Higher Chemistry

Unit 1 – Chemical Changes and Structure

Homework Booklet

Kirkcaldy High School

Controlling the Rate of Reaction

1. A pupil made the following observations on dripping taps. $45cm^3$ of water was collected from Tap A in 3 minutes. 340 cm^3 of water was collected from Tap B in 20 minutes. By calculating the average rate of loss of water from each tap, find out which tap was dripping faster.

Tap A: rate = $\dfrac{\textbf{change in quantity}}{\textbf{change in time}} = \dfrac{45}{3} = 15\ cm^3min^{-1}$

Tap B: rate = $\dfrac{\textbf{change in quantity}}{\textbf{change in time}} = \dfrac{340}{20} = 17\ cm^3min^{-1}$ **- faster**

2. A farmer records the weight of his pigs every Monday. Here is part of the record for one of the pigs.

Date	Weight (kg)	Date	Weight (kg)
Jan 1	75.85	Jan 22	83.70
Jan 8	76.50	Jan 29	86.90
Jan 15	79.10	Feb 5	92.10

Calculate the average rate of weight gain per week:-

i) during the first week

- **rate =** $\dfrac{\textbf{change in quantity}}{\textbf{change in time}} = \dfrac{76.50\text{-}75.85}{1} = 0.65\ kg\ week^{-1}$

ii) during the first three weeks

- **rate =** $\dfrac{\textbf{change in quantity}}{\textbf{change in time}} = \dfrac{83.70\text{-}75.85}{3} = 2.61\ kg\ week^{-1}$

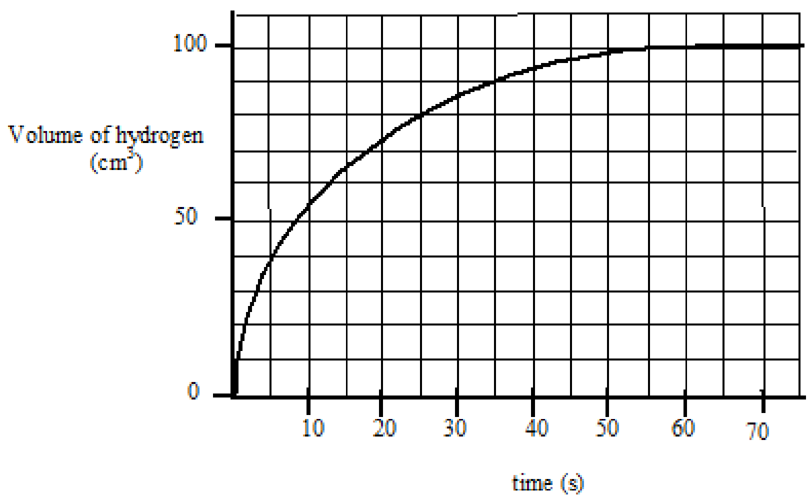
iii) over the five week period

- **rate =** $\dfrac{\textbf{change in quantity}}{\textbf{change in time}} = \dfrac{92.10.50\text{-}75.85}{5} = 3.25\ kg\ week^{-1}$

3. A pupil was attempting to measure the rate of a chemical reaction which produced a gas. After six seconds $8cm^3$ of gas had been collected. After ten seconds the total volume of gas collected was $14cm^3$. Calculate the average rate of the reaction during this time interval (from six to ten seconds).

- **rate =** $\dfrac{\textbf{change in quantity}}{\textbf{change in time}} = \dfrac{14\text{-}8}{10\text{-}6} = 1.5\ cm^3s^{-1}$

4. The graph below shows the volume of hydrogen gas released when a 10cm strip of magnesium (mass = 0.1g) was added to $30cm^3$ of 1 mol l^{-1} hydrochloric acid.



(a) Calculate the average rate of reaction

i. over the first 15 seconds

- **rate = $\frac{\text{change in quantity}}{\text{change in time}} = \frac{65-0}{15-0}$ = 4.3 cm^3s^{-1}**

ii. between 20 and 30 seconds

- **rate = $\frac{\text{change in quantity}}{\text{change in time}} = \frac{85-72}{30-20}$ = 1.3 cm^3s^{-1}**

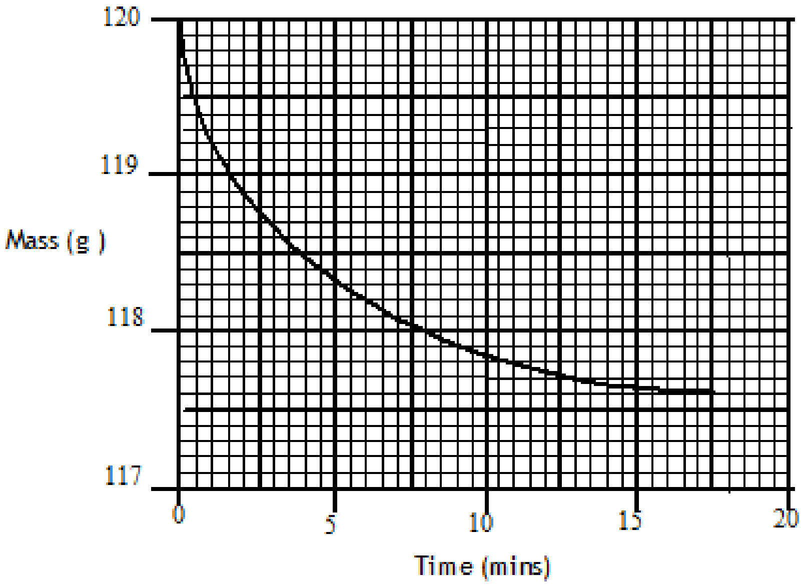
(b) How long did it take for the reaction to stop?

- **60 s**

(c) The graph shows that the rate of reaction changes as the reaction proceeds. Explain why it changes in this way.

- **The concentration of the acid is dropping so the reaction rate drops**

5. Marble chips (calcium carbonate), reacted with excess dilute hydrochloric acid.
The rate of reaction was followed by recording the mass of the container and the reaction mixture over a period of time. The results of the experiment are shown in the following graph.



(a) Write a balanced equation for the reaction.

- **$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$**

(b) Give a reason for the loss of mass of the container.

- **A gas is given off**

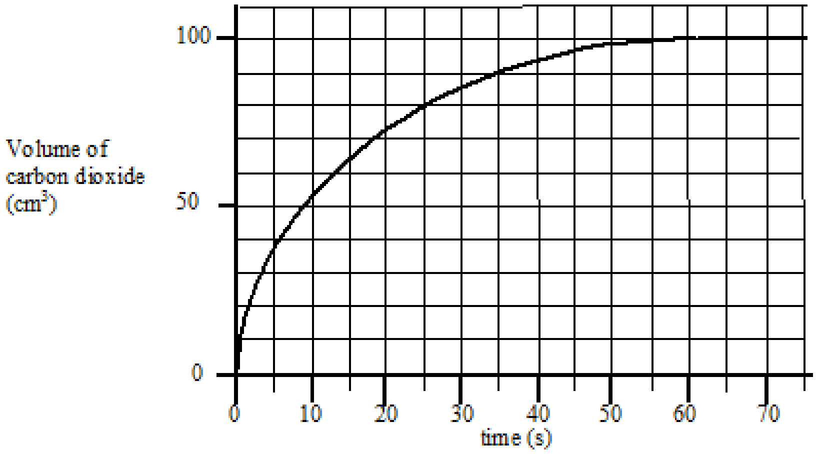
(c) Calculate the average rate of reaction over the first five minutes.

- **rate = $\frac{\text{change in quantity}}{\text{change in time}} = \frac{120-118.3}{5-0}$ = 0.3 cm^3min^{-1}**

(d) Why does the average rate of reaction decrease as the reaction proceeds?

- **The concentration of the acid is dropping so the reaction rate drops**

6. The results shown below were obtained when 0.42 g of powdered chalk was added to 20 cm^3 hydrochloric acid, concentration 2 mol l^{-1} (an excess of the acid).



(a) Sketch the graph and add a **solid** line to the graph to show what would happen if 0.42g of chalk lumps was used instead of powdered chalk.

- **A higher line that starts and ends at the same place as the original**

(b) Add a **dotted** line to the graph to show what would happen if 20 cm^3 of 3 mol l^{-1} hydrochloric acid was used instead of 2 mol l^{-1} hydrochloric acid.

- **A higher line that starts and ends at the same place as the original**

7. 1.0 g of zinc was placed in 20 cm^3 of 2 mol l^{-1} hydrochloric acid. After 20 seconds the zinc was removed, washed, dried and re-weighed. The remaining zinc weighed 0.35 g.

(a) Write a balanced chemical equation for the reaction.

- **$Zn + 2HCl \rightarrow ZnCl_2 + H_2$**

(b) Calculate the average rate of the reaction.

- **rate = $\frac{\text{change in quantity}}{\text{change in time}} = \frac{1-0.35}{20-0} = 0.0325\ cm^3s^{-1}$**

(c) Calculate the number of moles of hydrochloric acid used up in the 20 seconds.

- **$Zn \quad + \quad 2HCl \quad \rightarrow \quad ZnCl_2 \quad + \quad H_2$**
- **No. moles Zinc, n = $\frac{m}{GFM} = \frac{0.65}{65.5} = 0.01$ moles**
- **No. moles HCl, n = No. moles zinc x 2 = 0.01 x 2 = 0.02 moles**

8. A pupil was investigating the effect of temperature on the rate of a chemical reaction and obtained the following data.

Temperature (^oC)	15	25	33	37	44
Time for reaction to finish (s)	154	66.7	40	30.3	22.2
Relative rate (**s^{-1}**)	**0.006**	**0.015**	**0.025**	**0.033**	**0.045**

(a) Copy the table and calculate the relative rate of reaction at each temperature and add them to the table, putting the correct units in the brackets.

(b) Plot a graph of relative rate against temperature.

- **Correct graph with labels, units, $\frac{1}{t}$ vs T**

(c) Predict what the **relative rate** of the reaction will be at 50 ^0C.

- **0.06 s^{-1}**

(d) Use the graph to estimate the **time** for the reaction to finish at 40 ^0C.

- **35 s^{-1}**

9. (a) Explain

i) Why decreasing the particle size increases the rate of a chemical reaction.

- **The surface area for reaction increases**

ii) Why increasing the concentration speeds up a chemical reaction.

- **The number of particles increases so the probability of collisions increases**

(b) Give an everyday example of a reaction speeded up by

i) decreasing the particle size

- **Cutting up a steak so it cooks faster**

ii) increasing the concentration

- **Blowing on a fire**

10. When hydrochloric acid is added to a solution of sodium thiosulphate the following reaction takes place.

$$2HCl(aq) + Na_2S_2O_3(aq) \rightarrow 2NaCl(aq) + SO_2(g) + S(s) + H_2O(l)$$

Solid sulphur forms in the solution.

In one set of experiments the effect of varying the concentration of sodium thiosulphate was studied. Some of the volumes of solutions used are shown.

Volume of 0.05 mol l^{-1} $Na_2S_2O_3(aq)$/ cm^3	Volume of water/ cm^3	Volume of 0.1 mol l^{-1} $HCl(aq)$/ cm^3	Reaction time /s
200	0	5	20
160	**40**	**5**	25
120	**80**	**5**	33
80	**120**	**5**	50
40	**160**	**5**	100

(a) Copy and then complete the table to show the volumes of water and acid that would have been used.

(b) Describe how the reaction time could have been measured.

- **Time it takes for a cross placed underneath to disappear**

(c) Describe how the relative rate of reaction would be obtained from each of the results.

- **Relative rate =** $\mathbf{\frac{1}{t}}$

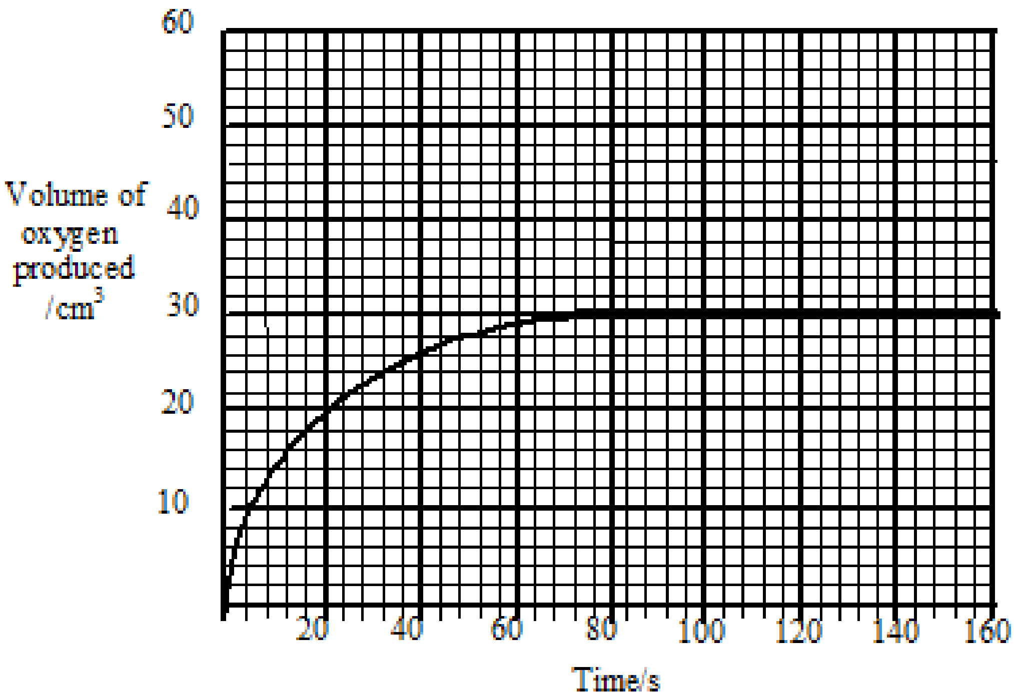
11. Hydrogen peroxide can be used to clean contact lenses. In this process, the enzyme catalase is added to break down hydrogen peroxide. The equation for the reaction is:

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

The rate of oxygen production was measured in three laboratory experiments using the same volume of hydrogen peroxide at the same temperature.

Experiment	Concentration of H_2O_2/ mol l^{-1}	Catalyst used
A	0.2	yes
B	0.4	yes
C	0.2	no

The curve obtained for experiment **A** is shown.



(a) Calculate the average rate of the reaction over the first 40 s.

- **rate = $\dfrac{\text{change in quantity}}{\text{change in time}} = \dfrac{26-0}{40-0} = 0.65\ cm^3s^{-1}$**

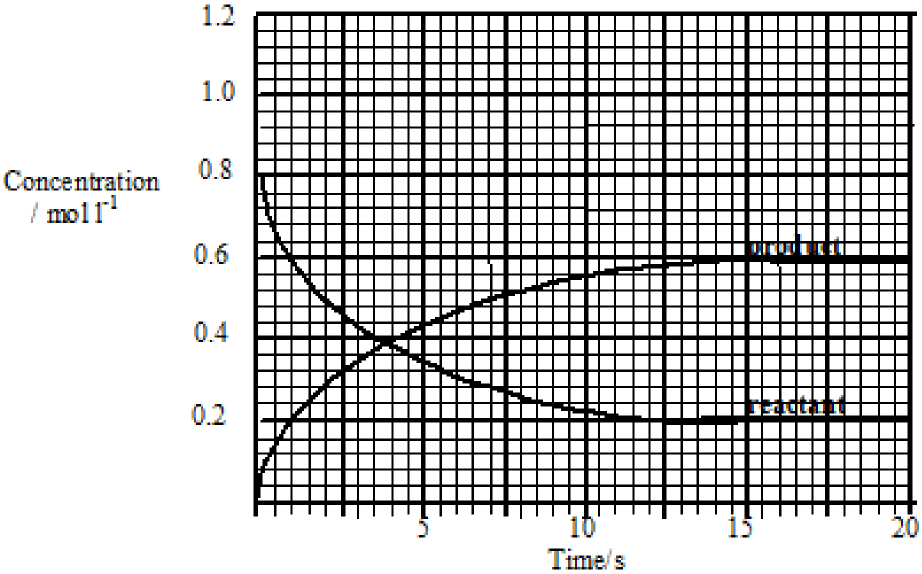
(b) Copy the graph and add curves to the graph to show the results of experiments B and C. Label each curve clearly.

- **B: Higher curve that starts in the same place as the original but ends up at 60 cm^3**
- **C: Lower curve that starts and ends in the same place as the original**

(c) Draw a labelled diagram of assembled lab apparatus which could be used to carry out these experiments.

- **Gas collected in an upside down measuring cylinder or gas syringe**
- **Gas pathway shown open**

12. The graph shows the concentrations of reactant and product as equilibrium is established in a reaction.



(a) Calculate the average rate of reaction over the first 10 s.

- **rate =** $\frac{\textbf{change in quantity}}{\textbf{change in time}} = \frac{\textbf{0.48-0}}{\textbf{10-0}}$ **= 0.048 cm^{3}s^{-1}**

(b) The equilibrium constant for a reaction is given the symbol ***K***
In this reaction ***K*** is given by:

$$\text{rate} = \frac{\text{equilibrium concentration of product}}{\text{equilibrium concentration of reactant}}$$

Calculate the value of ***K*** for the reaction.

- **rate =** $\frac{\textbf{equilibrium concentration of product}}{\textbf{equilibrium concentration of reactant}} = \frac{\textbf{0.6}}{\textbf{0.2}}$ **= 3**

(c) The reaction is repeated using a homogeneous catalyst.

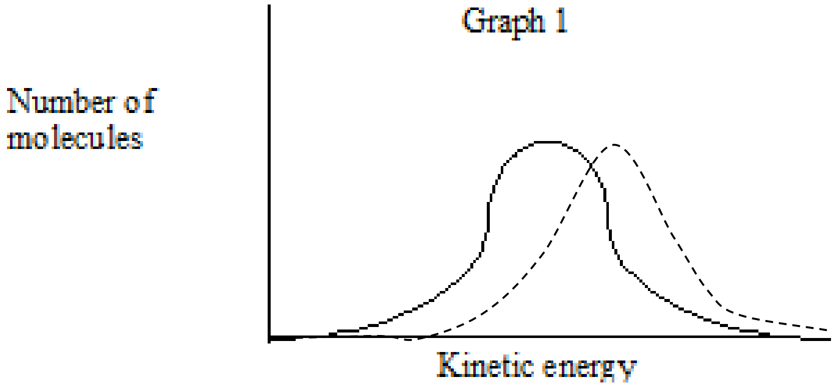
(i) What is meant by a homogeneous catalyst?

- **A catalyst that is in the same state as the reactants**

(ii) What effect would the introduction of the catalyst have on the value of ***K***?

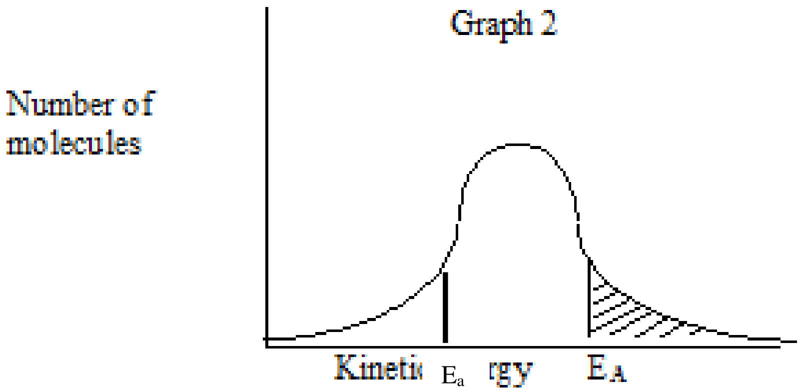
- **None**

13. (a) Graph 1 shows the distribution of kinetic energies of molecules in a gas at 30^{0}C.



Copy the graph and add a dotted line to show the distribution of kinetic energies at 20^{0}C.

(b) In Graph 2, the shaded area represents the number of molecules with the required energy of activation, E_A, for reaction to occur.



Copy the graph and draw a line to show how a catalyst affects the energy of activation.

(c) A collision involving molecules with the required energy of activation may **not** result in reaction. State a reason for this.

The molecules may not collide at the correct geometry to form a transition state.

14. What is meant by the term "activation energy"?

The minimum energy required to form a transition state. The minimum energy required for the reaction to happen.

15. The decomposition of an aqueous solution of hydrogen peroxide into oxygen and water can be catalysed by iodide ions, $I^-(aq)$, or by solid manganese (IV) oxide, $MnO_2(s)$.

For each of these catalysts state, with a reason, whether the catalysis is homogeneous or heterogeneous.

Iodide ions: homogeneous – same state as the water

Manganese oxide: hetereogeneous – different state to the water

16. An advice leaflet given to motorists when catalytic converters were first used states: "Cars fitted with catalytic converters must be run on unleaded petrol only."

(a) Outline the reasons for fitting catalytic converters, naming the substances reacting and what happens to them.

Catalytic converters remove potentially harmful and environmentally destructive gasses from the exhaust gases. Carbon monoxide is converted to carbon dioxide, nitrogen dioxide is converted to nitrogen and oxygen.

(b) Describe in terms of adsorption how catalysts work, and state the effect this has on the activation energy for the reaction.

In order to react, the gas molecule has to adsorb (stick) to the surface of the catalyst. The adsorbed catalyst/gas molecule complex is the transition state and the activation energy is the energy required for this to occur.

(c) Describe how a substance poisons a catalyst.

A catalyst can be poisoned if the reacting molecule sticks and cannot be removed.

(d) Explain the reason for the advice given at the start of the question.

Lead is a catalytic poison.

The Periodic Table: Bonding and Structure

17. a) Copy and complete the following statements.

i) Electronegativity is a measure of the **attraction** an atom in a covalent bond has for the **shared** electrons of the bond.

ii) In the Periodic table electronegativity **increases** across a period and **decreases** down a group.

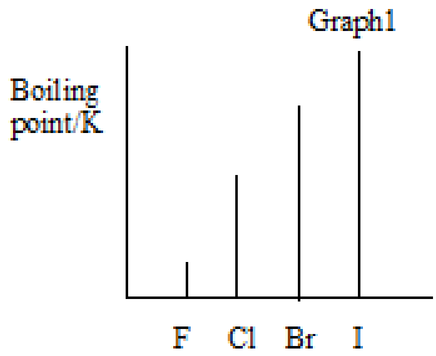
b) In each of the following pairs identify the element with the greater electronegativity.

i) phosphorus or **carbon** ii) silicon or **nitrogen**

18. All types of bonding involve electrostatic attraction between positively charged particles and negatively charged particles.
Copy and complete the table showing the three types of **strong** bonding force.

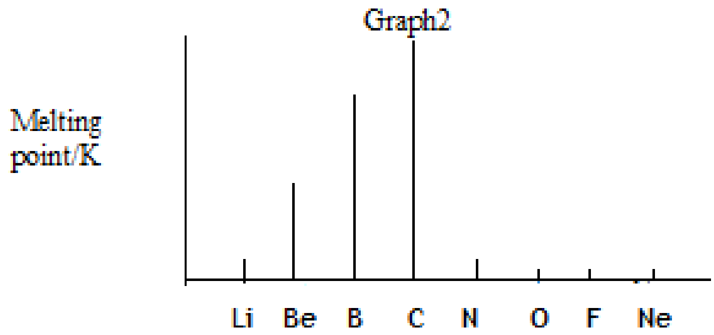
Type of bonding	Positively charged particles	Negatively charged particles
covalent	**nucleus**	shared electrons
ionic	**Positive ions**	**Negative ions**
Metallic	positive nucleus	**Delocalized electrons**

19. a) Graph 1 shows the boiling points of the Group 7 elements. Why do the boiling points increase down Group 7?



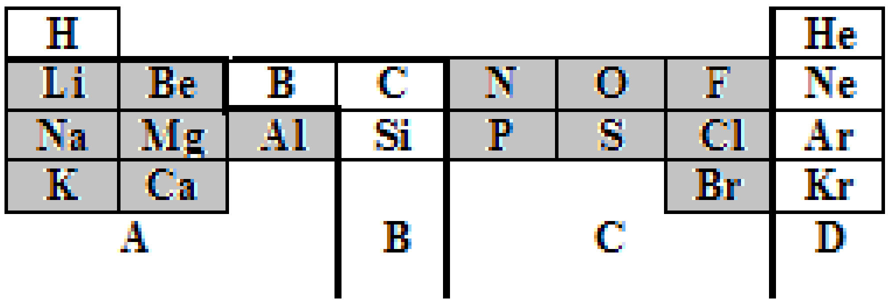
The London forces increase due to the increasing atomic size.

b) Graph 2 shows the melting points of elements from lithium to neon across the second period. Give reasons for the high melting points of boron and carbon.



Boron and carbon have a covalent network structure

20. The Periodic Table below has been divided into four sections - **A, B, C** and **D.**



a) State the type of structure in each of the four sections A, B, C and D.

A = Metallic
B = Covalent network
C= Covalent molecule
D = Mono atomic

b) In which section(s) will van der Waals' forces between the particles be significant.

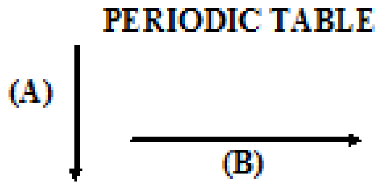
D

c) Using elements in the above table as examples, explain briefly the difference between a covalent molecular substance and a covalent network substance.

A covalent molecule is held to other molecules in a liquid or solid by weak van-der-waals forces (eg, polar-polar attractions in CO_2). In a covalent network, all the atoms are held together by strong covalent bonds.

Trends in the Periodic Table

21. a) Which arrow (A) or (B) indicates correctly a decrease in atomic size?

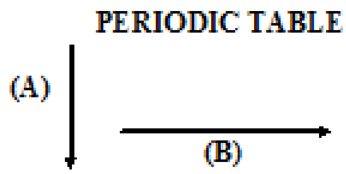


B

b) Explain why atomic size decreases in this way.

The nuclear charge is increasing along the period whereas the number of shells is not. Therefore the outer electrons "feel" a greater pull from the nucleus along the period and the atomic size decreases.

22. a) Define the **first ionization energy** of an element**.**



The energy required to remove an electron from every atom in a mole of atoms in the gas state.

b) Which arrow (A) or (B) indicates correctly a decrease in the first ionization energy of elements?

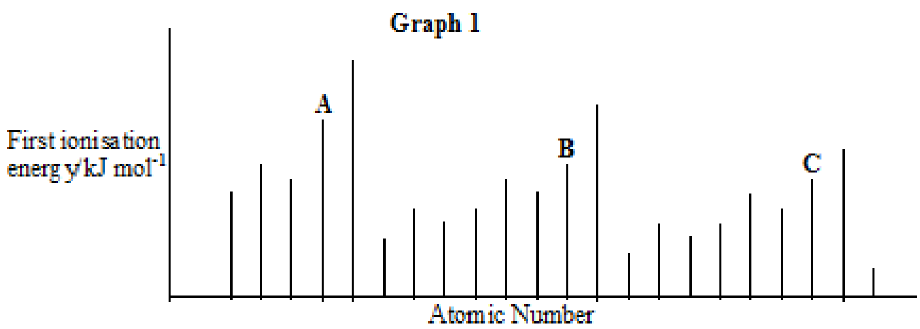
A

c) Give two reasons why the ionisation energy decreases in this way.

The number of electron shells in increasing so that the outer electron (i.e. the electron being removed) is further from the nucleus.

The inner shells of electrons "shield" the outer electron from the pull of the nucelus

23. The graph below shows the first ionisation energies of successive elements with increasing atomic number.



Elements A, B and C belong to the same group of the Periodic Table. Identify the group

Halogens (group 7)

24. Explain why the third ionisation energy of magnesium (7750 kJ mol^{-1}) is so much greater than the third ionisation of aluminium (2760 kJ mol^{-1}).

If you are removing a third electron from an ion that was neutral to start with, you are now removing a negatively charged electron from an ion with a 2+ charge. This is going to require more energy that removing an electron from a neutral or singly positively charged particle.

25. Ionisation energies can be found by applying an increasing voltage across test samples of gases until the gases ionise.
The results below were obtained from experiments using hydrogen atoms and then helium atoms.

Element	Voltage at which an atom of gas ionises/V	
hydrogen	13.6	no further change
helium	24.6	54.5

a) Why are there two results for helium but only one for hydrogen?

Helium has two electrons and hydrogen only has one.

b) (i) Write an equation which would represent the first ionisation energy of helium gas.

$He(g) \rightarrow He^{+}(g) + e^{-}$

(ii) Why is the first ionisation of helium higher than that of hydrogen?

The electron being removed from Helium is being removed from a full shell. This is going to require a lot of energy.

c) The ionisation energy, I.E. , can be found from:

I.E. = voltage x 1.6×10^{-19} J

Calculate a value for the first ionisation energy of helium.

I.E. = voltage x 1.6×10^{-19} J = $24.6 \times 1.6 \times 10^{-19}$ J = 3.936×10^{-18} J

Bonding in Compounds

26. There are many types of attractive force, some are weak and some are strong.

A positively charged ions and negatively charged ions	**B** temporary dipole and induced dipole	**C** positively charged nucleii and delocalised electrons
D permanent dipole and permanent dipole	**E** positively charged nucleii and shared electrons	

Identify the statement(s) referring to

a) van der Waals' forces
B

b) the forces between oxygen and hydrogen atoms in water
D

c) the intemolecular forces in HCl gas.
D, B

d) ionic bonds
A

27. The covalent bond in hydrogen chloride gas is polar and the molecule is polar. The covalent bonds in silicon tetrachloride are also polar. Explain why the silicon tetrachloride molecule is non-polar.
The silicon tetrachloride molecule is symmetric (tetrahedral) so all the polar bonds are pulling against each other.

28. There are many types of bonding force between atoms and molecules.

A permanent dipole to permanent dipole interactions	**B** non-polar covalent bonds	**C** hydrogen bonds
D ionic bonds	**E** metallic bonds	**F** London dispersion forces

a) Identify the three forces present in hydrogen fluoride.
A, C, F

b) Identify the force(s) present in

i) methane	**F**
ii) sodium chloride	**D**
iii) hydrogen bromide	**A, C**
iv) neon	**F**
v) oxygen	**B, F**

c) Identify the bond(s) and/or force(s) of attraction

i) responsible for the low boiling point of argon.	**F**
ii) that can exist **between** molecules.	**A, C, F**
iii) that allow electrons a lot of free movement.	**E**

29. Which of the compounds below have:

A Cl_2	B HF	C NaCl
D NH_3	E CH_3COCH_3	F SiO_2

Identify the substance(s) where the intermolecular forces are

a) van der Waals' forces **only.** **A**

b) hydrogen bonds **B, D, E**

30. Elements and compounds show a variety of structures.

A Cl_2	**B** Na	**C** NaCl
D SiO_2	**E** NH_3	**F** C(diamond)

Identify the substance(s)

a) with a tetrahedral arrangement of bonds in a covalent network.
D

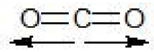
b) which can conduct electricity because of delocalised electrons.
B

c) with discrete covalent molecules
A, E

31. Many of the properties of water arise from the presence of polar O - H bonds which make the water molecules polar.
Carbon dioxide contains polar C = O bonds but its molecules are non-polar.

a) Explain this difference with the aid of diagrams of each molecule, showing polarities.

In the carbon dioxide molecule, the polar oxygen atoms "pull" electrons in opposite directions to the central carbon atom, the bonds are therefore polar (O is δ- and C is δ+) but because the molecule is symmetric (linear), the molecule overall ins non-polar.



b) Water is unusual in that in the solid form (ice) is less dense than the liquid form.
Explain why water behaves in this way.

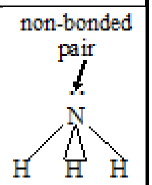
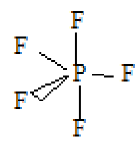
Water molecules are "bent" with O δ- and H δ+. The δ+ and δ+ connect when water freezes but the hydrogen bond between the O and H is very long. This leads to the very "open" structure of frozen water.

32. Both bonded and non-bonded pairs of electrons repel each other and this determines the shape of the molecule.
The following procedure is used to find the total number of pairs of electrons around a central atom.

(i) Note the number of electrons in the outer energy level (shell) of the central atom.
(ii) Note the number of other atoms present --- each atom provides one electron for bonding.
(iii) Add (i) and (ii) to give the total number of electrons.
(iv) Divide this number by two to give the number of electron pairs - both bonded and non-bonded pairs.

Example:- with ammonia, NH_3, N is the central atom.
(i) 2,5 = 5 electrons
(ii) 3H 3 x 1 = 3 electrons
(iii) Total = 8 electrons
(iv) 8 electrons gives four pairs.

Since NH_3 only has 3 bonds there is one non-bonded pair. The 4 pairs of electrons repel each other, giving the pyramid shape of the ammonia molecule as shown in the first row of the table.
Copy and complete the table.

Formula	Outer electrons in central atom	Total number of electrons	Bonded pairs	Non-bonded pairs	Molecular shape
NH_3	5	8	3	1	
CCl_4	4	**8**	4	0	**Tetrahedral**
$BeCl_2$	2	4	2	**0**	**Linear**
PF_5	**5**	10	5	0	

33.a) Predict the type of bonding (non-polar covalent, polar covalent or ionic) between the following elements.

i) magnesium and sulphur
ionic
ii) oxygen and phosphorus
polar covalent
iii) nitrogen and nitrogen
non-polar covalent
iv) fluorine and oxygen
polar covalent

b) Justify, by reference to electronegativity values, your answers to i) and iv).
There is an extremely large electronegativity difference between magnesium and sulphur so the electrons in a covalent bond would be attracted almost completely towards the sulphur. This effectively leads to an ionic attraction. For fluorine and oxygen the electronegativity difference is much lower so the bond would be polar covalent as opposed to ionic.

34. The boiling points of compounds depend on the intermolecular forces.

Name	Formula	Molecular mass	Boiling point (^0C)
butane	$CH_3CH_2CH_2CH_3$	**58amu**	- 0.5
propanone	CH_3COCH_3	**58amu**	56

a) Copy the table and calculate the molecular mass for each compound.

b) Explain why the boiling points are different.

Hydrogen bonding is present between the molecules in propanone. This leads to a greater attraction between the molecules and a higher boiling point. Only London forces are present in the bonding between butane molecules.

35. The Group 5 hydrides are covalent compounds.

Compound	Enthalpy of formation/kJ mol^{-1}	Boiling point/K
NH_3	-46	240
PH_3	+6	185
AsH_3	+172	218

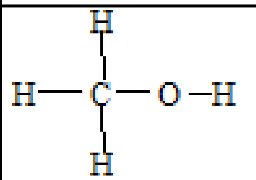
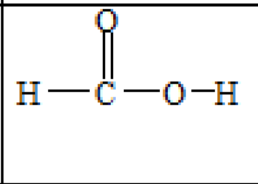
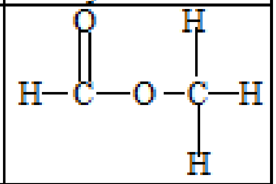
a) What is the trend in stability of the group 5 hydrides?

The enthalpy of formation increases as you go down the group 5 hydrides so the stability is decreases.

b) Explain why the boiling point of NH_3 is higher than the boiling point of PH_3 and AsH_3.

Hydrogen bonding is present between the molecules in NH_3. This leads to a greater attraction between the molecules and a higher boiling point.

36.The table below shows the boiling point, molecular mass and structure of the simplest alkanol, methanol, the simplest alkanoic acid, methanoic acid and the ester methyl methanoate which forms when the acid and the alkanol react together in a condensation reaction.

Compound	methanol	methanoic acid	methyl methanoate
Structure			
Molecular mass	32	46	60
Boiling point(^{0}C)	65	100	

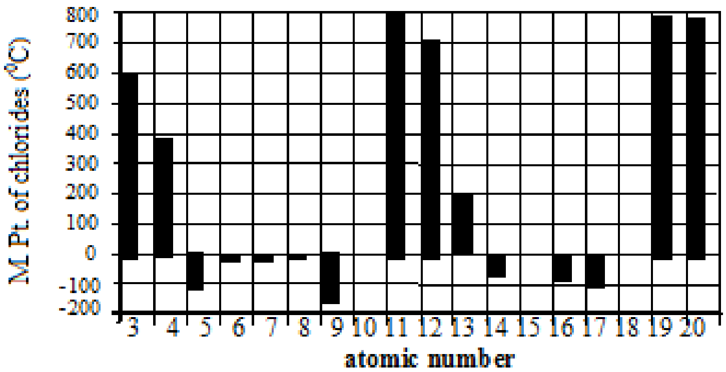
a) Using **molecular mass** as the **only criterion**, use the boiling points of methanol and methanoic acid to predict the boiling point of methyl methanoate and put it in the table.

~80 ^{0}C

b) The boiling point of the ester is in fact 32^{0}C. Explain in terms of the intermolecular forces why this value is so different from your prediction.

There is no hydrogen bonding between the molecules of methyl methanoate.

37. The bar chart shows the melting points of chlorides of elements 3 to 20 (with no bars for 10, 15 and 18).



a) Copy and complete this statement describing the pattern for these **melting points** related to the **Group number.**
In general as the Group number increases the melting point of the chloride **decreases**.

b) Explain why no values are given for elements 10 and 18.
They are noble gases and don't tend to form compounds

c) **From the bar chart**, state which of the chlorides has the weakest forces between the molecules.

Fluorine chloride

d) Predict a value for the melting point of the chloride of element 15.

-80 ^{0}C

38. Explain in terms of its structure and bonding why silicon carbide can be used to make stones for sharpening chisels and knives.

It is a covalent network structure so is therefore very strong and hard.

39. Consider the substances: potassium, bromine and potassium bromide.

a) Construct a table to show the type of bonding, the structure, the solubility or reaction with water, the state at room temperature and the electrical conductivity of the three substances.

Potassium – solid – reacts with water – conducts electricity
Bromine – liquid – doesn't react with water – doesn't conduct electricity
potassium bromide - doesn't react with water – conducts electricity as a liquid or solution

b) Explain the solubility of potassium bromide in water in terms of its bonding.

Potassium bromide is ionic and ionic substances tend to dissolve in water.

40. Lithium iodide is quite soluble in non-polar solvents e.g. white spirit (a mixture of hydrocarbons).

a) What does this statement suggest about the type of bonding in lithium iodide?

It is somewhat non-polar

b) State, with an explanation, whether you would expect lithium fluoride to be more or less soluble than lithium iodide in non-polar solvents.

Less. The electronegativity difference in lithium fluoride is much larger than the electronegativity difference in lithium iodide so lithium fluoride is more likely ot be ionic and therefore dissolve in a polar solvent.