Important Notes : Miscellaneous

General unreactivity of alkanes (including towards polar reagents) is due to : only contains STRONG C–H σ-bonds
: (alkanes are) non-polar molecules

NO_x reacts with unburnt hydrocarbons to form PEROXYACETYL NITRATE // PAN, which is a component of photochemical smog

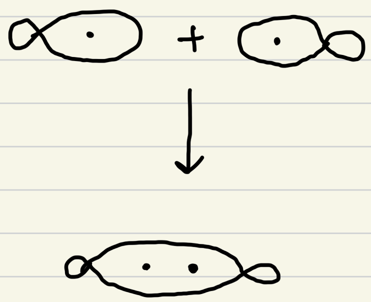
Lack of reactivity of N_2 : STRONG $N \equiv N$ triple bond ... requiring large amount of energy to overcome
: (N_2 is) non-polar molecule

NH_4NO_3 is added to soil to act as fertiliser // add nutrients for plants

$Ca(OH)_2$ is added to soil, to reduce the acidity of the soil - mg, to increase/raise the pH - mg

Describe how sp hybridised orbitals are formed:
↳ Mixing / overlap / combination of one s and one p orbital - mg

Sketch a diagram to show how two sp hybrid orbitals can form a σ-bond (SIGMA)



M1

M2 — mg

allotrope of phosphorus	formula	melting point / °C
white	P_4	44
red	P	590

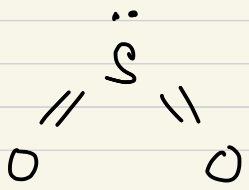
State the meaning of volatile: easily vaporised / easily evaporates / turns to gas easily

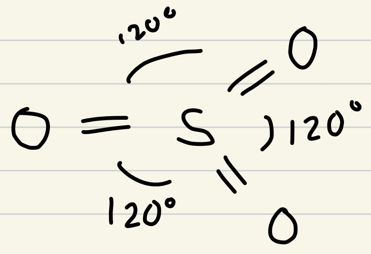
outer shell of electrons is full/has a complete octet or
valence shell of electrons is full/has a complete octet or
activation energy is too high or ionisation energy is too high
} unreactivity of Noble Gases (Ar)

Use of esters: solvents or perfumes or flavourings

Why Branched molecule has lower bp: straight chain molecules have larger surface area and thus can interact more with each other, straight chain molecules can pack more closely therefore stronger van der Waals' forces

1 lp
2 bp





0 lp
3 bp

Shape: non-linear or bent — trigonal planar

Bond Angle: 115° – 120° — 120°

Why ΔHi decreases down the Group

increasing distance of (outer) electron(s) from nucleus → write this alongside increase in atomic radius, whenever change in atomic radius is involved.

OR increasing distance of outer / valence shell from nucleus

increased shielding / screening (from inner shells)

reduces attraction

Oxidising agent: (Species) that gains e^- // e^- acceptor —ms

Why nitrogen and oxygen react during lightning strikes

(lightning) provides the (high) activation energy

When an ammonium salt is heated with a base, NH_3 gas is released because NH_3 is displaced // displaces NH_3 —ms

Explain what is meant by high vapour pressure.

M1 gas / vapour (particles / molecules) in equilibrium (with liquid / solid)

M2 greater proportion of gas (particles) than liquid (particles) (in comparison to a liquid of lower vapour pressure)

★★

Attempts to make Cl F7 have failed but IF7 has been prepared.

Suggest an explanation for the existence of IF7 and for the non-existence of Cl F7.

Ans) I atom is larger than Cl atom (1)

cannot pack 7 F atoms around Cl atom or can pack 7 F atoms around I atom (1)

Why metallic bond?

M1 comment explaining high melting point of Cu many strong metallic bonds OR many strong (electrostatic) attractions between cations and delocalised electrons OR strong bonds in giant metallic structure.

M2 comment explaining electrical conductivity of Cu delocalised electrons are free are to move through the structure (owtte)

Phosphorus (V) chloride reacts with water to give an acidic solution of pH 1-2 -ms

$NaCl$ "dissolves" in H_2O. PCl_5 "hydrolysed" by H_2O

Why is a molecule polar : Bond polarity description
: Molecule is not symmetrical
: Dipoles do not cancel out

Vapour Pressure : gas (particles / molecules) in equilibrium with liquid
gas (particles) exert a pressure (on the walls of a container / on the surface of the liquid)

Dynamic Equilibrium : rates of forward and reverse / backward reactions are equal
closed / sealed system / container
no change in measurable properties / no change in macroscopic properties / concentration of reactants and products remain constant

Physical properties typical of ceramic materials :
- strong / retain strength (over certain temperatures / conditions)
- non-conductors of electricity / electrical insulators
- high melting points

Examples: MgO, Al2O3,SiO2

Disproportionation : when a species is both oxidised and reduced

Enthalpy Change of formation : (enthalpy / energy change) when one mole of a compound/substance is formed from its elements in their standard states

Catalyst : increase the rate
lowers the activation energy
are regenerated at the end of the reaction

Nucleophile : a lone pair / e^- pair donor

Substitution Reaction : a reaction where an atom or group of atoms replaces another

Primary Alcohol: Compound which contains the $-CH_2OH$ group

E_a: The minimum amount of energy required for a collision to be effective

Chiral Centre: An atom that is bonded to 4 different atoms or groups of atoms

Even if e^- from the same shell are removed, sheilding effect decreases

Why does a molecule have a certain shape.
↳ It has x bond pairs and y lone pairs
↳ Relative strength of repulsion between e^- pairs, lone pair - lone pair > lone pair - bond pair > bond pair - bond pair

OR
↳ It only has x bond pairs and 0 lone pairs
↳ bonding pairs repel equally

Composition of products of incomplete combustion: CO, C, unburnt hydrocarbons, less CO_2 than complete combustion

Conditions for the formation of NO_x in an internal combustion engine
↳ high temperature
↳ high pressure

Role of NO_x in acid rain: Acts as catalyst. Catalyses formation of SO_3 from SO_2

$$2NO + O_2 \rightarrow NO_2$$
$$NO_2 + SO_2 \rightarrow SO_3 + NO$$
$$SO_3 + H_2O \rightarrow H_2SO_4$$

Sources of NO_x:
↳ Man-made: Internal Combustion Engines
↳ Natural: Lightning ϟ

Define Relative · Atomic Mass · Formula · Molecular: Weighted average mass of the · isotopes of an element · formula unit / molecule of a compound compared to the mass of the relative atomic mass unit, 1 / 12th the mass of a Carbon-12 isotope

Relative Isotopic Mass: mass of an isotope compared to 1/12th the mass of a Carbon-12 isotope

Why is the solution acidic after the hydrolysis of ACl_b
↳ Reaction produces hydrogen chloride gas that dissolves in water to produce hydrochloric acid which is a strong acid that dissociates completely in aq sol^n to give H^+ ions.

Why is $HI_{(g)}$ not prepared by reacting NaI and H_2SO_4
↳ $HI_{(g)}$ is not the only gaseous product, SO_2, H_2S and $I_{2(g)}$ also produced
↳ H_2SO_4 oxidises HI ora

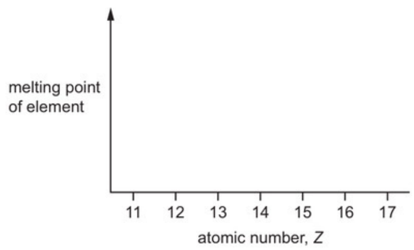
How does Al_2O_3 coating prevent the Al inside from reacting with water
↳ Al_2O_3 does not ~~react~~ with water
↳ Al_2O_3 insoluble in water

All Hydrogen Halides HX are colourless gases

Melting point of S_8 > Na

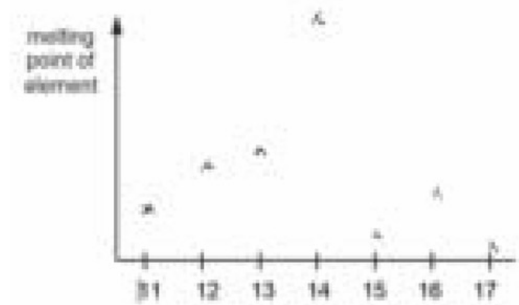
8

(b) Use the axes to sketch a graph that shows the trend in melting points of the elements with atomic numbers 11 to 17.



[2]

2(b)



How to write CarboCations : $CH_3C^+H_2$ OR $(CH_3CH_2)^+$

$3NO_2 + H_2O \longrightarrow 2HNO_3 + NO$

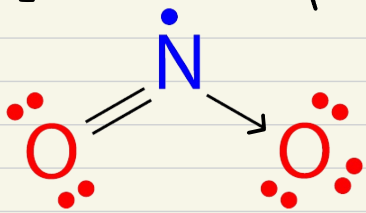
MCQs

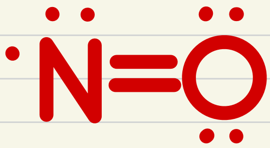
Under acidic conditions, amine functional group $(-NH_2)$ turns into $-NH_3^+$

Reaction of Halogenoalkanes with $NaOH_{(aq)}$ or H_2O or (aq) is hydrolysis reaction

$MgCl_{(s)} + H_2O_{(l)}$: gives a sol^n with pH 6.5
: NO steamy fumes
: does NOT react with Na_2CO_3
: Dissolves in H_2O NOT hydrolyses

NO_2 molecule is polar

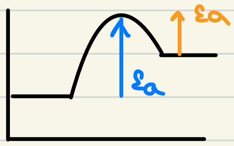
 : Bonding in NO_2 molecule

 : Bonding in NO molecule

pH of $Mg(OH)_2$ // MgO in water is 9

Acid-base reaction involving salt of carboxylic acid gives carboxylic acid, e.g. $CH_3COO^-Na^+$ gives CH_3COOH

In a reversible reac^n, the exothermic reac^n has a lower E_a than the endothermic reac^n



$\overset{+2}{3NO_2} + H_2O \longrightarrow \overset{+5}{2HNO_3} + \overset{+2}{NO}$

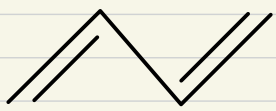
O : 1 R : 2

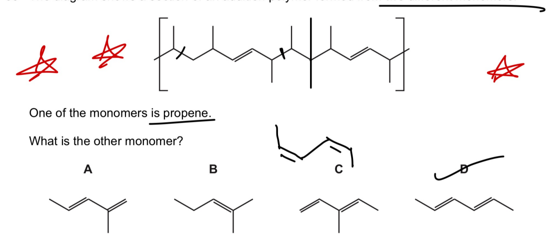
Cl_2 and HCl are acidic gas but H_2 is a neutral gas

Thermal Cracking: Any compound CAN be obtained from cracking as long as the no. of C atoms in the cracked compound < no. of C atoms in the original compound.

e.g. $C_{17}H_{36}$ CAN give $C_{16}H_{34}$, it is just less likely.

s.t.p. different from r.t.p. (conditions included at the back of qp)

monomers  can give repeat unit 


39 The diagram shows a section of an addition polymer formed from two different monomers.

One of the monomers is propene.

What is the other monomer?

A B C D

Halogen	Color in water	Color in hexane
Cl_2	**pale, yellow green**	**yellow green (not orange)**
Br_2	reddish	reddish or yellowish orange
I_2	colorless to violet	violet (pinkish)

Problem solving notes

- When finding the number of carbon atoms in a molecule from m peak and m+1 peak using the equation. Whatever result you get round it off. If you get 5.6 then round up to 6 carbon atoms and if you get 5.4 round down to 5 carbon atoms
- In a precipitation reaction you will always get the hydroxide of the salt, never the oxide, wherever applicable

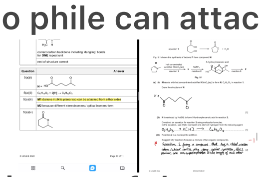
(b) $LiAlH_4$ cannot be used in aqueous solution because it reacts with water to produce LiOH(aq), H_2(g) and a white precipitate which is soluble in excess sodium hydroxide.

Identify the white precipitate.

~~Al_2O_3~~ $Al(OH)_3$ .... [1] −1

-
- Bond energy value is positive
- Propanol reactions with HCN better under alkaline conditions than acidic conditions (*ionised H+ and CN- revert back to HCN molecule due to the presence of strong acid, excess H+ ions, in acidic conditions)*
- If question says AVERAGE bond energy the mcq option where only one bond is broken is less preferable if there is another option where multiple bonds are broken and that energy is divided

When a chiral molecule is formed from a planar molecule, upon addiction reaction, both enantiomers will form in almost equal amounts because the electro/neucleo phile can attack from 2 directions, both sides of the plane, with equal ease.



When a chiral molecule is formed after a Sn2 substitution reaction with only one of the enantiomers of a chiral molecule, only one type of enantiomer will be formed after the subsequent substitution reaction (the formation of the other enantiomer will be negligible) because it is a tetrahedral molecule and there is limited space from where the nucleophile

can attack. Nucleophile can mainly attack from one direction(where the bond is in the process of being broken) thus only one enantiomer forms.

(c) 1 cm^3 of $MgCl_2(aq)$ is placed in a test-tube. A few drops of $AgNO_3(aq)$ are added, followed by 1 cm^3 of dilute $NH_3(aq)$.

State in full what is observed in this experiment.

upon adding AgNO3(aq), white ppt is formed. upon adding NH3(aq), the existing white ppt dissolves and give colourless soln but new white ppt is now also formed that does not dissolve. [2]

-1

white ppt / solid/
solid (re)dissolves (on addition of NH,)

All the ammonia is bonded with Ag. No NH3 left to give ppt with Mg (Mg(OH)2). The reaction with AgX with NH3 will always take priority over ppt reaction with Mg. If the question specified "excess" NH3 then it would have further reacted with Mg to give white ppt.

Nitrile can only be reduced using hydrogen in the presence of nickel catalyst NOT platinum catalyst

M1: reaction less vigorous (down the group)
M2: Any two of the following for one mark:
- electronegativity decreases
- less attractive to e – addition
- weaker oxidising agent
- greater nuclear charge outweighing increased shielding (ENC argument)

The hydrogen halides HCl, HBr and HI are all colourless gases at room temperature.

(a) The hydrogen halides can be formed by reacting the halogens with hydrogen.

Describe and explain the relative reactivity of the halogens down the group when they react with hydrogen to form HCl, HBr and HI.

This question asked you to explain the relative reactivity of halogens down the group whereas the one in the mock reactivities of halogens with hydrogen in terms of bond energy

When is it difficult to distinguish between two molecules using IR-SPECTROSCOPY?

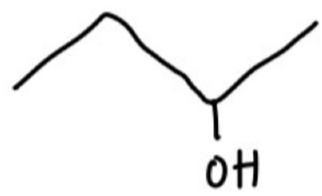
- **The molecules have the same bonds (despite of different functional groups) who have similar or overlapping absorption ranges thus their peaks/absorptions will be similar**
- **Both molecules have bonds in similar environments :** ***what this means is that both***

molecules have very similar structural formulae (eg. same carbon skeleton with with functional groups in the same positions expect 1 different functional group in the same position). This makes it difficult to judge the different molecules because this results in very similar finger print region(finger print region is below 1500cm-1 and is unique for all molecules of the same structural formula.

H

name ....butan-1-ol....................

J



name ....butan-2-ol....................

H has a higher boiling point that J because in H the OH bond is in the open thus hydrogen bonds can form more more directions (more interaction) however in J the OH bond is more restricted, hydrogen bonds cannot form from all directions (due to the presence of the addiction all alkyl group) and thus less interaction .

- [ ] Iodoform reaction is a hydrolysis reaction
  Reminders – 18/4/24, 9:00 AM

Morning

- [ ] Must answer reactivity of metals in terms of ionisation energy
  Reminders – 9:00 AM

- [ ] AlCl3 is always simple molecular structure
  Reminders – 9:00 AM

- [ ] Similar environment(similar structural formulae) of molecules will make it difficult to distinguish between them using ir spectroscopy because they will have similar finger print region
  Reminders – 9:00 AM

AS Chemistry-2022-2024-syllabus

3 Subject content - AS Level subject content

Isotopes of a particular element have the same chemical properties because they have the same number of electrons//same electronic structure. They have slightly different physical properties, such as small differences in density or small differences in mass, because they have different numbers of neutrons.

14

sub-shells 4 describe the electronic configurations to include the number of electrons in each shell, sub-shell and orbital 5 explain the electronic configurations in terms of energy of the electrons and inter-electron repulsion 6 determine the electronic …

15

describe and sketch the shapes of s and p orbitals describe a free radical as a species with one or more unpaired electrons

15

…ies are due to the attraction between the nucleus and the outer electron 6 explain the factors influencing the ionisation energies of elements in terms of nuclear charge, atomic/ionic radius, shielding by inner shells and sub-shells and spin-pair repulsion …

15

electrons are arranged outside the nucleus in energy levels or quantum shells. These principal energy levels or principal quantum shells (symbol n) are numbered according to how far they are from the nucleus. Ground State (of an atom): the lowest possible energy level of the atom in question.

15

1 define the unified atomic mass unit as one twelfth of the mass of a carbon-12 atom 2 define relative atomic mass, A_r, relative isotopic mass, relative molecular mass, M_r, and relative formula mass in terms of the unified atomic mass unit

16

The empirical formula of a compound is the simplest whole number ratio of the number of atoms of each element present in one molecule or formula unit of the compound. The molecular formula of a compound shows the number of atoms of each element present in a molecule. The formula for an ionic compound is always its empirical formula.

16

(When performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question. When rounding up or down, candidates should ensure that significant figures are neither lost unnecessarily nor used beyond what is justified (see also Mathematical requirements section).)

16

Some compounds can form crystals which have water as part of their structure. This water is called the water of crystallisation. A compound containing water of crystallisation is called a hydrated compound, e.g. hydrated copper(II) sulfate, $CuSO_4 \cdot 5H_2O$ A compound which does not contain water of crystallisation is called an anhydrous compound e.g. anhydrous copper(II) sulfate, $CuSO_4$

16

1 define electronegativity as the power of an atom to attract electrons to itself 2 explain the factors influencing the electronegativities of the elements in terms of nuclear charge, atomic

17

…shielding by inner shells and sub -shells 3 state and explain the trends in electronegativity across a period and down a group of the Periodic Table 4 use the differences in Pauling electronegativity values to predict the formation of ionic and covalent bonds

17

If the difference in electronegativity for two elements is less than the difference in electronegativity of two elements that form a covalent bond, then the first two elements will also form a covalent bond.

17

1 define ionic bonding as the electrostatic attraction between oppositely charged ions (positively charged cations and negatively charged anions) 2 describe ionic bonding including the examples of sodium chloride, magnesium oxide and calcium fluoride

17

A

1 define metallic bonding as the electrostatic attraction between positive metal ions and delocalised electrons

17

A

1 define covalent bonding as electrostatic attraction between the nuclei of two atoms and a shared pair of electrons

17

A

(b) understand that elements in period 3 can expand their octet including in the compounds sulfur dioxide, SO_2, phosphorus pentachloride, PCl_5, and sulfur hexafluoride, SF_6

18

A

• σ bonds are formed by direct overlap of orbitals between the bonding atoms • π bonds are formed by the sideways overlap of adjacent p orbitals above and below the σ bond

18

A

• bond energy as the energy required to break one mole of a particular covalent bond in the gaseous state • bond length as the internuclear distance of two covalently bonded atoms

18

A

(b) use bond energy values and the concept of bond length to compare the reactivity of covalent molecules

18

A

• BF_3 (trigonal planar, 120°) • CO_2 (linear, 180°) • CH_4 (tetrahedral, 109.5°) • NH_3 (pyramidal, 107°) • H_2O (non-linear, 104.5°) • SF_6 (octahedral, 90°) • PF_5 (trigonal bipyramidal, 120° and 90°)

18

A

predict the shapes of, and bond angles in, molecules and ions analogous to those specified in 3.5.1

18

(b) use the concept of hydrogen bonding to explain the anomalous properties of H_2O (ice and water): • its relatively high melting and boiling points • its relatively high surface tension • the density of the solid ice compared with the liquid water

19

(a) describe van der Waals' forces as the intermolecular forces between molecular entities other than those due to bond formation , and use the term van der Waals' forces as a generic term to describe all intermolecular forces

19

• instantaneous dipole – induced dipole (id-id) force, also called London dispersion forces

19

• permanent dipole – permanent dipole (pd-pd) force, including hydrogen bonding

19

– permanent dipole force between molecules where hydrogen is bonded to a highly electronegative atom state that, in general, ionic, covalent and metallic bonding are stronger than intermolecular forces

19

3.5 (dot-and-cross diagrams may include species with atoms which have an expanded octet or species with an odd number of electrons)

19

1 explain the origin of pressure in a gas in terms of collisions between gas molecules and the wall of the container 2

understand that ideal gases have zero particle volume and no intermolecular forces of attraction 3 state and use the ideal gas equation p…

19

Why is water a good solvent for ionic compounds **water molecules are polar; ions are charged; attraction between een water molecules and ions;**

19

…describe, in simple terms, the lattice structure of a crystalline solid which is: (a) giant ionic, including sodium chloride and magnesium oxide (b) simple molecular, including iodine, buckminsterfullerene C_{60} and ice (c) giant molecular, including silicon(IV) oxide, graphite and diamond (d) giant metallic, including copper

20

2 describe, interpret and predict the effect of different types of structure and bonding on the physical properties of substances, including melting point, boiling point, electrical conductivity and solubility 3 deduce the type of structure and bonding pre…

Insoluble: giant covalent Soluble: simple molecular: gets hydrolysed by water : giant ionic: dissociates in water

20

(a) standard conditions (this syllabus assumes that these are 298 K and 101 kPa) shown by $\ominus$.

20

(b) enthalpy change with particular reference to: reaction, ΔH_r , formation, ΔH_f, combustion, ΔH_c, neutralisation, ΔH_{neut} 4 understand that energy transfers occur during chemical reactions because of the breaking and making of chemical bonds 5 use bond e…

20

bond breaking) to calculate enthalpy change of reaction, ΔHr 6 understand that some bond energies are exact and some bond energies are averages 7 calculate enthalpy changes from appropriate experimental results, including the use of the relationships q = m…

20

(a) understand what is meant by a reversible reaction

A reaction in which the products can react to re-form the original reactants is called a reversible reaction

21

2 define Le Chatelier's principle as: if a change is made to a system at dynamic equilibrium, the position of equilibrium moves to minimise this change 3 use Le Chatelier's principle to deduce qualitatively (from appropriate information) the effects of cha…

21

value of the equilibrium constant for a reaction 10 describe and explain the conditions used in the Haber process and the Contact process, as examples of the

Compromising temperature, compromising more yield for faster rate

21

…describe strong acids and strong bases as fully dissociated in aqueous solution and weak acids and weak bases as partially dissociated in aqueous solution 5 appreciate that water has pH of 7, acid solutions pH of below 7 and alkaline solutions pH of above

7 6 explain qualitatively the differences in behaviour between strong and weak acids …

22

…ow 7 and alkaline solutions pH of above 7 6 explain qualitatively the differences in behaviour between strong and weak acids including the reaction with a reactive metal and difference in pH values by use of a pH meter, universal indicator or conductivity …

22

Many indicators can be considered as weak acids in which the acid (HIn) and its conjugate base (In–) have different colours. Adding an acid to this indicator solution shifts the position of equilibrium to the left. There are now more molecules of colour A. Adding an alkali shifts the position of equilibrium to the right. There are now more ions of colour B. The colour of the indicator depends on the relative concentrations of HIn and In–. The colour of the indicator during a titration depends on the concentration of H +ions present. *The pH range of the indicator must have some overlap with the pH range of the end point of the reaction. The endpoint of the indicator must fall within the pH range of the end point of the reaction.*

22

weak alkalis 10 select suitable indicators for acid-alkali titrations , given appropriate data (pK $_a$values will not be used)

22

changes on the rate of a reaction 3 use experimental data to calculate the rate of a reaction

22

1 define activation energy, E $_A$, as the minimum energy required

A for a collision to be effective 2 sketch and use the Boltzmann distribution to explain the significance of activation energy 3 explain qualitatively, in terms both of the Boltzmann distribution…

22

A (a) explain that, in the presence of a catalyst, a reaction has a different mechanism, i.e. one of lower activation energy

22

A P_4O_{10}, SO_2), chlorine (to give NaC*l*, MgCl_2, A*l* Cl_3, SiCl_4, PCl_5) and water (Na and Mg only) 2 state and explain the variation in the oxidation number of the oxides (Na$_2$O, MgO, Al_2O$_3$, P_4O_{10}, SO_2 and

23

SO_3 only) and chlorides (NaC*l*, MgCl_2, A*l* Cl_3, SiCl_4, PC*l* $_5$only) in terms of their outer shell (valence shell) electrons 3 describe, and write equations for, the reactions, if any, of the oxides Na$_2$O, MgO, Al_2O$_3$, SiO$_2$, P_4O_{10}, SO_2

23

A SO_2 and SO_3 and the hydroxides NaOH, Mg(OH)$_2$, A *l* (OH)$_3$ including, where relevant, amphoteric behaviour

23

A with water including the likely pHs of the solutions obtained 6 explain the variations and trends in…

23

9.2.5 in terms of bonding and electronegativity 7 suggest the types of chemical bonding present in the chlorides and oxides from observations of their

Probably a good idea to just mention the electronegativity in the answer as it's often a marking point

23

…escribe the colours and the trend in volatility of chlorine, bromine and iodine 2 describe and explain the trend in the bond strength of the halogen molecules 3 interpret the volatility of the elements in terms of instantaneous dipole–induced dipole forces

24

1 describe the relative reactivity of the elements as oxidising agents 2 describe the reactions of the elements with hydrogen and explain their relative reactivity in these reactions 3 describe the relative thermal stabilities of the hydrogen halides and e…

M1: reaction less vigorous (down the group) M2: Any two of the following for one mark: * electronegativity decreases * less attractive to e – addition * weaker oxidising agent * greater nuclear charge outweighing increased shielding (ENC argument)

24

hot aqueous sodium hydroxide and recognise these as disproportionation reactions 2 explain, including by use of an equation, the use of chlorine in water purification to include the production of the active species HOC *l* and C *l* O _ which kill bacteria.

25

1 explain the lack of reactivity of nitrogen, with reference to triple bond strength and lack of polarity 2 describe and explain:

25

…state and explain the natural and man-made occurrences of oxides of nitrogen and their catalytic removal from the exhaust gases of internal combustion engines 4 understand that atmospheric oxides of nitrogen (NO and NO_2) can react with unburned hydrocarbons to form peroxyacetyl nitrate, PAN, which is a component of photochemical smog 5 describe the role of NO and NO_2 in the formation of acid rain both directly and in their catalytic role …

25

form peroxyacetyl nitrate, PAN, which is a component of photochemical smog 5 describe the role of NO and NO_2 in the formation of acid rain both directly and in their catalytic role in the oxidation of atmospheric sulfur dioxide

25

…hydrocarbon as a compound made up of C and H atoms only 2 understand that alkanes are simple hydrocarbons with no functional group 3 understand that the compounds in the table on page 26 and 27 contain a functional group which dictates their physical and chemical properties…

28

…organic molecules as either straight-chained, branched or cyclic 2 describe and explain the shape of, and bond angles in, molecules containing sp, sp^2 and sp^3 hybridised atoms 3 describe the arrangement of σ and π bonds in molecules containing sp, sp^2 and sp^3 hybridised atoms 4 understand and use the term planar when describing the arrangement of atoms in organic molecules, for example ethene

28

(a) addition of hydrogen to an alkene in a hydrogenation reaction, H_2 (g) and Pt/Ni catalyst and heat (b) cracking of a longer chain alkane, heat with Al_2O_3 2 describe:

29

…cracking can be used to obtain more useful alkanes and alkenes of lower M_r from heavier crude oil fractions 5 understand the general unreactivity of alkanes, including towards polar reagents in terms of the strength of the C–H bonds and their relative lack of polarity 6 recognise the environmental consequences of carbon monoxide, oxides of nitrogen and unburnt hydrocarbons arising from the combustion of alkanes in the internal combustion engine and of their catalytic removal

Why is cracking important? Cracking is a very important process for three reasons: 1. Cracking is used to produce petrol. Our need for petrol is greater than our need for diesel oil or lubricating oil. Diesel and lubricating oils are made up of long-chain alkanes. Through cracking, these long-chain molecules can be converted into petrol (short-chain alkanes). 2. Cracking is used to produce short-chain alkenes. long-chain alkane—>shorter-chain alkane + alkene Short-chain alkenes such as ethene and propene are used as starting materials for making ethanol and plastics. 3. Cracking is used to produce hydrogen. Hydrogen is a by-product in the cracking of alkanes.

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(a) elimination of HX from a halogenoalkane by ethanolic NaOH and heat (b) dehydration of an alcohol, by using a heated catalyst (e.g. Al_2O_3) or a concentrated acid (c) cracking of a longer chain alkane

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(i) hydrogen in a hydrogenation reaction, H_2 (g) and Pt/Ni catalyst and heat

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(iii) a hydrogen halide, HX(g) at room temperature

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A (d) addition polymerisation exemplified by the reactions of ethene and propene describe the use of aqueous bromine to show the presence of a C=C bond describe the mechanism of electrophilic addition in alkenes, using bromine / ethene and hydrogen bromide /…

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(c) substitution of an alcohol, e.g. by reaction with HX or KBr with H_2SO_4 or H_3PO_4; or with PCl_3 and heat; or with PCl_5; or with $SOCl_2$ 2 classify halogenoalkanes into primary, secondary and tertiary 3 describe the following nucleophilic substitution react…

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A (d) the reaction with aqueous silver nitrate in ethanol as a method of identifying the halogen present as exemplified by bromoethane

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A …ure of the two, depending on structure describe and explain the different reactivities of halogenoalkanes (with particular reference to the relative strengths of the C–X bonds as exemplified by the reactions of halogenoalkanes with aqueous silver nitrates)

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(f) formation of esters by reaction with carboxylic acids and concentrated H_2SO_4 or H_3PO_4 as catalyst as exemplified by ethanol

Not HCl probably

31

…r change from orange to green deduce the presence of a $CH_3CH(OH)$– group in an alcohol, $CH_3CH(OH)$–R, from its reaction with alkaline I_2 (aq) to form a yellow precipitate of triiodomethane and an ion, RCO_2^- explain the acidity of alcohols compared with water

31

(d) esterification with alcohols with concentrated H_2SO_4 as catalyst

Probably not HCl

32

(e) reduction by $LiAlH_4$ to form a primary alcohol

32

Formation of esters requires **concentrated H2SO4** catalyst but hydrolysis of esters require **dilute acid or alkali**

32

(a) the condensation reaction between an alcohol and a carboxylic acid with concentrated H_2SO_4 as catalyst 2 describe the hydrolysis of esters by dilute acid and by dilute alkali and heat

32

…of an addition polymer obtained from a given monomer 3 identify the monomer(s) present in a given section of an addition polymer molecule 4 recognise the difficulty of the disposal of poly(alkene)s, i.e. non-biodegradability and harmful combustion products

33

1.1 × abundance of M^+ ion 6 deduce the presence of bromine and chlorine atoms in a compound using the $M+2$ peak

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