molecules with the same molecular formula but different structural formulae?

.... [1]

(d) Dodecane is an alkane containing 12 carbon atoms. Ethanol can be manufactured from dodecane in a two-stage process.

In **stage 1**, each molecule of dodecane is converted into three molecules of ethene and one molecule of another hydrocarbon.

(i) Name the process which occurs in **stage 1**.

.... [1]

(ii) Write a chemical equation for the reaction which occurs in **stage 1**.

.... [2]

In **stage 2**, ethene reacts with steam to produce ethanol.

(iii) State **two** conditions needed for **stage 2**.

1 ....

2 ....

[2]

(iv) Name the type of reaction which occurs in **stage 2**.

.... [1]

(v) Suggest how to test the purity of the ethanol produced.

....

.... [2]

(e) Ethanol can also be manufactured by the fermentation of glucose, $C_6H_{12}O_6$.

(i) State **two** conditions needed for the fermentation of glucose.

1 ....

2 ....
[2]

(ii) Complete the chemical equation for the fermentation of glucose.

$$C_6H_{12}O_6 \rightarrow ......C_2H_5OH + ...$$

[2]

(iii) One disadvantage of fermentation is that the maximum concentration of ethanol produced is about 15%.

Suggest why the concentration of ethanol produced by fermentation does **not** exceed 15%.

....

.... [1]

(iv) Give **one** other disadvantage of manufacturing ethanol by fermentation.

.... [1]

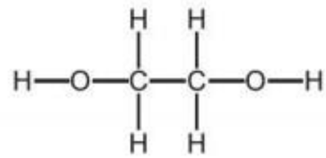
(v) Give **one** advantage, other than cost, of manufacturing ethanol by fermentation.

.... [1]

(vi) Suggest the name of a process to obtain ethanol from a mixture of ethanol and water.

.... [1]

(f) Ethane-1,2-diol has the following structure.



(i) Write the empirical formula of ethane-1,2-diol.

.... [1]

(ii) Ethane-1,2-diol can undergo condensation polymerisation but cannot undergo addition polymerisation.

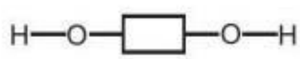
Explain why ethane-1,2-diol **cannot** undergo addition polymerisation.

....

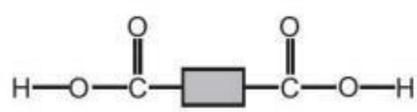
.... [1]

(iii) Ethane-1,2-diol undergoes condensation polymerisation with molecule **Y**.

The diagrams represent the structures of ethane-1,2-diol and molecule **Y**.



ethane-1,2-diol



molecule **Y**

Draw the condensation polymer formed between ethane-1,2-diol and molecule **Y**.
Show **one** repeat unit. Show all of the atoms and all of the bonds in the linkage.

Gasoline is used as a fuel for cars. It is a mixture of hydrocarbons.

(a) Name the raw material from which gasoline is obtained.

.... [1]

(b) One of the compounds in gasoline is heptane, C_7H_{16}. Heptane is a saturated hydrocarbon.

(i) What is meant by the term *saturated hydrocarbon*?

saturated ....

....

hydrocarbon ....

....

[3]

(ii) To which homologous series does heptane belong?

.... [1]

(iii) Give **two** characteristics of an homologous series.

1 ....

2 ....

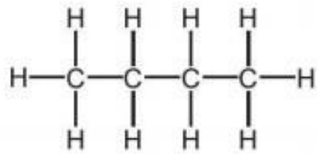
[2]

(iv) Complete the chemical equation for the complete combustion of heptane.

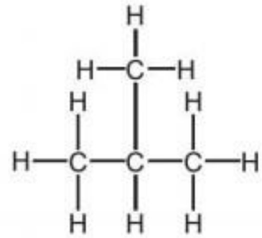
C_7H_{16} + ...O_2 $\rightarrow$ .... + ...

[2]

(d) The formula C_4H_{10} represents two structural isomers, **A** and **B**.



A



B

(i) Name isomer **A**.

.... [1]

(ii) What is meant by the term *structural isomers*?

....

....

.... [2]

(iii) Isomer **B** reacts with chlorine in a substitution reaction.

Give the conditions required for the reaction to occur and draw the structures of **two** possible products, **one** of which is organic and **one** of which is **not** organic.

conditions ....

structures of products

The alkenes and alkanes are both examples of homologous series which are hydrocarbons.

(a) What is meant by the term *hydrocarbon*?

....

.... [2]

(b) Give **three** characteristics of an homologous series.

1 ....

2 ....

3 ....
[3]

(c) Name and draw the structure of the second member of the alkene homologous series.
Show all of the atoms and all of the bonds.

name ....

structure

(d) Alcohols can be made from alkenes.

Name the reagent and conditions needed to convert an alkene into an alcohol.

....

.... [2]

(e) The alcohol butanol, $CH_3CH_2CH_2CH_2OH$, can be converted into a carboxylic acid with four carbon atoms.

(i) Name the carboxylic acid formed from butanol and draw its structure. Show all of the atoms and all of the bonds.

name ....

structure

[2]

(ii) Ethanoic acid can be formed from ethanol by fermentation. It can also be formed by the addition of a suitable chemical reagent.

Name the reagent needed to convert ethanol into ethanoic acid.

.... [2]

.... [2]

(iii) State the type of chemical change which occurs when ethanol is converted into ethanoic acid.

.... [1]

(f) Describe how a student could prepare the ester methyl ethanoate in a school laboratory. In your description give

- the names of the **two** starting organic chemicals,
- the essential reaction conditions needed,
- a chemical equation for the reaction.

....

....

....

....

Synthetic polyamides are made by condensation polymerisation.

(a) **(i)** What is meant by the term *condensation polymerisation*?

....

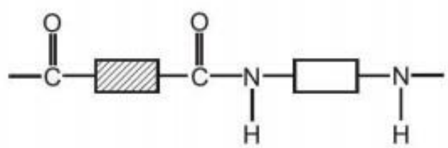
....

.... [3]

(ii) Name another type of polymerisation.

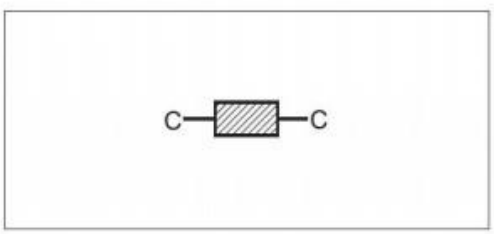
.... [1]

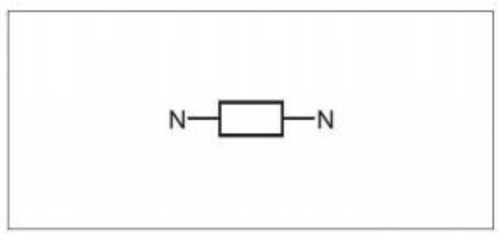
(b) One repeat unit of a synthetic polyamide is represented by the following structure.



(i) Draw a ring around the amide link. [1]

(ii) Complete the diagrams to show the structures of the monomers used to produce the synthetic polyamide. Show all the missing atoms and bonds.





[2]

(iii) Name an example of a synthetic polyamide.

.... [1]

(c) Proteins and synthetic polyamides have similarities and differences.

(i) Name the type of compounds that are the monomers used to make up proteins.

.... [1]

(ii) Starting with a sample of protein, describe how to produce, separate, detect and identify the monomers which make it up.

Your answer should include

- the name of the process used to break down the protein into its monomers,
- the name of the process used to separate the monomers,
- the method used to detect the monomers after they have been separated,
- the method used to identify the monomers after they have been separated.

....

....

....

....

.... [4]

[Total: 13]

Molten zinc chloride can undergo electrolysis. The ions in the electrolyte are Zn^{2+} and Cl^{-}. The electrodes are made of graphite.

a. Give the name of the electrolyte.

____

b. Which electrode do the chloride ions move towards?

____

c. Describe what happens to the chloride ions at this electrode.

____

____

d. Are the zinc ions oxidised or reduced during electrolysis?

____

The products of the electrolysis of potassium bromide solution are hydrogen and bromine. Explain why potassium is not produced.

____

____

____

Molten zinc chloride can undergo electrolysis. The ions in the electrolyte are Zn^{2+} and Cl^{-}. The electrodes are made of graphite.

e. Give the name of the electrolyte.

____

f. Which electrode do the chloride ions move towards?

____

g. Describe what happens to the chloride ions at this electrode.

____

____

h. Are the zinc ions oxidised or reduced during electrolysis?

____

The products of the electrolysis of potassium bromide solution are hydrogen and bromine. Explain why potassium is not produced.

____

____

____

Dan wanted to find out where zinc should be placed in the reactivity series.

(i) What tests should Dan do to find the correct position of zinc in the reactivity series?

....

....

....

1 mark

(ii) How would Dan use his test results to decide where to put zinc in the reactivity series?

....

....

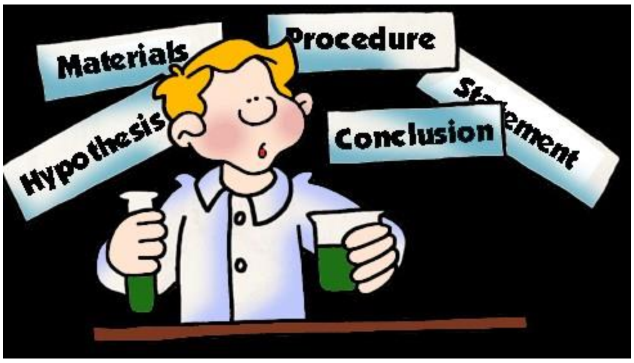
....

1 mark

maximum 5 marks

CRITERIA B AND C

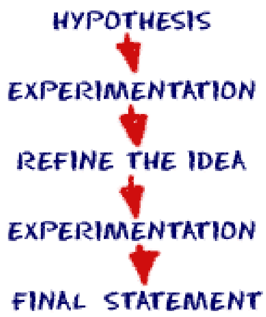
Criterion B - SCIENTIFIC INQUIRY - PLANNING



This process of investigation is often referred to as the **scientific method** and it is defined in many textbooks and science courses as a linear set of steps through which a scientist moves: from observation through experimentation and eventually to a conclusion.

Key Info

The scientific method is a way to ask and answer scientific questions by making observations and doing experiments. The steps of the scientific method are to:



- **Ask a Question**
- **Do Background Research**
- **Construct a Hypothesis**
- **Test Your Hypothesis by Doing an Experiment**
- **Analyze Your Data and Draw a Conclusion**
- **Communicate Your Results**

It is important for your experiment to be a fair test. A "fair test" occurs when you change only one factor (variable) and keep all other conditions the same.

Laboratory Reports.

Lab reports are an important component of the scientific method. Writing up the results of a scientific investigation is one way of communicating your results and conclusions. Other scientists can test your findings by repeating your method. Your investigation should therefore provide all of the necessary information for planning future experiments.

The following information gives you some idea about what you should include in your lab report.

1. RESEARCH QUESTION

Define the problem. Be specific.

Check whether you have:

<u>STATED</u> the purpose of the investigation as an accurate question and <u>INCLUDED</u> SPECIFIC INFORMATION e.g. Times, temperatures, size. The question should include the INDEPENDENT VARIABLE and the DEPENDENT VARIABLE

2. JUSTIFICATION

Why is worthy of study?

A JUSTIFICATION connecting the research question (RQ) directly to the STATEMENT OF INQUIRY (SoI) for that unit. You should start the section by copying the SOI.

Use these 3 headings to help focus your justification:

- Aim of the experiment (this should be short but expand a little on the research question)
- How is your experiment related to real life
- How is your experiment related to scientific research

3. HYPOTHESIS

Predict, and then explain, why you think your prediction is correct using your own experiences and scientific reasoning.

a. How do you think the independent variable will affect the dependent variable?
b. <u>WRITE A</u> HYPOTHESIS WHICH IS CLEAR AND TESTABLE
c. <u>EXPLAIN</u> YOUR HYPOTHESIS USING SCIENTIFIC REASONING (if you use any sources these MUST be listed in Works cited at the end of the investigation but in-text referenced here)

4. VARIABLES

Identify the following variables in your experiment – use this table:

Independent Variable. The **<u>single</u>** factor that you are changing/comparing to test your hypothesis. This should be changed at least 5 times in order to draw a line of best fit/trend line	How will you change this variable
Dependent Variable. What you need to measure to test your hypothesis	How will you measure this variable

Control Variables. What you need to keep the same in your experiment to be a fair test. You should be looking for a minimum of 4-**5** key controlled variables – no major controlled one should be missing	How will you control these variables?

5. APPARATUS

You need to give a list all the important equipment you are going to use in your experiment. This ensures that you remember what to collect and if someone reads your plan they can

see what equipment is required. Try to include the size of the equipment e.g. beaker 250ml, or range for the equipment e.g. thermometer 0-100^{0}C or digital balance +- 0.001g

Apparatus name	Size/type or Range (see above)	Quantity
eg thermometer	0-100°C	2

6. DIAGRAM

For most experiments a diagram is useful. A very easy way to do this is to take a photograph and label the equipment. If not, use a drawing package to draw it neatly. *PS If you are doing an electrical circuit there MUST be a circuit diagram*

7. METHOD

Write out the steps you will follow in your experiment – do it step by step. Say clearly and exactly what you will so that someone else could repeat it exactly.

It should include:

a. What you are going to observe or measure e.g. time, temperature (This is your dependent variable)
b. How you are going to measure or observe it
c. What factor(s) you will have to change during the experiment (This is your independent variable). You will be comparing experiments – what is the key thing you are changing?
d. How you will change your independent variable
e. State the major controlled variables and describe how you will attempt to keep them the same?
f. You must plan for a minimum of three repeat trials

8. RISK ASSESSMENT

You will only be given credit for risks associated with your experiment. If you write down things that will not apply, they will be ignored. Think carefully about what risks are associated with your experiment in terms of making it safe for you and being responsible for the equipment/living organisms. Also consider any ethical considerations.

Rubric - Criterion B: Inquiring and Designing

MYP 5 (Grade 9-10)

<table>
<tr><th>Level</th><th>MYP Level Descriptors (i, ii, iii, iv)</th></tr>
<tr><td>0</td><td>The student does not reach a standard described by any of the descriptors below.</td></tr>
<tr><td>1-2</td><td>The student is able to:
i. state a problem or question to be tested by a scientific investigation
○ The research question is unclear.
○ An attempt is made to justify the question.
ii. outline a testable hypothesis
○ The hypothesis attempts to connect the independent variable to a prediction about the results for the dependent variable but is unclear.
○ There is no valid explanation.
iii. outline the variables
○ The independent and dependent variables are correctly stated.
○ 1-2 valid controlled variables are stated.
iv. design a method, with limited success.
○ A risk assessment is attempted, but may not be relevant.
○ Materials are listed but important information is missing.
○ Steps are very brief and lacking significant detail and clarity.
○ If a diagram was appropriate, it is missing.</td></tr>
<tr><td>3-4</td><td>The student is able to:
i. outline a problem or question to be tested by a scientific investigation
○ The research question mentions either the independent or the dependent variable.
○ The justification paragraph makes a basic but logical connection between the unit's statement of inquiry and the investigation.
ii. formulate a testable hypothesis using scientific reasoning
○ The hypothesis somewhat connects the independent variable to a prediction about the results for the dependent variable but is unclear.
○ The scientific explanation is limited and/or unclear.
iii. outline how to manipulate the variables, and outline how relevant data will be collected
○ The independent and dependent variables are correctly stated but only briefly described, OR one is correctly explained in detail.
○ 2+ valid controlled variables are listed and with at least a brief explanation of how they will be kept constant. (OR 4+ valid controlled variables but without explanation.)
iv. design a safe method in which he or she selects materials and equipment.
○ A relevant risk assessment is made, with an attempt at consideration of ethical and environmental issues.
○ Materials are listed with size and numbers; no important materials are missing.
○ Steps are brief and at times are lacking detail and clarity, but the basic process is present.
○ A diagram is attempted if appropriate.</td></tr>
</table>

5-6	The student is able to: i. describe a problem or question to be tested by a scientific investigation ○ The research **question** is **mostly clear** and mentions both the independent and dependent variables. ○ The justification paragraph makes a **logical connection** between the unit's statement of inquiry and the investigation. ii. formulate and explain a testable hypothesis using scientific reasoning ○ The hypothesis connects the independent variable to a prediction about the results for the dependent variable. ○ It is explained using mostly relevant and correct scientific reasoning. iii. describe how to manipulate the variables, and describe how sufficient, relevant data will be collected ○ The independent **and** dependent variables are **correct** and **described**. *There are **5** different conditions of the independent variable that will be investigated.* ○ **3-4+ valid** controlled variables are listed and a **detailed description** is given as to how they will be kept constant so reliable data can be collected. iv. design a complete and safe method in which he or she selects appropriate materials and equipment. ○ A relevant and **complete** risk assessment is made, with no important safety concern missing, and with **some** consideration of ethical and environmental issues. ○ Materials are listed correctly with **appropriate** size and numbers; nothing **important** is missing. ○ Steps are clear and nothing **important** is missing; however some of the steps may not be necessary or may be slightly out of order. ○ A diagram is **included** if appropriate.
7-8	The student is able to: i. explain a problem or question to be tested by a scientific investigation ○ The research **question** is clear and specifically mentions key details of both the independent and dependent variables. ○ The justification paragraph makes a **detailed and logical connection** between the unit's statement of inquiry and the investigation. ii. formulate and explain a testable hypothesis using correct scientific reasoning ○ The hypothesis **connects** the independent variable to **specific predictions** about the results for the dependent variable. ○ It is explained using **relevant, correct scientific** reasoning. iii. explain how to manipulate the variables, and explain how sufficient, relevant data will be collected ○ The independent **and** dependent variables are **correct** and **explained** in specific detail. *There are **5** different conditions of the independent variable that will be investigated.* ○ **5+ valid** controlled variables are listed and **detailed explanation** is given as to how they will be kept constant so reliable data can be collected. No major control variable is missing. iv. design a logical, complete and safe method in which he or she selects appropriate materials and equipment. ○ A relevant, **complete** and **thorough** risk assessment is made, **with** consideration of ethical and environmental issues. ○ Materials are listed correctly with appropriate size and numbers; **nothing is missing.** ○ Steps are logical, necessary, clear, and in the right order; **nothing is missing.** ○ A **clear and complete** diagram is included if appropriate.

Criterion C -PROCESSING AND EVALUATING

9. RESULTS – data collection, table and graphs

In order to carry out a successful investigation, the following practices should be carried out:

- Each student should prepare their own data tables to record their results.
- Each student should write down accurate measurements or observations as the experiment is carried out. These must be recorded to a appropriate number of significant figures.
- There should be as many different and suitable measurements as possible to reliably test your hypothesis.
- The measurements should be repeated at least three times, if possible.
- Where appropriate, your should record qualitative data.

Example of a table

HEADING: independent variable (with units)

a descriptive title that includes i.v., d.v. and, if appropriate, a key controlled variable

HEADING: dependent variable (with units)

Time taken to dissolve different masses of salt in 100 mL of room temperature water

	Time (s)			
Mass of salt (g)	Trial 1	Trial 2	Trial 3	Average
5	12	10	13	11.7
10	18	19	20	19.0
15	22	22	23	22.3
20	36	38	37	37.0
25	67	70	68	68.3

5 measures for the i.v.

data for three trials present

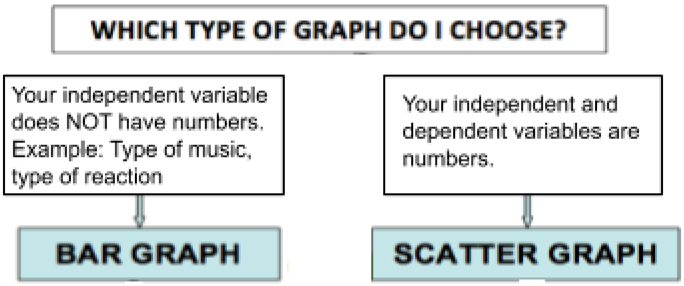
trials have been averaged

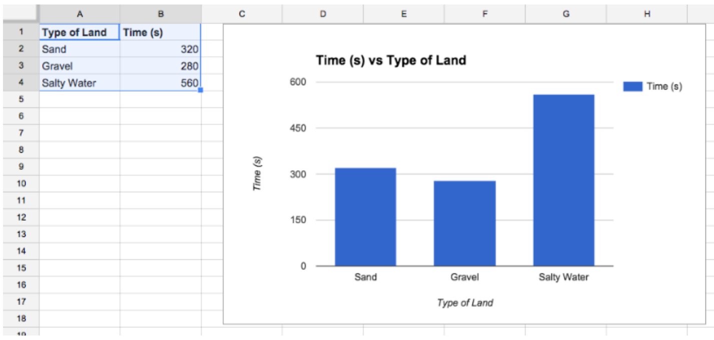
N.B. The data presented in the cells do not have units next to them. Only numbers are presented. Units must be found in the headings.

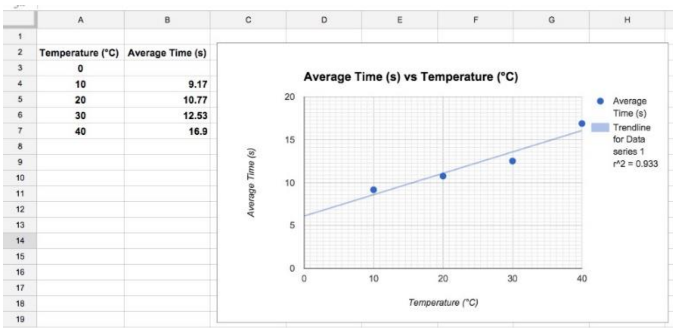
The following are suggestions for processing and presenting your data:

1. Tables for presenting data should be clear and well organized. In addition to an overall title, each column or row should contain a heading and units.
2. Your data should be transformed in some way, including calculating the average of repeated measurements.

3. Present your results in an appropriate graph (SEE BELOW). Graphs should be titled and axes labelled correctly including units. Use a spreadsheet program (click for help).
4. If a scattergraph has been used, a line of best fit/trend line should be drawn and the R^2 value shown.







Graph Checklist

What Makes for a Good Graph?	For a Good Graph, You Should Answer "Yes" to Every Question
Have you selected the appropriate graph for the data you are displaying?	Yes / No
Does your graph have a title?	Yes / No
Have you placed the independent variable on the x-axis and the dependent variable on the y-axis?	Yes / No
Does your graph have the proper scale (the appropriate high and low values on the axes)?	Yes / No
Is your data plotted correctly and clearly? If a scattergraph does your graph have a suitable trend line?	Yes / No
If you have a trend line have you added R^2 value?	Yes / No
Are all major and minor gridlines present?	Yes / No

10. CONCLUSION

The following are suggestions for writing your conclusion:

1. Summarize what you have found out in your investigation – THIS SHOULD BE YOUR OPENING PARAGRAPH.

2. Identify any trends, patterns or relationships in the data and comment on any anomalous results, explaining the reason for the anomaly occurring

3. Explain your conclusion clearly, comparing with your hypothesis and demonstrating an understanding of the scientific reasoning behind your results.

4. Compare your findings with any published data and suggest some reasons for discrepancies that may have occurred.

11. RESPONSE TO HYPOTHESIS

1. Discuss whether or not the data supports your hypothesis.

2. Discuss whether or not the published data supports your hypothesis.

12. EVALUATION

To evaluate your science experiments use the following table:

Limitations of the experiment (you need a minimum of 5 valid limitations for the highest levels)	**Effect on experiment**	**Improvements/Extensions to the Experiment** *Give reasons why they will improve the reliability of evidence collected - state any apparatus that should be used*

FINISH WITH A PARAGRAPH ON RELIABILITY – USE THE INFORMATION ABOVE AND THE PROMPTS BELOW TO HELP YOU WRITE THIS SECTION

- How **RELIABLE** was your method? WERE YOUR REPEATS CONSISTENT?
- Do you think that the data you collected using your method is **VALID**? DID YOU CONTROL YOUR VARIABLES?
- Of the problems listed above, which one(s) significantly affected the reliability of the experiment? Explain.

NOTE: Extension does not mean a different research question or independent variable

Science - Criterion C: Processing and Evaluating

MYP 5 (Grade 9-10)

Level	MYP Level Descriptors (i, ii, iii, iv)
0	The student **does not** reach a standard described by any of the descriptors below.
1-2	The student is able to: i. collect and present data in numerical and/or visual forms • Raw data is collected in a basic **table** but with MANY errors. • Data is presented in a basic **graph** but with MANY errors. ii. interpret data • A **correct** concluding statement is made from the data *(or a concluding statement is made with a description of pa1erns in the result, but it is incorrect).* iii. state the validity of a hypothesis based on the outcome of a scientific investigation • A statement of whether the hypothesis is correct or incorrect is made, WITH • a brief but CORRECT justification referring to the experiment's data **OR** research. iv. state the validity of the method based on the outcome of a scientific investigation • State **at least one valid limitation** of the written method or **valid problem** with the experiment itself, **AND** • a **brief statement** of how that limitation/problem affects the **accuracy** of the method or results. v. state improvements or extensions to the method. • 1+ improvements or extensions to overcome problems/limitations are **stated**, with **at least one** that is **valid.**
3-4	The student is able to: i. correctly collect and present data in numerical and/or visual forms • Raw data is **collected into one or two tables** with **SOME** of the following **correct**: ○ a specific title ○ clearly organized rows and columns ○ headings with units ○ correct, complete data ○ repeat experiments are averaged correctly ○ IF formulae are used, one sample calculation is shown with working ○ specific qualitative observations given IF appropriate • Data is presented in the **correct type of graph** with **SOME** of the following **correct**: ○ a specific title ○ appropriate size (approx. ½ page) ○ appropriate scale (major & minor gridlines) ○ appropriate range on the axes ○ specifically labeled axes with units ○ graph is correct (bars, points, lines) ○ IF scatter graph - there should be a line of best fit and R^2 value (and labeled line of best fit NOT linear) ii. accurately interpret data and explain results • A **correct** concluding statement is made from the data WITH • a basic **description** of pattern(s) in the results, AND • an explanation of the results which may not be fully logical or scientific. iii. outline the validity of a hypothesis based on the outcome of a scientific investigation • A statement of whether the hypothesis is correct or incorrect is made, WITH • a CORRECT justification referring to the experiment's data **OR** research. iv. outline the validity of the method based on the outcome of a scientific investigation • (2-3) **valid limitations** of the written method or **valid problems** with the experiment itself, **AND** • a brief but valid **explanation** of how **each** limitation/problem affects the **validity** of the method and/or **accuracy** of the results v. outline improvements or extensions to the method that would benefit the scientific investigation. • (2-3) clear and realistic improvements or extensions to overcome problems/limitations are given.

5-6	The student is able to: i. correctly collect, organize and present data in numerical and visual forms • Raw data is **collected into one or two <u>tables</u>** with **MOST** of the following **correct**: ○ a specific title ○ clearly organized rows and columns ○ headings with units ○ correct, complete data ○ repeat experiments are averaged correctly ○ IF formulae are used, one sample calculation is shown with working ○ specific qualitative observations given IF appropriate • Data is presented in the **correct type of <u>graph</u>** with **MOST** of the following **correct**: ○ a specific title ○ appropriate size (approx. ½ page) ○ appropriate scale (major & minor gridlines) ○ appropriate range on the axes ○ specifically labeled axes with units ○ graph is correct (bars, points, lines) ○ IF scatter graph - there should be a line of best fit and R^2 value (and labeled line of best fit NOT linear) ii. accurately interpret data and explain results using scientific reasoning • A **correct** conclusion is made from the data WITH • an **<u>explanation</u>** of pattern(s) in the results AND • a **correct** explanation of **why** this happened using **scientific reasoning <u>OR</u>** • a **correct** explanation of whether or not **research** supports your conclusion, using reliable sources. iii. discuss the validity of a hypothesis based on the outcome of a scientific investigation • A statement of whether the hypothesis is correct or incorrect is made, WITH • a CORRECT justification referring to the experiment's data **<u>AND</u>** research. iv. discuss the validity of the method based on the outcome of a scientific investigation • (3-4) **<u>relevant</u> limitations** of the written method or **<u>key</u> problems** with the experiment itself, **<u>AND</u>** • a **<u>clear</u> explanation** of how **<u>each</u>** limitation/problem affects the **validity** of the method or **accuracy** of the results. v. describe improvements or extensions to the method that would benefit the scientific investigation. • (3-4) clear, realistic and **specific improvements or extensions** to overcome problems/limitations are given, WITH • a brief **explanation** of how to this would improve the reliability of the investigation.
7-8	The student is able to: i. correctly collect, organize, transform and present data in numerical and visual forms • Raw data is **collected into one or two <u>tables</u>** with **ALL** of the following **correct**: ○ a specific title ○ clearly organized rows and columns ○ headings with units ○ correct, complete data ○ repeat experiments are averaged correctly ○ IF formulae are used, one sample calculation is shown with working ○ specific qualitative observations given IF appropriate • Data is presented in the **correct type of <u>graph</u>** with **ALL** of the following **correct**: ○ a specific title ○ appropriate size (approx. ½ page) ○ appropriate scale (major & minor gridlines) ○ appropriate range on the axes ○ specifically labeled axes with units ○ graph is correct (bars, points, lines) ○ IF scatter graph - there should be a line of best fit and R^2 value (and labeled line of best fit NOT linear) ii. accurately interpret data and explain results using correct scientific reasoning • A **correct** conclusion is made from the data WITH • an **<u>explanation</u>** of pattern(s) in the results AND • a **correct** explanation of **why** this happened using **scientific reasoning <u>AND</u>** • a **correct** explanation of whether or not **research** supports your conclusion, using reliable sources. iii. evaluate the validity of a hypothesis based on the outcome of a scientific investigation • A statement of whether the hypothesis is correct or incorrect is made, WITH • a CORRECT justification evaluating the strengths and limitations of the hypothesis compared to the experiment's data **<u>AND</u>** research. iv. evaluate the validity of the method based on the outcome of a scientific investigation • (5+) **<u>relevant</u> limitations** of the written method (including any key problems with the experiment itself), **<u>AND</u>** • a **<u>clear</u> explanation** of how **<u>each</u>** limitation/problem affects the **validity** of the method and/or **accuracy** of the results. v. explain improvements or extensions to the method that would benefit the scientific investigation. • (5+) clear, realistic and **specific improvements or extensions** to overcome problems/limitations are given, WITH • an accurate **explanation** of how this would improve the reliability of the investigation.

SCIENCE SAFETY CONTRACT

The scientific laboratory is a place of adventure and discovery. Some of the most important events in scientific history have happened in laboratories. One of the first things any scientist learns is that working in the lab can be an exciting experience. However, the laboratory can also be dangerous if proper safety rules are not followed at all times. In order to prepare for a safe year in the laboratory, read over the following safety rules. Then read them a second time. Make sure you understand each rule. If you do not, ask your teacher to explain any rules of which you are unsure. When you are satisfied that you understand all the rules on this list, sign and date the contract in the place provided. Signing this contract tells your teacher that you are aware of the rules of the laboratory.

Dress Code

1. Many materials in the laboratory can cause eye injury. To protect yourself from possible injury, always wear safety goggles whenever you are working with chemicals, burners, or any substance that might get into your eyes. If you wear contact lenses, let your teacher know.
2. Covered shoes need to be worn TO ALL SCIENCE LESSONS. Even if you are not doing an experiment yourself, there may be broken glass fragments or chemicals from previous lessons around.
3. Tie back long hair in order to keep it away from any chemicals, burners, candles, or other laboratory equipment.
4. Any article of clothing or jewelry that can hang down and touch chemicals and flames should be removed or tied back before working in the laboratory.

General Lab Safety Rules

5. Read all directions for an experiment several times. Follow the directions exactly as they are written. If you are in doubt about any part of the experiment, ask your teacher for assistance.
6. Never perform activities that are not authorized by your teacher. Always obtain permission before "experimenting" on your own.
7. Never handle any equipment unless you have specific permission.
8. Take extreme care not to spill any material in the laboratory. If spills occur, ask your teacher about the proper clean up procedure. Never simply pour chemicals and other substances into the sink or trash container unless you have explicit permission to do so.
9. Never eat in the laboratory. Wash your hands before and after each experiment.

First Aid

10. Report all accidents to your teacher immediately.
11. Learn what to do in case of specific accidents, such as getting acid in your eyes or on your skin (rinse with lots of cool water for 20 minutes).
12. Know where the first aid kit is located.
13. Know where and how to report an accident or fire. Find the location of the fire extinguisher, fire blanket, and fire alarm.

Heating and Fire Safety

14. Never use any heat source, such as candles or a burner, without safety goggles.
15. Never heat any chemical that you are not instructed to heat. A chemical that is harmless when cool can be dangerous when heated.
16. Always maintain a clean work area and keep all materials away from flames.
17. Make sure you know how to light a burner properly. NEVER leave a lighted burner unattended.
18. Always point a test tube or bottle that is being heated away from you and others. Chemicals can splash or boil out of a heated test tube.
19. Never heat a liquid in a closed container. The expanding gases produced may blow the container apart.
20. Never pick up any container that has been heated without first holding your hand above it. If you can feel the heat on your hand, the container may be too hot to handle. Always use a clamp or tongs when handling hot containers or equipment.

Chemical Safety

21. Never mix chemicals for the "fun of it". You might produce a dangerous, possibly explosive substance.
22. Never touch, taste, or smell any chemical that you do not know for a fact is harmless. Many chemicals are poisonous. If you are instructed to smell the fumes in an experiment, always gently wave your hand over the opening of a container and direct the fumes toward your nose. Do not inhale the fumes directly from the container.
23. Use only those chemicals needed in the activity. Keep all lids closed when a chemical is not being used. Notify your teacher

when chemicals are spilled.

24. Dispose of all chemicals as instructed by your teacher. To avoid contamination, never return chemicals to their original containers.
25. Be extra careful when working with acids and bases. Pour such chemicals over the sink, not over your workbench.
26. When diluting an acid, always pour the acid into the water. NEVER pour water into the acid.
27. Rinse any acids off your skin or clothing with water. Immediately notify your teacher of any acid spill.

Glassware Safety

28. Glass tubing should never be forced into a rubber stopper. A turning motion and lubricant will be helpful when inserting glass tubing into rubber stoppers or rubber tubing. Your teacher can demonstrate the proper way to insert glass tubing.
29. Never heat glassware that is not thoroughly dry. Use a wire gauze to protect glassware from the flame of a burner or candle.
30. Keep in mind that hot glassware will not always appear hot. Never pick up glassware without first checking to see if it is hot.
31. Never use broken or chipped glassware. If glassware breaks, notify your teacher. Do not immediately pick it up.
32. Never eat or drink from laboratory glassware. Always thoroughly clean any glassware before putting it away.

Using Sharp Instruments

33. Handle scalpels or razor blades with extreme care. Never cut any material toward you or anyone else—ALWAYS cut away from you.
34. Notify your teacher immediately if you are cut in the laboratory.

End-of-Experiment Rules

35. When an experiment is completed, always clean up your work area and return all equipment to its proper place.
36. Wash your hands after every experiment.
37. Make sure all burners and candles are extinguished before leaving the laboratory.

Other Safety Rules

38. Never run in the laboratory.
39. Never put insoluble solids (such as used matches) in the sinks. Follow directions for the proper disposal of insoluble wastes.
40. Do NOT bring your bags into the laboratory. Bags should be left in lockers
41. Stay at your workstation – do not move around and disrupt other students at their workstations.
42. Work in a quiet, businesslike manner.
43. Do not touch other people’s materials, specimens or equipment.
44. Treat the technicians with respect and consideration.

Criterion D - HELP FOR SPECIFIC COMMAND TERMS

Here is an example for each key command term from the Criterion D rubric, based on the subject of mammals

Command Term	**Question**	**Answer**
State	**State the group of animals that make milk for their babies.**	Mammals
Outline	**Outline the characteristics of mammals.**	Mammals have ● mammary glands ● hair ● live birth **or** Mammals have mammary glands, hair and live birth
Summarize	**Summarize the characteristics of mammals.**	Mammals are a group of animals that have mammary glands that make milk for their offspring. Mammals also have hair generally over their bodies. Also, mammals usually give birth to live young that were developing in the mother.
Describe	**Describe the characteristics of mammals.**	Mammals are a group of animals from the order *mammalia*. They have mammary glands that produce milk for their offspring. An example is the female human who produces milk from her breasts to feed her child. Mammals also have hair which help them keep a constant body temperature.
Explain	**Explain the characteristics of mammals.**	Mammals are a group of animals from the order mammalia which comes from the class chordata. They have mammary glands. All male and female genders have mammary glands however milk is produced by the female after the birth of her young for nutrition. This is called lactation. The milk is produced by changes in hormones such estrogen, progesterone, and prolactin. Mammals also have hair. The function of this hair is to provide warmth, camouflage (ex: polar bear), sensitivity, and protection (ex.: porcupine).
Discuss	**Discuss the benefits and limitations of being a mammal.**	Mammals are a very successful group of animals as their offspring are more likely to survive. One reason for this is because the offspring develop in utero and are protected from predation, temperature changes, and other environmental factors. This is because the milk produced for the mother is a readily available source of energy. Mammals also tend to care for their young for a long period of time. Although this is seen in other types of animals, it is most commonly observed in mammals (in-text citation). A disadvantage of this, is that female mammals require a lot of energy for the rearing of her young and as a result cannot produce as many young as other

		groups of species.
Evaluate	**Evaluate this statement: 'Mammals are the most successful species on Earth.'**	This would be a concluding statement in which you would use the scientific information from your discussion above to decide whether mammals are the most successful species on earth or not - you would have to give arguments for your final decision

EXAMPLES OF INVESTIGATIONS

ˆ Solar Cooker - design a cooker and compare them with other cookers in the class - investigate factors that make the cookers the most efficient.

ˆ The effect of solutes on the boiling point of water or other liquids

ˆ Electrochemistry - look at a factor that affects the voltage produced by an electrochemical cell

ˆ Flexibility of a material

ˆ Which is the best fuel to take on a camping trip?

ˆ Acid Rain investigation - look at the effects of acid rain on different types of concrete

ˆ Investigating Soil

ˆ Use of water in toilets - different toilets could be compared (home v school, energy saving toilets v non energy saving, placing of a brick in cistern, etc.) or even a comparison of the methods used to look for accuracy.

ˆ Fruit batteries

ˆ Electrical conductivity of aqueous solutions of Ionic compounds

ˆ Popcorn Lab - compare bags to see if the numbers are roughly the same. Use an estimation technique - testing the accuracy of the machinery being used to fill the bags

ˆ Creating a sound-proof room - create a soundproof room and compare various recyclable or easy to get hold of materials

IN OMNIA PARATUS
INDUS
INTERNATIONAL SCHOOL

You are going to investigate the rate of reaction between solution S and solution T at different temperatures. When these chemicals react they form iodine. Sodium thiosulfate solution and starch solution can be used to show how fast the reaction proceeds.

Read all the instructions carefully before starting the experiments.

Instructions

You are going to do four experiments.

Experiment 1

● Place the conical flask on the white tile. Use measuring cylinder A to add 10cm3

of solution S

to the conical flask.

● Use measuring cylinder B to add 20cm3

of sodium thiosulfate solution to the conical flask.

● Use the teat pipette to add about 1cm3

of starch solution to the mixture.

● Measure and record the initial temperature of the mixture in the table.

● Use measuring cylinder C to start the reaction by adding 10cm3

of solution T to the conical flask.

Start the timer immediately and swirl the mixture.

● Measure the time taken for the mixture to turn blue-black and record the time taken in the table.

● Measure and record the final temperature of the mixture in the table.

● Empty the conical flask and rinse it with distilled water.

Experiment 2

● Repeat Experiment 1 but first heat the mixture of solution S, sodium thiosulfate solution and starch solution in the conical flask to about 30°C.

● Measure and record the initial temperature of the mixture in the table.

● Use measuring cylinder C to add 10cm3

of solution T to the conical flask. Start the timer

immediately and swirl the mixture.

● Measure the time taken for the mixture to turn blue-black and record the time taken in the table.

● Measure and record the final temperature of the mixture in the table.

● Empty the conical flask and rinse it with distilled water.

Experiment 3

● Repeat Experiment 2 but first heat the mixture of solution S, sodium thiosulfate solution and starch solution in the conical flask to about 40°C before adding solution T.

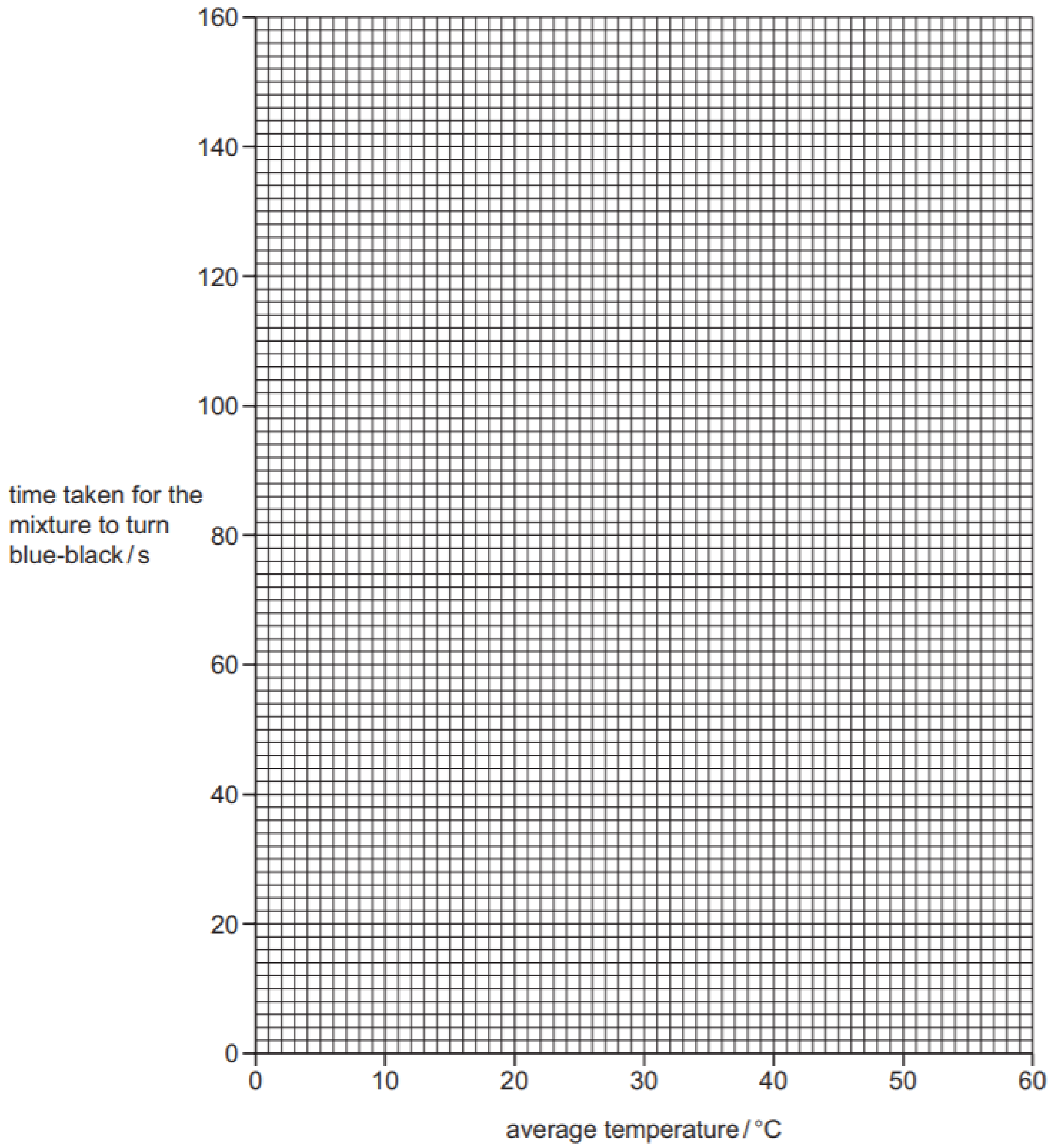
Experiment 4

● Repeat Experiment 2 but first heat the mixture of solution S, sodium thiosulfate solution and starch solution in the conical flask to about 50°C before adding solution T.

(a) Record your results from Experiments 1–4 in the table. Calculate the average temperatures to complete the table.

experiment number	initial temperature / °C	final temperature / °C	average temperature / °C	time taken for the mixture to turn blue-black / s
1				
2				
3				
4				

(b) Plot your results for Experiments 1–4 on the grid. Draw a smooth line graph.



[4]

(c) From your graph, deduce the average temperature needed for the mixture to turn blue-black in 60s. Show clearly on the grid how you worked out your answer.

.... [3]

(d) (i) In which experiment, 1, 2, 3 or 4, was the rate of reaction greatest?

.... [1] (ii) Explain, in terms of particles, why the rate of reaction was greatest in this experiment.

....

....

.... [2] (e) Pipettes or burettes could be used to measure the volumes of solution S and the sodium thiosulfate solution more accurately. State and explain one other way to improve the accuracy of the results of these experiments. way to improve the accuracy

.... explanation

....

.... [2]

(f) A student predicted that using a burette to add solution T would improve the accuracy of the results of these experiments. Suggest why the student's prediction would not improve the accuracy of the results of these experiments.

....

.... [2]

[Total: 18]

2 You are provided with two solids, solid P and solid Q.

Do the following tests on solid P and solid Q, recording all of your observations at each stage.

(a) Describe the appearance of:

solid P ....

solid Q ....

[1]

tests on solid P

Divide solid P into three portions.

(b) (i) Place the first portion of solid P in a boiling tube. Add a small piece of aluminium foil and

about 2cm3

of aqueous sodium hydroxide to the boiling tube. Heat the mixture and test the

gas produced with indicator paper.

Record your observations.

...

...

.. [2]

(ii) Use your results from (b)(i) to identify the gas produced.

.. [1]

(c) (i) Place the second portion of solid P in a test-tube. Add about 2cm3

of distilled water to the

test-tube. Stopper and shake the test-tube to dissolve solid P. Add a few drops of dilute

nitric acid and about 1cm3

of aqueous barium nitrate.

Record your observations.

.. [1]

(ii) What conclusion can you draw about solid P from your observations in (c)(i)?

.. [1]

(d) Do a flame test on the third portion of solid P.

Record your observations.

.. [1]

(e) Identify solid P.

.. [2]

tests on solid Q

(f) Heat about half of solid Q in a hard glass test-tube.

Record your observations.

...

.. [2]

Add about 4cm3

of distilled water to the rest of solid Q in a test-tube. Stopper and shake the

test-tube to dissolve solid Q.

Divide the solution into two equal portions in two test-tubes.

(g) (i) Add an excess of aqueous sodium hydroxide to the first portion of the solution. Leave the

mixture to stand for approximately 5 minutes.

Record your observations.

.. [2]

(ii) Add a few drops of dilute nitric acid and about 1cm3

of aqueous silver nitrate to the second

portion of the solution.

Record your observations.

.... **[1]**

(h) What conclusions can you draw about the identity of solid Q?

....

.... **[2]**

[Total: 16]

3 Some cleaning products are mixtures. The three substances present in a cleaning product are listed in the table.

substance	state at room temperature	physical property
sodium carbonate	solid	melts at 858 °C
ethanol	liquid	boils at 78 °C
limonene	liquid	boils at 176 °C

Use the information in the table to plan an experiment to obtain a sample of each substance from a mixture of the three substances.

You are provided with a mixture of the three substances and common laboratory apparatus.

....
....
....
....
....
....
.... **[6]**

[Total:6]

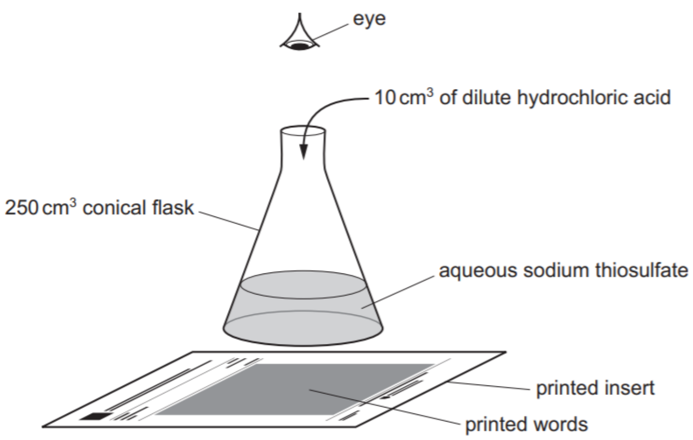
You are going to investigate the rate of reaction between dilute hydrochloric acid and aqueous sodium thiosulfate. When these chemicals react they form a precipitate which makes the solution

go cloudy. The formation of this precipitate can be used to show how fast the reaction proceeds.

Read all the instructions carefully before starting the experiments.

Instructions

You are going to do five experiments using the apparatus shown.



Experiment 1

- **Use the large measuring cylinder to pour 50cm3**

of aqueous sodium thiosulfate into the conical

flask. Place the conical flask on the printed insert.

- **Fill the small measuring cylinder with 10cm3**

of dilute hydrochloric acid.

- **Add the dilute hydrochloric acid to the solution in the conical flask. Start the timer immediately**

and swirl the mixture.

- **View the conical flask from above and measure the time taken for the printed words to disappear**

from view. Record the time taken in the table on page 4.

- **Immediately pour the contents of the conical flask into the quenching bath and rinse the conical**

flask with distilled water.

Experiment 2

- **Use the large measuring cylinder to pour 40cm3**

of aqueous sodium thiosulfate into the conical

flask, followed by 10cm3 of distilled water. Place the conical flask on the printed insert.

- **Fill the small measuring cylinder with 10cm3**

of dilute hydrochloric acid.

- **Add the dilute hydrochloric acid to the solution in the conical flask. Start the timer immediately**

and swirl the mixture.

- **View the conical flask from above and measure the time taken for the printed words to disappear**

from view. Record the time taken in the table on page 4.

- **Immediately pour the contents of the conical flask into the quenching bath and rinse the conical**

flask with distilled water.

Experiment 3

- **Repeat Experiment 2 but use 35cm3 of aqueous sodium thiosulfate, 15cm3**

of distilled water

and 10cm3

of dilute hydrochloric acid.

Experiment 4

- **Repeat Experiment 2 but use 30cm3 of aqueous sodium thiosulfate, 20cm3**

of distilled water

and 10cm3

of dilute hydrochloric acid.

Experiment 5

- **Repeat Experiment 2 but use 10cm3 of aqueous sodium thiosulfate, 40cm3**

of distilled water

and 10cm3

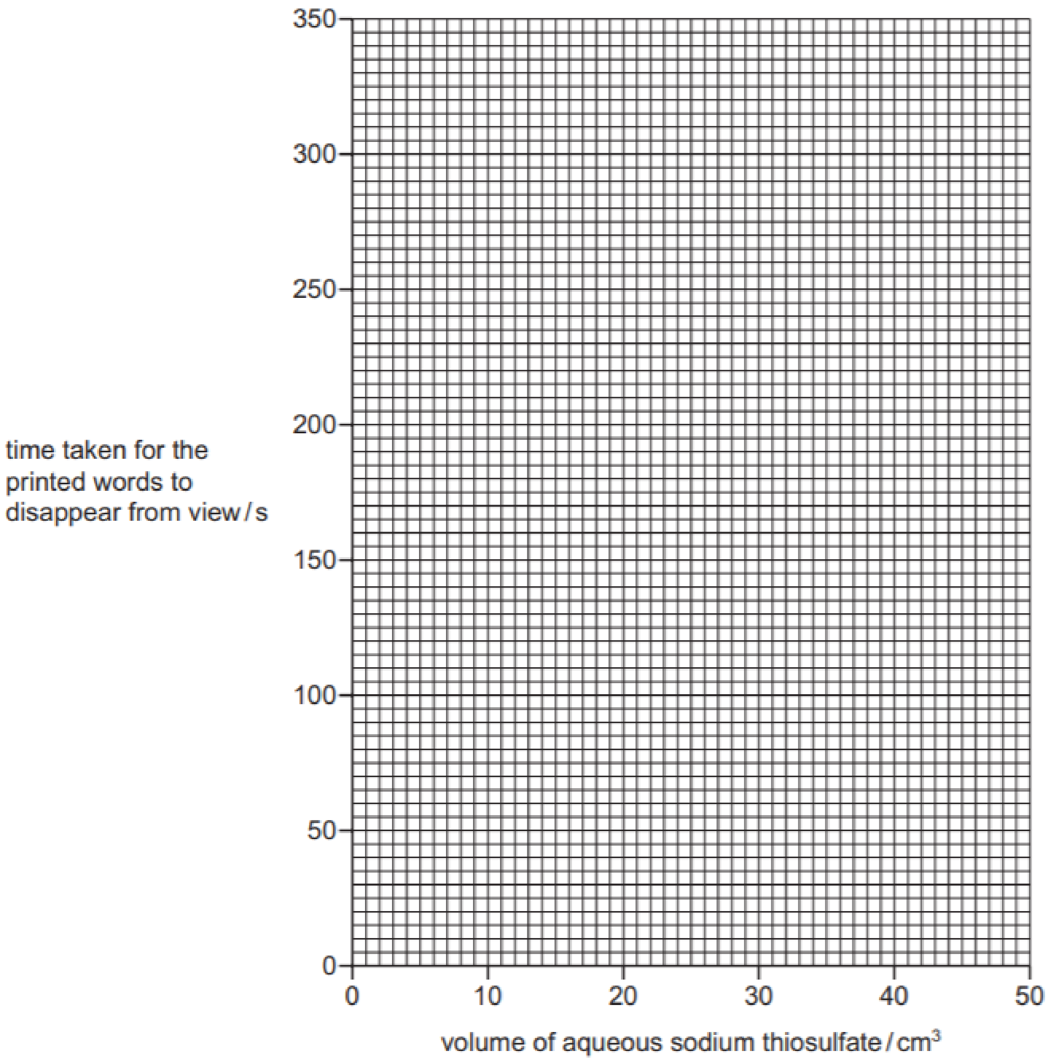
of dilute hydrochloric acid.

(a) Complete the table.

experiment	volume of aqueous sodium thiosulfate / cm^3	volume of distilled water / cm^3	time taken for the printed words to disappear from view / s
1			
2			
3			
4			
5			

[3]

(b) Plot your results from Experiments 1–5 on the grid. Draw a smooth line graph



[3]

(c) Describe the appearance of the mixture in the conical flask at the end of each experiment.

.... [1]

(d) (i) From your graph, deduce the time taken for the printed words to disappear from view

if Experiment 2 were repeated using 20cm3 of aqueous sodium thiosulfate and 30cm3

of distilled water.

Show clearly on the grid how you worked out your answer.

.... s [2]

(ii) The rate of reaction can be calculated using the equation shown.

rate of reaction = 1

time taken

Calculate the rate of reaction using your answer from (d)(i).

.... [1]

(e) (i) In which experiment, 1, 2, 3, 4 or 5, was the rate of reaction greatest?

.... [1]

(ii) Explain, in terms of particles, why the rate of reaction was greatest in this experiment.

....

....

.... [2]

(f) Give the name of a more accurate piece of apparatus for measuring volumes than a measuring cylinder.

.... [1]

(g) Suggest the effect on the results of using a 100cm3 conical flask instead of a 250cm3 conical flask. Explain your answer.

....

.... [2]

(h) Sketch on the grid the graph you would expect if all of the experiments were repeated at a lower temperature. Clearly label your graph. [1]

[Total: 17]

2 You are provided with two substances, solution A and solid B.

Do the following tests on the substances, recording all of your observations at each stage.

tests on solution A

Divide solution A into four approximately equal portions in three test-tubes and one boiling tube.

(a) Test the pH of the first portion of solution A.

pH = ... [1]

(b) Add a strip of magnesium ribbon to the second portion of solution A in a test-tube. Shake the mixture.

Record your observations.

....

.... [1]

(c) Add a spatula measure of copper(II) oxide to the third portion of solution A in a boiling tube.

Warm the mixture gently.

Record your observations.

.... [1]

Keep the fourth portion of solution A for the test in (e).

tests on solid B

(d) Use a spatula to place approximately half of solid B into a hard glass test-tube. Heat solid B

gently then strongly. Leave the hard glass test-tube to stand for approximately 1 minute.

Record your observations.

.... [2]

(e) Add the rest of solid B to the fourth portion of solution A in a test-tube. Test the gas produced.

Record your observations.

test ....

result ....

[2]

Keep the solution from the test in (e) for the test in (f).

Add an approximately equal volume of distilled water to the solution from the test in (e). Shake the

solution and divide it into three approximately equal portions in two test-tubes and one boiling tube.

(f) (i) Add a few drops of aqueous sodium hydroxide to the first portion of the solution in a

test-tube.

Record your observations.

.... [1]

(ii) Now add an excess of aqueous sodium hydroxide to the mixture.

Record your observations.

.... [1]

(g) (i) Add a few drops of aqueous ammonia to the second portion of the solution in a test-tube.

Record your observations.

.... [1]

(ii) Now add an excess of aqueous ammonia to the mixture.

Record your observations.

.... [1]

(h) Add a small piece of aluminium foil and about 2cm3 of aqueous sodium hydroxide to the third portion of the solution in a boiling tube. Gently warm the mixture. Test the gas produced with indicator paper.

Record your observations.

.... [2]

(i) Identify solution A.

.... [2]

(j) Identify solid B.

.... [2]

[Total: 17]

Potassium chloride is a salt that dissolves in water.

The solubility of a salt is the mass in grams of the salt that dissolves in 100cm3 of water at a particular temperature.

Plan an investigation to determine the solubility of potassium chloride in water at 40°C.

You are provided with potassium chloride and common laboratory apparatus.

....

....

....

....

....

....

....

....

.... [6]

[Total: 6]

You are going to investigate what happens to the temperature when two different solids, W and X, dissolve in water.

Read all the instructions carefully before starting the experiments.

Instructions

You are going to carry out two experiments.

(a) Experiment 1

- Put the polystyrene cup into the 250cm3

beaker for support.

- Use the measuring cylinder to pour 30cm3

of distilled water into the polystyrene cup.

- Measure the initial temperature of the water and record it in the table at time = 0 seconds.

- Add all of solid W to the water, start the timer and stir the solution continuously with the

thermometer.

- Measure the temperature of the solution every 10 seconds for 90 seconds.

- Record your results in the table.

time / s	0	10	20	30	40	50	60	70	80	90
temperature of the solution / °C										

[2]

(b) Experiment 2

- Empty the polystyrene cup and rinse it with water. Put the polystyrene cup back into the

250cm3

beaker.

- Use a measuring cylinder to pour 30cm3

of distilled water into the polystyrene cup.

- Measure the initial temperature of the water and record it in the table at time = 0 seconds.

- Add all of solid X to the water, start the timer and stir the solution continuously with the

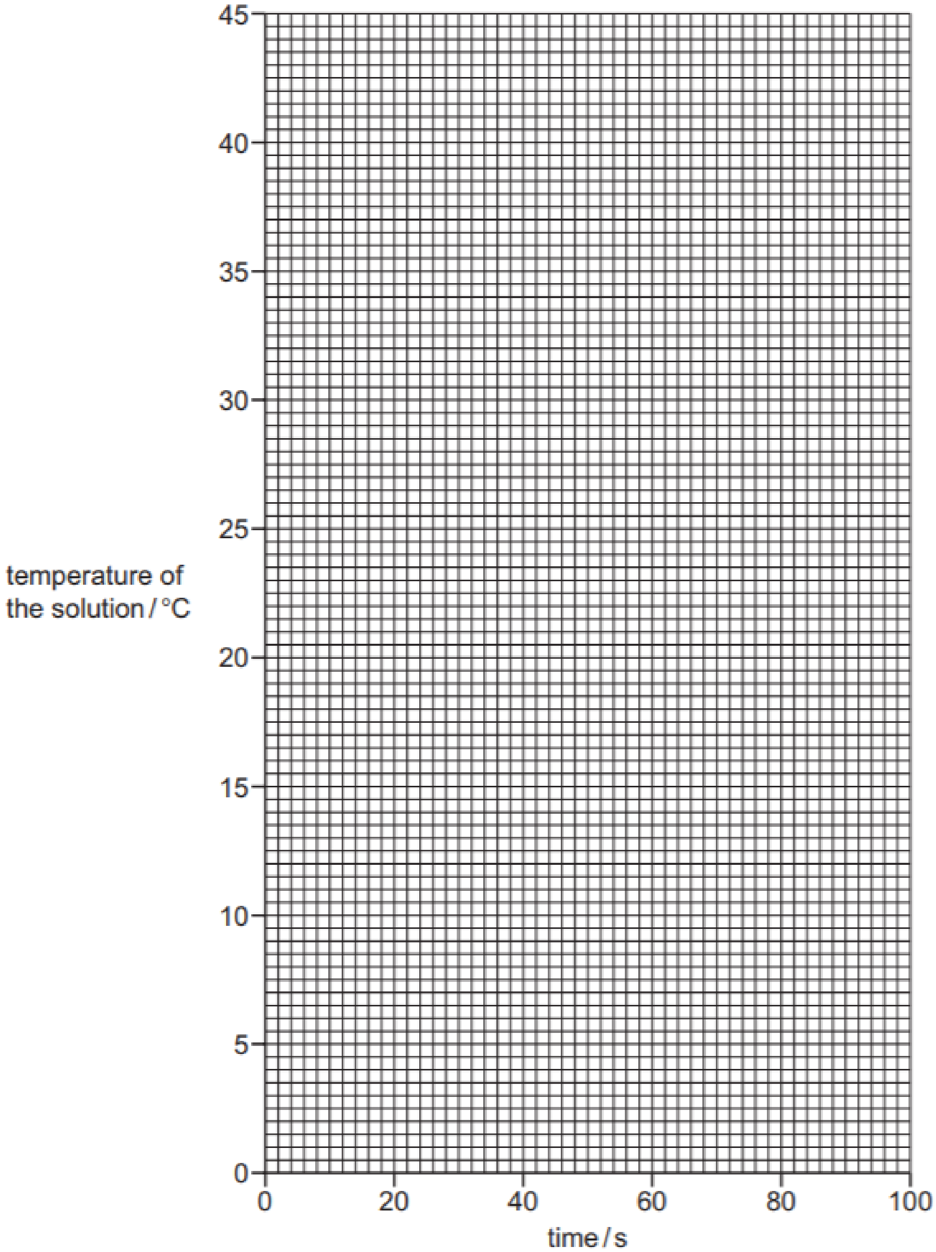
thermometer.

- Measure the temperature of the solution every 10 seconds for 90 seconds.
- Record your results in the table

time / s	0	10	20	30	40	50	60	70	80	90
temperature of the solution / °C										

[2]

(c) Plot your results for Experiments 1 and 2 on the grid. Draw two smooth line graphs. Clearly label your lines.



[4]

(d) (i) From your graph, deduce the temperature of the solution in Experiment 1 after 15 seconds.

Show clearly on the grid how you worked out your answer.

.... °C [2]

(ii) From your graph, deduce the time taken for the temperature of the solution in Experiment 2 to change by 6°C from the initial temperature.

Show clearly on the grid how you worked out your answer.

.... s [2]

(e) Use your results to identify the type of energy change that occurs when solid X dissolves in water.

.... [1]

(f) Predict the temperature of the solution in Experiment 2 after 1 hour. Explain your answer.

....

.... [1]

(g) State two sources of error in these experiments. Give one improvement to reduce each of these sources of error.

source of error 1 ....

improvement 1 ....

source of error 2 ....

improvement 2 ....

[4]

(h) When carrying out the experiments, what would be a disadvantage of taking the temperature readings only every 30 seconds?

....

.... [1]

[Total: 19]

Washing soda crystals are crystals of hydrated sodium carbonate, $Na_2CO_3.10H_2O$. When exposed to the air, some of the water is lost from the crystals and a new substance is formed. This process occurs faster in hotter climates.

Plan an experiment to determine the percentage of water by mass present in the new substance.

You are provided with common laboratory apparatus.

....

....

....

....

....

....

....

....

.... [6]

[Total: 6]

You are going to investigate the reactions of four different metals. Copper, magnesium, iron and zinc will be used.

Read all the instructions below carefully before starting the experiments.

Instructions

Experiment 1

By using a measuring cylinder, pour 15 cm3 of dilute sulphuric acid into the boiling tube provided. Measure the initial temperature of the acid and record it in the table. Add the 1 g sample of zinc powder to the acid in the boiling tube and stir the mixture with the thermometer. Record the maximum temperature reached and any observations in the table. Remove the thermometer and rinse with water.

Experiment 2

Repeat Experiment 1, using 1 g of iron instead of zinc. Record the maximum temperature reached and any observations in the table.

Experiment 3

Repeat Experiment 1, using the 0.5 g sample of magnesium. Test the gas given off with a lighted splint.

Experiment 4

Repeat Experiment 1, using 1 g of copper. Record all results in the table.

Table of results

Experiment	Metal	Temperature of acid / °C		Observations
		initial	maximum	
1	zinc			
2	iron			
3	magnesium			
4	copper			

[10]

(a) Use your results and observations to answer the following questions.

(i) Which metal is most reactive with sulphuric acid?

....[1]

(ii) Give two reasons why you chose this metal.

1. ....

2. ....[2]

(iii) Name the gas given off in Experiment 3.

....[1]

You are now going to investigate the reaction between magnesium and aqueous copper(II) sulphate.

Experiment 5

Rinse the thermometer with water at room temperature. By using a measuring cylinder pour 5 cm3 of aqueous copper(II) sulphate into a test-tube. Measure the initial temperature of the solution and record it in the table below.

Add the 0.2 g sample of magnesium powder to the test-tube and record the maximum temperature reached. Record all of your observations in the table.

Table of results

initial temperature of aqueous copper(II) sulphate.. °C

maximum temperature reached when magnesium added .... °C
[2]

observations when magnesium was added to aqueous copper(II) sulphate

....

....

....[3]

(b) How do your observations show that the reaction of magnesium with aqueous copper(II) sulphate is exothermic?

....[1]

(c) What type of exothermic reaction occurs when magnesium is added to aqueous copper(II) sulphate?

....[1]

(d) Use your results from Experiments 1 to 5 to put the four metals in order of reactivity.

least reactive...

....

....

most reactive...

IN OMNIA PARATUS
INDUS
INTERNATIONAL SCHOOL

CRITERIA D

POSSIBLE CRITERION D IDEAS – REFLECTING ON THE IMPACT OF SCIENCE

CHEMICAL ENERGY

- ❒ Alternative/renewable energy/fuels of the Future – solutions to the impact of fossil fuels
- ❒ Climate change – students choose a specific country and then look at solutions to limit climate change effects in that country (IDU with Individuals and Societies)
- ❒ The future of transportation

CHEMISTRY AND CHEMICAL REACTIONS

- ❒ Industrial Chemistry - students to choose one industrially made chemical and look at its impact on society
- ❒ The use of Batteries
- ❒ The use of plastics
- ❒ Natural v Synthetic (Textiles, Pesticides)
- ❒ Recycling metals

ELECTRICITY

- ❒ Electrical devices – what was their invention a solution for?
- ❒ Energy conservation
- ❒ Renewable energy – solar, hydro-electric, wind, tidal, etc.

HEALTH AND DISEASE

- ❒ Controlling Diseases
- ❒ Medical advances – good for Living Systems units

OTHERS

- ❒ Should countries try to put a human on Mars?
- ❒ Crop production techniques – maximizing Photosynthesis
- ❒ Famous Scientist – what effect has their discovery had on science?
- ❒ What is the role of science in preventing/predicting/limiting the impact of natural disasters?

ECOLOGY

- ❒ Ecological Footprints - complete an Ecological Footprint and then write an essay to look at the impact of this on society
- ❒ Deforestation – how this could be prevented?

GUIDANCE FOR WRITING ABOUT FACTORS

Think about these definitions/questions when you evaluate the benefits and limitations of this application of science in regards to one world factors.

ECONOMIC is defined as "pertaining to the production, distribution, and use of income, wealth, and commodities." The costs and benefits of using science. Who pays or receives the benefit? Is it worth it? How much will it cost? Are the potential benefits worth the cost? Who pays for it, private companies, the Government, consumers? (Obesity will cost governments millions but so does changing people's eating habits and diet- is it worth attempting to deal with it?)

SOCIAL is defined as "of or pertaining to the life, welfare, and relations of human beings in a community."

The people affected or involved, from a few people up to the global community. How will this application of science affect society, will it affect everyone or a select group? Will it unfairly affect one group of society over another? Will everyone have access to the science/solution or only the privileged (An oil refinery is a local problem but global warming is truly global)

ETHICAL is defined as "pertaining to or dealing with morals or the principles of morality; pertaining to right and wrong in conduct." Is it right or wrong to use science in this way and if we do, how ought we to do it? (Aborting foetuses is something we have to do in a caring way even if we think it is right or wrong. Who decides whether it is right or wrong, the Government, religious groups, individuals? There are ways of doing things and often a law to guide us - animal testing is highly policed).

POLITICAL is defined as "of, pertaining to, or involving the state or its government." Is the government involved directly or are powerful groups trying to influence people or the government or the UN etc? Notice how much diet advice is given in books by non-experts. Why would the Government be for or against this the application of this science and how would it affect other countries?

ENVIRONMENTAL is defined as "of, relating to, or associated with the environment." Is the issue an environmental one in some way, directly or indirectly? (Pollution is direct but what about binge-drinking?)

CULTURAL is defined as "of, relating to, or associated with the behaviors and beliefs characteristic of a particular social, ethnic, or age group." A problem in one place is seen differently in another for reasons such as faith, tradition and different belief values. Would people from different cultures be affected the same way or have the same opinion on the scientific solution. Some religions are against certain medical procedures, from organ transplants to blood donations. Some cultures rely heavilly upon the natural ecosystem to gain their food and shelter, others live in a manufactured society and are diconnected from the natural world. Certain religions have rules about how males and females should interact, or what foods to eat.

Resources for self – study and content development

https://znotes.org/cie-igcse/chemistry-0620

https://www.savemyexams.co.uk/igcse-chemistry-cie-new/revision-notes/

https://www.physicsandmathstutor.com/chemistry-revision/igcse-cie/

http://astarchemistry.com/igcse/cie/notes/

https://igcsesciencenotes.wordpress.com/chemistry0620/

https://ibrecap.com/

https://www.chemactive.com/wp-content/uploads/2018/03/MYP-Chemistry-Notes.pdf

ALL THE BEST!!