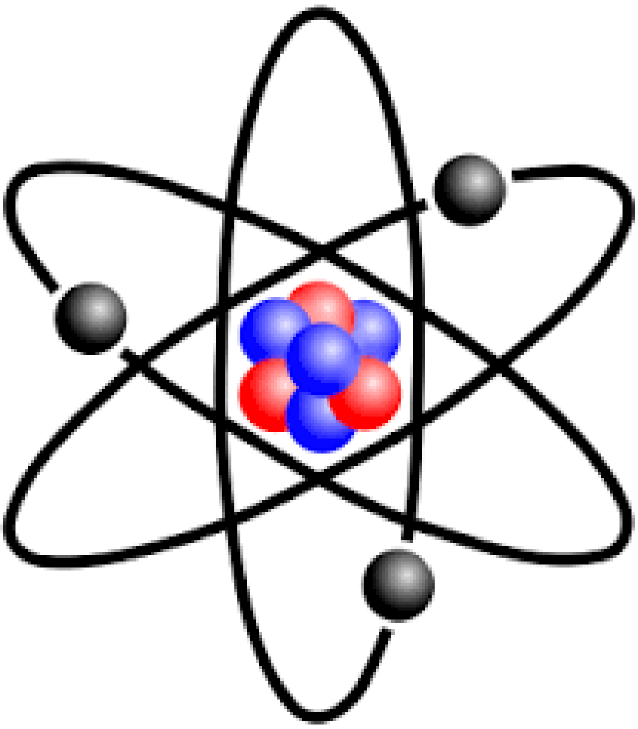
Core essential knowledge for triple chemistry topics paper 3 and 4



Student name

____

Please note that paper 4 will include topic 1 core knowledge. You will need to learn the core knowledge for topic 1 for each of your exams.

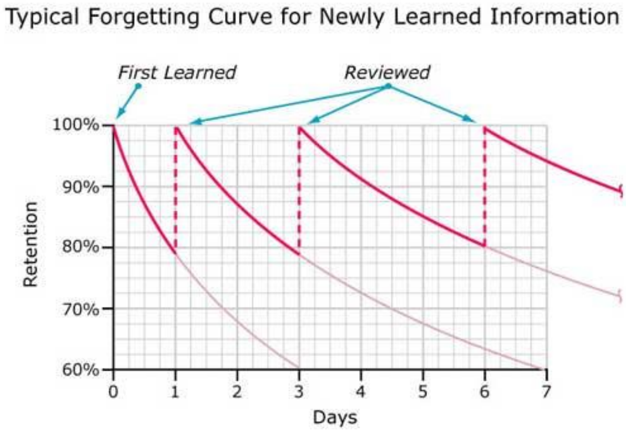
Contents

Chemistry code of success

In order to be successful we need to work together and support each other. The code of success is designed to share with you the things that need to happen for you to be successful. Success is not based on intelligence. Success is based on hard work, practice and repetition, learning from mistakes and a desire to be the best you can.

Core knowledge is a crucial part of your chemistry studies.
Students that have a good recall of their core knowledge have a greater chance of success in chemistry.

This booklet contains the core knowledge that you must learn during the next two years of your studies. There is also a quizlet that you can use to practice your core knowledge. This can be found with Quizlet by searching for buildyouriceberg.

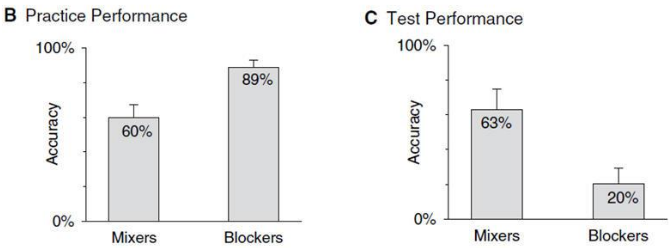


To help students learning and remember these facts the following strategies are recommended:

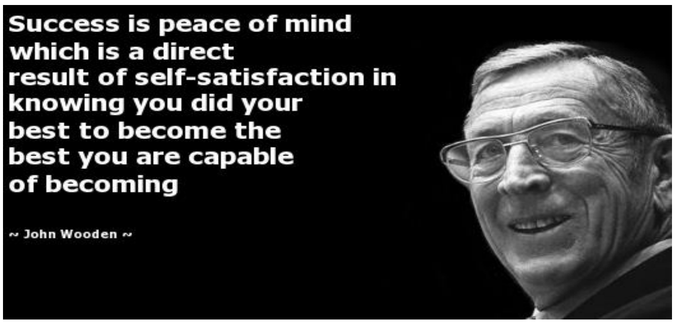
a) *Regular quizzing* - Three 15-20 minute sessions a week would help greatly.

b) *Spacing* – leaving time between each sessions will allow students to forget information. This forgetting and relearning strengthens the recall of information.

c) *Interleaving* – cover a different topic in each revision session helps to strengthen a student's memory of the core knowledge.



Parents/carers if you are able to support your child with the learning of this material then that is a great help. You can quiz them by asking them the questions within the book. The answers have been included to support the use of this booklet.
Remember that your success in chemistry is your responsibility. Your success will be based on how hard you work.



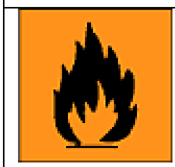
Paper 3 – Topic 0 key concepts

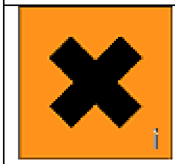
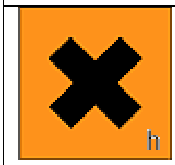
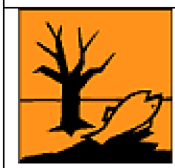
1. Complete the table below to show the formulas for the elements, compounds and ions shown.

Elements	**Formula**	**Compounds**	**Formula**	**Ions**	**Formula**
Hydrogen		Water		Hydrogen	
Sodium		Carbon dioxide		Chloride	
Chlorine		Hydrochloric acid		Sodium	
Nitrogen		Sulfuric acid		Hydroxide	
Oxygen		Nitric acid		Carbonate	
Helium		Sodium chloride		Sulfate	
Carbon		Magnesium oxide		Nirate	
Aluminium		Aluminium oxide		Oxide	

2. Give two reasons why hazard symbols are included on chemical containers.
   i)____
   ii) ____

3. Complete the table below to show the name, description and precautions which need to be taken for each hazard symbol.

Symbol	Name	Description	Precautions
			
			

			
i			
h			
			
			
			

4. Complete the table to show the number and type of each atom in the following formulas.

	Number and type of atoms
CO_2	
$4CO_2$	
$5CO_2$ + $4H_2O$	
$Co(OH)_2$	
$Al_2(SO_4)_3$	
$4C_2H_5OH$ + $12O_2$	

5. Balance the following equations

1. Ca + HF $\rightarrow$ CaF_2 + H_2	2. NaBr + Cl_2 $\rightarrow$ NaCl + Br_2
3. Zn + H_2O $\rightarrow$ ZnO + H_2	4. Li + O_2 $\rightarrow$ Li_2O
5. CH_4 + O_2 $\rightarrow$ CO_2 + H_2O	6. Al_2O_3 $\rightarrow$ Al + O_2
7. N_2 + H_2 $\rightarrow$ NH_3	9. Fe_2O_3 + CO $\rightarrow$ Fe + CO_2
10. Li + H_2O $\rightarrow$ LiOH + H_2	11. CO_2 + H_2O $\rightarrow$ $C_6H_{12}O_6$ + O_2

6. What do each of the state mean?

State symbol	meaning
s	
l	
g	
aq	

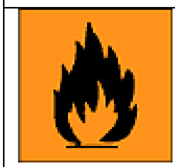
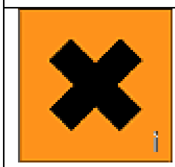
Paper 3 – Topic 0 key concepts answers

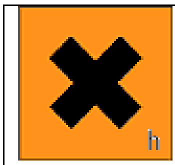
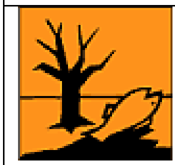
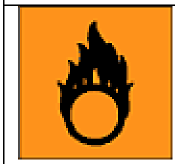
1. Complete the table below to show the formulas for the elements, compounds and ions shown.

Elements	Formula	Compounds	Formula	Ions	Formula
Hydrogen	H	Water	H_2O	Hydrogen	H^+
Sodium	Na	Carbon dioxide	CO_2	Chloride	Cl^-
Chlorine	Cl	Hydrochloric acid	HCl	Sodium	Na^+
Nitrogen	N	Sulfuric acid	H_2SO_4	Hydroxide	OH^-
Oxygen	O	Nitric acid	HNO_3	Carbonate	CO_3^{2-}
Helium	He	Sodium chloride	$NaCl$	Sulfate	SO_4^{2-}
Carbon	C	Magnesium oxide	MgO	Nirate	NO_3^-
Aluminium	Al	Aluminium oxide	Al_2O_3	Oxide	O^{2-}

2. Give two reasons why hazard symbols are included on chemical containers.
   i) to show everyone the potential dangers associated with the chemical inside the container.
   ii) to show everyone the precautions they need to take with the chemical in the laboratory.

3. Complete the table below to show the name, description and precautions which need to be taken for each hazard symbol.

Symbol	Name	Description	Precautions
	*Explosive*	*They could explode*	*Keep these materials away from heat and fire.*
	*Highly flammable*	*Catches fire easily*	*Keep away from sparks and oxidising agents. Wear eye protection*
	*Toxic*	*Cause death – if swallowed, breathed in or absorb through skin*	*Wear gloves, wear mask, wear eye protection, use a fume cupboard*
	*irritant*	*Irritates skin, eyes, lungs*	*Wear gloves, eye protection, handle with care and wipe up any spills*

	harmful	Similar to toxic but not as dangerous	Handle with care, gloves, eye protection, immediately wash off spills from skin
	*corrosive*	*Eat away at objects*	*Eye protection, gloves, wipe up spills, handle with care*
	*Hazardous to the environment*	*Harms or kills living things in the environment*	*Don't dispose of these chemicals down the drain. Dispose carefully in a controlled way.*
	*Oxidising agents*	*Provide oxygen which makes other substances burn more fiercely*	*Keep away from highly flammable chemicals, eye protection*

4. Complete the table to show the number and type of each atom in the following formulas.

	Number and type of atoms
CO_2	*1 x carbon, 2 x oxygen*
$4CO_2$	*4 x carbon, 8 x oxygen*
$5CO_2 + 4H_2O$	*5 x carbon, 10 x oxygen and 8 x hydrogen, 4 x hydrogen*
$Co(OH)_2$	*1 x cobalt, 2 x oxygen, 2 x hydrogen*
$Al_2(SO_4)_3$	*2 x aluminium, 3 x sulfur, 12 x oxygen*
$4C_2H_5OH + 12O_2$	*8 x carbon, 24 x hydrogen, 4 x oxygen and 24 x oxygen*

5. Balance the following equations

8. $Ca + HF \rightarrow CaF_2 + H_2$ $Ca + 2HF \rightarrow CaF_2 + H_2$	9. $NaBr + Cl_2 \rightarrow NaCl + Br_2$ $2NaBr + Cl_2 \rightarrow 2\,NaCl + Br_2$
10. $Zn + H_2O \rightarrow ZnO + H_2$ $Zn + H_2O \rightarrow ZnO + H_2$	11. $Li + O_2 \rightarrow Li_2O$ $4Li + O_2 \rightarrow 2\,Li_2O$
12. $CH_4 + O_2 \rightarrow CO_2 + H_2O$ $CH_4 + 2\,O_2 \rightarrow CO_2 + 2H_2O$	13. $Al_2O_3 \rightarrow Al + O_2$ $2Al_2O_3 \rightarrow 4\,Al + 3O_2$

14. N_2 + H_2 → NH_3 *N_2 + $3H_2$ → $2NH_3$*	12. Fe_2O_3 + CO → Fe + CO_2 *Fe_2O_3 + $3\ CO$ → $2Fe$ + $3\ CO_2$*
13. Li + H_2O → $LiOH$ + H_2 *$2Li$ + $2\ H_2O$ → $2\ LiOH$ + H_2*	14. CO_2 + H_2O → $C_6H_{12}O_6$ + O_2 *$6CO_2$ + $6H_2O$ → $C_6H_{12}O_6$ + $6\ O_2$*

6. What do each of the state mean?

State symbol	**meaning**
s	*Solid*
l	*Liquid*
g	*Gas*
aq	*Aqueous, a solution formed when a solid dissolves in water*

Paper 3 – Topic 1 atomic structure

1. What discovery caused the Dalton model of the atom to change over time?

____

____

2. Draw the structure of an atom of lithium. Include the correct number of protons, neutrons and electrons and label the nucleus and energy levels.

Diagram of labelled atom

3. Complete the following table

Particle	Location in the atom	Mass	Charge
Proton			
Neutron			
Electron			

4. Explain why atoms have no overall charge

____

____

5. The nucleus of an atom is very _____________ compared to the overall size of the atom.

6. Where is most of the mass of an atom found?

____

7. Describe the meaning of the following terms

Atomic number	
Mass number	
Relative atomic mass	

8. What is the definition of an atom?

____

____

9. What is the definition of an element?

____

____

10. What is the definition for an isotope?

____

____

11. Calculate the number of protons, neutrons and electrons in two isotopes of carbon.

Atom		Atom
12		14
C		C
6		6
	protons	
	Neutrons	
	electrons	

12. What is the formula used to calculate the relative atomic mass of an element from the relative mass and abundance of its isotope?

____

____

13. The mass and abundance of boron's isotopes are 19.9% boron-10 and 80.1% boron-11. Use this information to calculate the relative atomic mass for boron.

____

____

14. Mendeleev used the _________________ of the elements to arrange them in the periodic table. He left _________________ because he knew that some elements had not yet been discovered. He made _________________ about the undiscovered elements and these were very close to the actual properties of that element.

15. Mendeleev believed that he had arranged the elements in order of _______________________. This was not always true due to the ____.

16. The periodic table contains elements in horizontal rows called _________________ and vertical columns called ________________ which have ________________ chemical properties.

17. The metals are on the ___________________ of the periodic table whilst the non-metals are on the ___________________.

18. Describe the relationship between an atoms location in the periodic table and it electron structure. You should use calcium (2, 8, 8, 2) as your example and refer to the period and group which calcium is found in. (2)

____

____

____

19. Complete the electron configurations for the elements in the table.

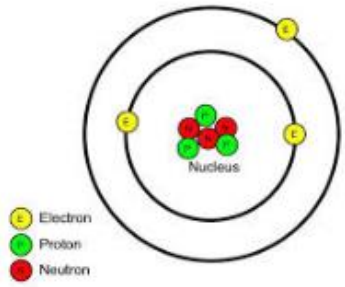
Element	Electron structure
Carbon	
Sodium	
Argon	
Potassium	

Paper 3 – Topic 1 atomic structure answers

1. What discovery caused the Dalton model of the atom to change over time?
   The discovery of sub atomic particles – electrons, protons and neutrons.

2. Draw the structure of an atom of lithium. Include the correct number of protons, neutrons and electrons and label the nucleus and energy levels.

Diagram of labelled atom



3. Complete the following table

Particle	Location in the atom	Mass	Charge
Proton	***nucleus***	***1***	***+1***
Neutron	***nucleus***	***1***	***0***
Electron	***Energy level***	***1/1840***	***-1***

4. Explain why atoms have no overall charge (1)
   they have the same number of positive protons as they have negative electrons.

5. The nucleus of an atom is very ***small*** compared to the overall size of the atom.

6. Where is most of the mass of an atom found?
   inside the nucleus

7. Describe the meaning of the following terms

Atomic number	***The number of protons in the nucleus***
Mass number	***The number of protons and neutrons***
Relative atomic mass	***This is the average mass of the atoms (or isotopes) in an element relative to 1/12th the mass of carbon-12.***

8. What is the definition of an atom?
   an atom is the simplest particle of an element which has the same chemical properties as the element

9. What is the definition of an element?
   an element is a substance which contains only one type of atom

10. What is the definition for an isotope?
   Isotopes are atoms of the same element which have different numbers of neutrons but the same number of protons.

11. Calculate the number of protons, neutrons and electrons in two isotopes of carbon.

Atom		Atom
12		14
C		C
6		6
6	protons	***6***
6	Neutrons	***8***
6	electrons	***6***

12. What is the formula used to calculate the relative atomic mass of an element from the relative mass and abundance of its isotope?

Relative atomic mass	$$= \frac{\%\ \text{abundance atom 1}}{100} \times \text{atomic mass atom 1} + \frac{\%\ \text{abundance atom 2}}{100} \times \text{atomic mass atom 2}$$

13. The mass and abundance of boron's isotopes are 19.9% boron-10 and 80.1% boron-11. Use this information to calculate the relative atomic mass for boron.
   relative atomic mass $= \frac{19.9}{100} \times 10 + \frac{80.1}{100} \times 11 = 10.8$

14. Mendeleev used the ***properties*** of the elements to arrange them in the periodic table. He left ***gaps*** because he knew that some elements had not yet been discovered. He made ***predictions*** about the undiscovered elements and these were very close to the actual properties of that element.

15. Mendeleev believed that he had arranged the elements in order of increasing ***relative atomic mass***. This was not always true due to the ***relative abundance of isotopes for some pairs of elements***.

16. The periodic table contains elements in horizontal rows called ***periods*** and vertical columns called ***groups*** which have **similar** chemical properties.

17. The metals are on the ***left*** of the periodic table whilst the non-metals are on the ***right***.

18. Describe the relationship between an atoms location in the periodic table and it electron structure. You should use calcium (2, 8, 8, 2) as your example and refer to the period and group which aluminium is found in. (2)
calcium is in group 2 because it has two electrons in the outer energy level. Calcium is in the fourth period because it has four energy levels.

19. Complete the electron configurations for the elements in the table.

Element	Electron structure
Carbon	***2, 4***
Sodium	***2, 8, 1***
Argon	***2, 8, 8***
Potassium	***2, 8, 8, 1***

<u>Paper 3 – Topic 1 ionic bonding</u>

1. How is an ionic bond formed?

____

____

2. What is meant by

   a) cation ____

   b) anion ____

3. What type of ion will an atom form if it gains electrons? Explain your answer.

____

____

4. What type of ion will an atom form if it loses electrons? Explain your answer.

____

5. What is the definition of an ion?

____

6. Complete the following table for the ions listed.

Ion	Formula of ion	Mass number of atom	Atomic number for atom	Number of			Electron arrangement
				Protons	Neutrons	electrons	
Sodium		23	11				
Magnesium		24	12				
Oxide		16	8				
Chloride		35	17				
Beryllium		9	4				
Fluoride		19	9				
Sulfide		32	16				
Potassium		39	19				
Lithium		7	3				
Aluminium		27	13				
Nitride		14	7				

7. Draw diagrams to show the electron arrangement for each stage of the reaction between sodium and chlorine.

Sodium atom before bonding	Chlorine atom before bonding	Sodium ion after bonding	Chloride ion after bonding

8. Explain why the sodium ion and chloride ion become charged.

____

____

9. Describe how an ionic compound is held together.

____

____

10. Use the information below to find the formula of the compounds in the table.

Information Mg^{2+} O^{2-} Na^{+} F^{-} CO_3^{2-} OH^{-} SO_4^{2-}			
Sodium fluoride	Magnesium fluoride	Sodium carbonate	Magnesium oxide
Sodium sulfate	Magnesium sulfate	Sodium oxide	Magnesium hydroxide

11. Draw a diagram of an ionic lattice and describe what is meant by the phrase 'ionic lattice'.

____

____

12. Explain why sodium chloride has a high melting and boiling point. (2)

____

____

13.Complete the following table

	Does it conduct electricity	Explanation
Solid Sodium chloride		
Molten Sodium chloride		
Dissolved sodium chloride		

Paper 3 – Topic 1 ionic bonding answers

1. How is an ionic bond formed?
   Metals atoms transfer electrons to non-metal atoms.

2. What is meant by
   a) cation ***a positively charged metal ion***
   b) anion ***a negatively charged non-metal ion***

3. What type of ion will an atom form if it gains electrons? Explain your answer.
   It will form a negative ion because electrons are negative

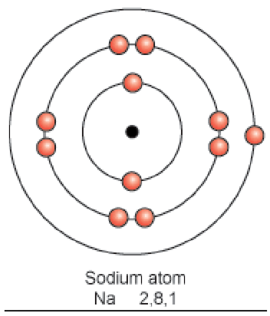
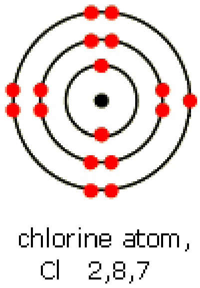
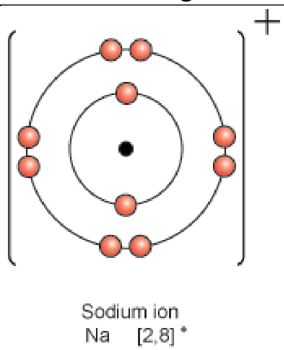
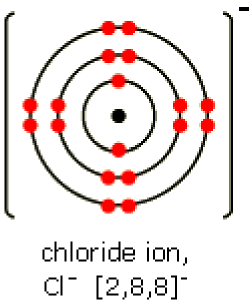
4. What type of ion will an atom form if it loses electrons? Explain your answer.
   it will form a positive ions as it loses negative electrons

5. What is the definition of an ion?
   A charged particle

6. Complete the following table for the ions listed.

Ion	Formula of ion	Mass number of atom	Atomic number for atom	Number of			Electron arrangement
				Protons	Neutrons	electrons	
Sodium	Na^{+}	23	11	***11***	***12***	***10***	***2, 8***
Magnesium	Mg^{2+}	24	12	***12***	***12***	***10***	***2, 8***
Oxide	O^{2-}	16	8	***8***	***8***	***10***	***2, 8***
Chloride	Cl^{-}	35	17	***17***	***18***	***18***	***2, 8, 8***
Beryllium	Be^{2+}	9	4	***4***	***5***	***2***	***2***
Fluoride	F^{-}	19	9	***9***	***10***	***10***	***2, 8***
Sulfide	S^{2-}	32	16	***16***	***16***	***18***	***2, 8, 8***
Potassium	K^{+}	39	19	***19***	***20***	***18***	***2, 8, 8***
Lithium	Li^{+}	7	3	***3***	***4***	***2***	***2***
Aluminium	Al^{3+}	27	13	***13***	***14***	***10***	***2, 8***
Nitride	N^{3-}	14	7	***7***	***7***	***10***	***2, 8***

7. Draw diagrams to show the electron arrangement for each stage of the reaction between sodium and chlorine.

Sodium atom before bonding	Chlorine atom before bonding	Sodium ion after bonding	Chloride ion after bonding
			

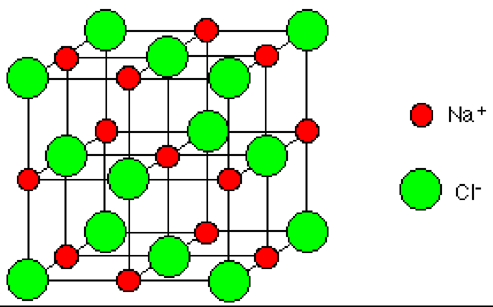
8. Explain why the sodium ion and chloride ion become charged.
The sodium ion has a positive charge because it has lost electrons
The chloride ion has a negative charge because it has gained an electron

9. Describe how an ionic compound is held together.
there is an electrostatic attraction between the positive sodium ions and negative chloride ions

10. Use the information below to find the formula of the compounds in the table.

Information Mg^{2+} O^{2-} Na^{+} F^{-} CO_3^{2-} OH^{-} SO_4^{2-}			
Sodium fluoride ***NaF***	Magnesium fluoride ***MgF_2***	Sodium carbonate ***Na_2CO_3***	Magnesium oxide ***MgO***
Sodium sulfate ***Na_2SO_4***	Magnesium sulfate ***$MgSO_4$***	Sodium oxide ***Na_2O***	Magnesium hydroxide ***$Mg(OH)_2$***

11. Draw a diagram of an ionic lattice and describe what is meant by the phrase 'ionic lattice'.



An ionic lattice is a regular repeated pattern of positive and negative ions which are held together by electrostatic attractions.

12.Explain why sodium chloride has a high melting and boiling point. (2)

The electrostatic attraction between the ions is strong and requires a lot of energy to break

13.Complete the following table

	Does it conduct electricity	Explanation
Solid Sodium chloride	***No***	***Ions are held in a lattice by an electrostatic attraction. They are not able to move***
Molten Sodium chloride	***Yes***	***The electrostatic attractions have been broken and the ions are free to move.***
Dissolved sodium chloride	***Yes***	***The electrostatic attractions have been broken and the ions are free to move.***

Paper 3 – topic 1 covalent bonding

1. What is a covalent bond?

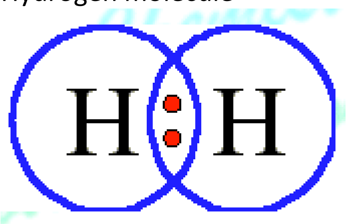
   ____

   ____

2. What is formed when a covalent bond is formed between atoms?

   ____

3. Draw diagrams to show the covalent bonds in the following molecules. Hydrogen has been completed for you.

Hydrogen molecule 	Hydrogen chloride
Water	Methane
Oxygen	Carbon Dioxide

4. Complete the table below to explain why simple covalent compounds have the following properties:

Property	Reason
Low melting and boiling point	
Poor conductors of electricity	

5. Diamond and graphite are different forms of ______________________. They are examples of ____________________________ molecules.

6. Draw a diagram to show the structure of diamond and graphite which are both giant covalent substances.

Diamond	Graphite

7. Explain how the structures of diamond and graphite can be used to explain the following properties.

Diamond	Graphite
Very high melting and boiling point *__Reason__*	
Hard substance *__Reason__*	*Softer substance* *__Reason__*
Will not conduct electricity *__Reason__*	*Will conduct electricity* *__Reason__*

8. Explain why diamond and graphite have the following uses.

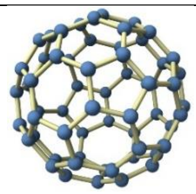
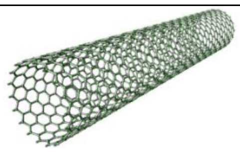
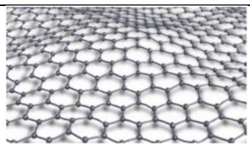
Substance	**Use**	**Reason**
Diamond	Cutting tools	
Graphite	Electrodes	
	lubricant	

9. What are fullerenes?

____

____

10. Complete the table below to show the structures and properties of the fullerenes.

Fullerene name	Diagram	Structure, bonding	Properties	Uses
Buckminster				
Nanotube				
Graphene				

11. Most metals are _______________ solids and have __________ melting points, ___________ density and are ____________ conductors of electricity. Non-metals have _______________ boiling points and are ___________ conductors of electricity.

12.Draw a labelled diagram to show the structure of a metal.

13.Use the structure of metals to explain why metals have the following properties

Property	Reason
Malleable	
Conduct electricity	

14.Complete the table below.

Substance	Melting point (°C)	Boiling point (°C)	Solubility in water	Does it conduct electricity as a solid?	Does it conduct electricity in solution?	Type of bonding in substance
Sodium chloride						
Magnesium sulfate						
Hexane						
Liquid paraffin						
Silicon (IV) oxide (sand)						
Copper Sulfate						
Sucrose (sugar)						

Paper 3 – topic 1 covalent bonding answers

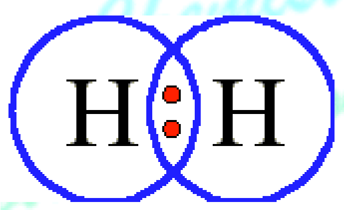
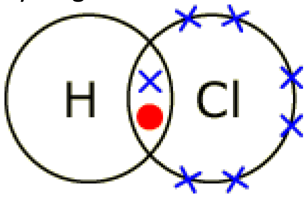
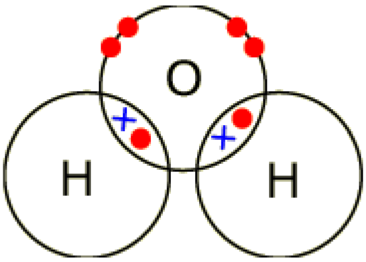
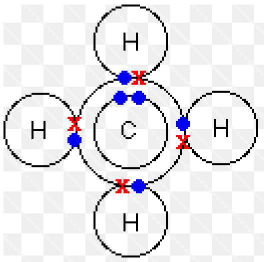
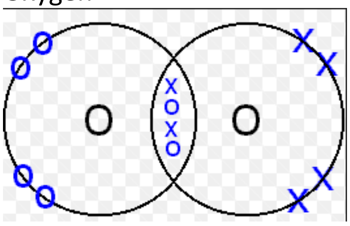
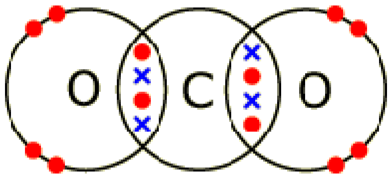
1. What is a covalent bond?

   A covalent bond is formed when two atoms share a pair of electrons.

2. What is formed when a covalent bond is formed between atoms?

   A molecule

3. Draw diagrams to show the covalent bonds in the following molecules. Hydrogen has been completed for you.

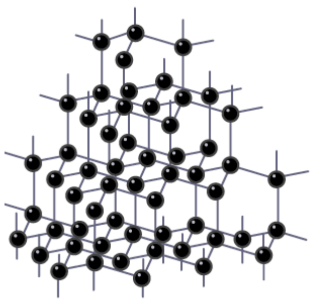
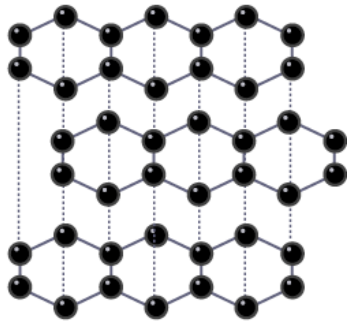
Hydrogen molecule	Hydrogen chloride
	
Water 	Methane 
Oxygen 	Carbon Dioxide 

4. Complete the table below to explain why simple covalent compounds have the following properties:

Property	Reason
Low melting and boiling point	***There are weak forces of attraction between molecules which do not require a lot of energy to break them.***
Poor conductors of electricity	***There are no free electrons because they are used in bonding.***

5. Diamond and graphite are different forms of ***carbon***. They are examples of ***giant covalent*** molecules.

6. Draw a diagram to show the structure of diamond and graphite which are both giant covalent substances.

Diamond	Graphite
	

7. Explain how the structures of diamond and graphite can be used to explain the following properties.

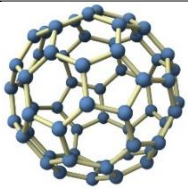
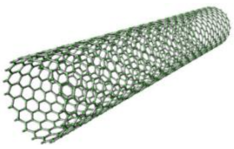
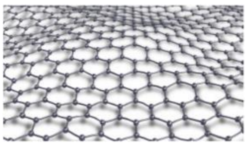
Diamond	Graphite
Very high melting and boiling point Reason ***There are lots of strong covalent bonds which need to be broken. This requires a lot of energy.***	
Hard substance Reason ***The structure of the atoms is a rigid structure which stops the atoms from moving.***	*Softer substance* Reason ***The atoms are arranged in layers. The layers are held together by weak forces of attraction which are means the layers can slide.***
Will not conduct electricity Reason ***Each carbon atom has 4 bonds so there are no free electrons.***	*Will conduct electricity* Reason ***Each carbon atom has 3 bonds to there are free, delocalised electrons which can move.***

8. Explain why diamond and graphite have the following uses.

Substance	**Use**	**Reason**
Diamond	Cutting tools	***Diamond is a hard material due to its rigid structure.***
Graphite	Electrodes	***Each carbon has three bonds so there are free electrons which can move and conduct electricity.***
	lubricant	***The layers can slide as they are only held together by weak forces of attraction.***

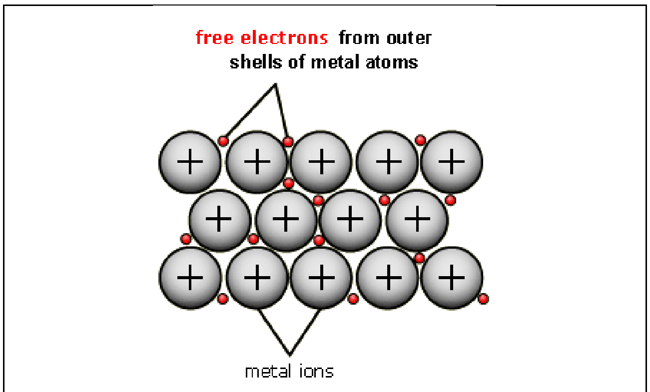
9. What are fullerenes?
A fullerene is a simple carbon molecule in which each carbon has three covalent bonds.

10. Complete the table below to show the structures and properties of the fullerenes.

Fullerene name	Diagram	Structure, bonding	Properties	Uses
Buckminster		***Each carbon has three bonds its formula is C_{60} The molecule is a ball shape***		
Nanotube		***The graphene is rolled to form the nanotube***	***Very strong Good conductor of electricity very light***	***Sports equipment e.g. tennis rackets Semiconductors in electrical circuits***
Graphene		***It is a sheet of carbon atoms The sheet is just one atom thick***	***It is very light it is very strong It is a good conductor of electricity***	

11. Most metals are ***shiny*** solids and have ***high*** melting points, ***high*** density and are ***good*** conductors of electricity. Non-metals have ***low*** boiling points and are ***poor*** conductors of electricity.

12.Draw a labelled diagram to show the structure of a metal.



13.Use the structure of metals to explain why metals have the following properties

Property	Reason
Malleable	***The layers of ions are able to slide***
Conduct electricity	***The delocalised electrons are able to move about to carry the electrical current***

14.Complete the table below.

Substance	Melting point (°C)	Boiling point (°C)	Solubility in water	Does it conduct electricity as a solid?	Does it conduct electricity in solution?	Type of bonding in substance
Sodium chloride	***High***	***High***	***Yes***	***No***	***Yes***	***Ionic***
Magnesium sulfate	***High***	***High***	***Yes***	***No***	***Yes***	***Ionic***
Hexane	***Low***	***Low***	***No***	***No***	***No***	***Simple covalent***
Liquid paraffin	***Low***	***Low***	***No***	***No***	***No***	***Simple covalent***
Silicon (IV) oxide (sand)	***v.high***	***v.high***	***No***	***No***	***No***	***Giant covalent***
Copper Sulfate	***Decomposes***	***NA***	***Yes***	***No***	***Yes***	***Ionic***
Sucrose (sugar)	***Low***	***Low***	***Yes***	***No***	***No***	***Simple covalent***

Paper 3 – Topic 1 masses and calculations

1. Calculate the relative formula mass for the following compounds

Compound	Relative formula mass		Compound	Relative formula mass
Methane, CH_4			ethanol, CH_3CH_2OH	
Sodium chloride, $NaCl$			iron(III) sulphate, $Fe_2(SO_4)_3$	
Ammonia, NH_3			lead nitrate, $Pb(NO_3)_2$	
Sodium hydroxide, $NaOH$			aluminium sulfate, $Al_2(SO_4)_3$	
Magnesium hydroxide, $Mg(OH)_2$			sulphur dioxide, SO_2	
Sulfuric acid, H_2SO_4			iron(III) oxide, Fe_2O_3	

2. What is the definition of one mole?

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3. Calculate the mass for the number of moles of each substance below.

a) 0.8 moles of nitric acid, HNO_3?	b) 0.25 moles of sulphur molecules, S_8?
c) 0.175 mole of sulphate ions, SO_4^{2-}?	d) 4 mole of iron (II) chloride, $FeCl_2$?
e) 4 moles of manganese (II) bromide, $MnBr_2$?	f) 0.006 moles of ethane molecules, C_2H_6?

4. Calculate number of moles in the following mass of each substance.

a) beryllium in 18g?	b) iron in 32g?
c) carbon dioxide, CO_2, in 180g?	d) ethene, C_2H_4, in 75g?
e) calcium carbonate, $CaCO_3$, in 340g?	

5. What number is Avogadro's constant equal to (hint – this is equal to the number of particles in one mole of any substance)?

____

6. How many oxygen atoms are there in the following substances?

a) Two moles of oxygen molecules, O_2.	b) One mole of sulfur dioxide, SO_2.
c) Four moles of sulfuric acid, H_2SO_4.	d) Fours mole of water, H_2O.

7. How many particles are there in the following substances?

e) chloride ions in one of magnesium chloride, $MgCl_2$.	f) sulfur atoms in two moles of sulfur, S_8.
g) sodium ions in one mole of sodium carbonate, Na_2CO_3.	h) sulfate ions in three moles of aluminium sulfate, $Al_2(SO_4)_3$.

8. How many moles are there in each of the following?

i) 2.847×10^{23} atoms of iron.	j) 3.433×10^{24} molecules of water.
k) 5.317×10^{24} atoms of carbon.	l) 6.134×10^{23} molecules of ethanol.
m) 4.327×10^{24} atoms of sodium.	n) 7.421×10^{24} molecules of carbon dioxide.

9. What is the definition for the empirical formula?

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10. Find the empirical formula for the following molecular formulas.

Formula of molecule	C_2H_6	$C_6H_{12}O_6$	C_6H_6	$C_3H_6O_2$
Empirical formula				

11. A compound contains 0.31g of phosphorus and 1.07g of chlorine. Calculate the empirical formula based on this information.

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12. A compound contained 66.7g of carbon, 11.1g of hydrogen and 22.2g of oxygen. Its relative formula mass was 72. Calculate the empirical formula for this compound and then the molecular formula.

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13. Write the word and balanced symbol equation, including state symbols, for the reaction between solid magnesium and oxygen gas to form solid magnesium oxide.

Word equation		+		→	
Balanced symbol equation					

14. Complete the table below to show the risks and precautions you should take.

Risk	Precaution

15. Write a brief method to explain how to carry out the experiment to find the empirical formula for magnesium oxide.

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16.What does the law of conservation of mass state?

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17. Use the conservation of mass to explain what happens during a precipitation reaction.

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18. Use the conservation of mass to explain what happens during a reaction which produces a gas and is carried out in an open test tube.

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19. Calculate the mass of carbon dioxide you could obtain by adding hydrochloric acid to 15 g of calcium carbonate:

Calcium carbonate	+	Hydrochloric acid	→	Calcium chloride	+	water	+	Carbon dioxide
$CaCO_3$	+	$2HCl$	→	$CaCl_2$	+	H_2O	+	CO_2

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20. Why is the mass of product limited by the reactant which is not in excess?

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21. What is the equation used to calculate the concentration of a solution?

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22. Calculate the concentration when 2.1g of of sodium hydrogencarbonate, $NaHCO_3$, in 250cm^3 of solution?

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Paper 3 – Topic 1 masses and calculations answers

1. Calculate the relative formula mass for the following compounds

Compound	Relative formula mass		Compound	Relative formula mass
Methane, CH_4	***1xC=1x12=12*** ***4xH=4x1=4*** ***Total = 16***		ethanol, CH_3CH_2OH	***2xC=2x12=24*** ***6xH=6x1=6*** ***1xO=1x16=16*** ***Total = 46***
Sodium chloride, NaCl	***1xNa=1x23=23*** ***1xCl=1x35.5=35.5*** ***total = 58.5***		iron(III) sulphate, $Fe_2(SO_4)_3$	***2xFe=2x56=112*** ***3xS=3x32=96*** ***12xO=12x16=192*** ***Total = 400***
Ammonia, NH_3	***1xN=1x14=14*** ***3xH=3x1=3*** ***Total = 17***		lead nitrate, $Pb(NO_3)_2$	***1xPb=1x207=207*** ***2xN=2x14=28*** ***6xO=6x16=96*** ***Total = 331***
Sodium hydroxide, NaOH	***1xNa=1x23=23*** ***1xO=1x16=16*** ***1xH=1x1=1*** ***Total = 40***		aluminium sulfate, $Al_2(SO_4)_3$	***2xAl=2x27=54*** ***3xS=3x32=96*** ***12xO=12x16=192*** ***Total = 342***
Magnesium hydroxide, $Mg(OH)_2$	***1xMg=1x24=24*** ***2xO=2x16=32*** ***2xH=2x1=2*** ***Total = 58***		sulphur dioxide, SO_2	***1xS=1x32=32*** ***2xO=2x16=32*** ***Total = 64***
Sulfuric acid, H_2SO_4	***2xH=2x1=2*** ***1xS=1x32=32*** ***4xO=4x16=64*** ***Total = 98***		iron(III) oxide, Fe_2O_3	***2xFe=2x56=112*** ***3xO=3x16=48*** ***Total = 160***

2. What is the definition of one mole?

one mole of a substance contains 6.02 x 10^{23} atoms

3. Calculate the mass for the number of moles of each substance below.

a) 0.8 moles of nitric acid, HNO_3?	b) 0.25 moles of sulphur molecules, S_8?
mass = no of moles x mass 1 mole ***= 0.8 x 63*** ***= 50.4g***	***mass = no of moles x mass 1 mole*** ***= 0.25 x 63*** ***= 64g***

c) 0.175 mole of sulphate ions, SO_4^{2-}? ***mass = no of moles x mass 1 mole*** ***= 0.175 x 96*** ***= 16.8g***	d) 4 mole of iron (II) chloride, $FeCl_2$? ***mass = no of moles x mass 1 mole*** ***= 4 x 127*** ***= 508g***
e) 4 moles of manganese (II) bromide, $MnBr_2$? ***mass = no of moles x mass 1 mole*** ***= 4 x 215*** ***= 860g***	f) 0.006 moles of ethane molecules, C_2H_6? ***mass = no of moles x mass 1 mole*** ***= 0.006 x 30*** ***= 0.18g***

4. Calculate number of moles in the following mass of each substance.

a) beryllium in 18g? ***No of moles =*** $\frac{\textbf{mass}}{\textbf{Mass of 1 mole}}$ $= \frac{18}{9} = 2$	b) iron in 32g? ***No of moles =*** $\frac{\textbf{mass}}{\textbf{Mass of 1 mole}}$ $= \frac{32}{56} = 0.57$
c) carbon dioxide, CO_2, in 180g? ***No of moles =*** $\frac{\textbf{mass}}{\textbf{Mass of 1 mole}}$ $= \frac{180}{44} = 4.09$	d) ethene, C_2H_4, in 75g? ***No of moles =*** $\frac{\textbf{mass}}{\textbf{Mass of 1 mole}}$ $= \frac{75}{28} = 2.68$
e) calcium carbonate, $CaCO_3$, in 340g? ***No of moles =*** $\frac{\textbf{mass}}{\textbf{Mass of 1 mole}}$ $= \frac{340}{100} = 3.4$	

5. What number is Avogadro's constant equal to (hint – this is equal to the number of particles in one mole of any substance)?
   6.02×10^{23}

6. How many oxygen atoms are there in the following substances?

e) Two moles of oxygen molecules, O_2. ***No of atoms = number of moles x Avogadro's constant*** ***= 4 x 6.02x10^{23}*** ***= 2.4 x 10^{24} atoms***	f) One mole of sulfur dioxide, SO_2. ***No of atoms = number of moles x Avogadro's constant*** ***= 2 x 6.02x10^{23}*** ***= 1.2 x 10^{24} atoms***
g) Four moles of sulfuric acid, H_2SO_4. ***No of atoms = number of moles x Avogadro's constant*** ***= 16 x 6.02x10^{23}*** ***= 2.4 x 10^{24} atoms***	h) Fours mole of water, H_2O. ***No of atoms = number of moles x Avogadro's constant*** ***= 4 x 6.02x10^{23}*** ***= 2.4 x 10^{24} atoms***

7. How many particles are there in the following substances?

e) chloride ions in one mole of magnesium chloride, $MgCl_2$. ***No of ions = number of moles x Avogadro's constant*** ***= 2 x 6.02x10^{23}*** ***= 1.2 x 10^{24} atoms***	f) sulfur atoms in two moles of sulfur, S_8. ***No of atoms = number of moles x Avogadro's constant*** ***= 16 x 6.02x10^{23}*** ***= 9.62 x 10^{24} atoms***
g) sodium ions in one mole of sodium carbonate, Na_2CO_3. ***No of ions = number of moles x Avogadro's constant*** ***= 2 x 6.02x10^{23}*** ***= 1.2 x 10^{24} atoms***	h) sulfate ions in three moles of aluminium sulfate, $Al_2(SO_4)_3$. ***No of ions = number of moles x Avogadro's constant*** ***= 9 x 6.02x10^{23}*** ***= 5.4 x 10^{24} atoms***

8. How many moles are there in each of the following?

i) 2.847×10^{23} atoms of iron. ***No of moles =*** $\frac{\textbf{No of particles}}{\textbf{Avogadro's constant}}$ $= \frac{2.847 \times 10^{23}}{6.02 \times 10^{23}} = 0.47$	j) 3.433×10^{24} molecules of water. ***No of moles =*** $\frac{\textbf{No of particles}}{\textbf{Avogadro's constant}}$ $= \frac{3.433 \times 10^{24}}{6.02 \times 10^{23}} = 5.7$
k) 5.317×10^{24} atoms of carbon. ***No of moles =*** $\frac{\textbf{No of particles}}{\textbf{Avogadro's constant}}$ $= \frac{5.317 \times 10^{23}}{6.02 \times 10^{23}} = 8.8$	l) 6.134×10^{23} molecules of ethanol. ***No of moles =*** $\frac{\textbf{No of particles}}{\textbf{Avogadro's constant}}$ $= \frac{6.134 \times 10^{23}}{6.02 \times 10^{23}} = 1.01$
m) 4.327×10^{24} atoms of sodium. ***No of moles =*** $\frac{\textbf{No of particles}}{\textbf{Avogadro's constant}}$ $= \frac{4.327 \times 10^{24}}{6.02 \times 10^{23}} = 7.18$	n) 7.421×10^{24} molecules of carbon dioxide. ***No of moles =*** $\frac{\textbf{No of particles}}{\textbf{Avogadro's constant}}$ $= \frac{7.421 \times 10^{24}}{6.02 \times 10^{23}} = 12.3$

9. What is the definition for the empirical formula?
   the simplest ratio for the number and type of atoms in a formula

10. Find the empirical formula for the following molecular formulas.

Formula of molecule	C_2H_6	$C_6H_{12}O_6$	C_6H_6	$C_3H_6O_2$
Empirical formula	**CH_3**	**CH_2O**	**CH**	**$C_3H_6O_2$**

11. A compound contains 0.31g of phosphorus and 1.07g of chlorine. Calculate the empirical formula based on this information.

	Phosphorus	Chlorine
mass	***0.31***	***1.07***
Relative atomic mass	***31***	***35.5***
No of moles	***0.31/31 = 0.01***	***1.07/35.5 = 0.03***
Divide by smallest number	***0.01/0.01 = 1***	***0.03/0.01 = 3***
Empirical formula	***PCl_3***	

12. A compound contained 66.7g of carbon, 11.1g of hydrogen and 22.2g of oxygen. Its relative formula mass was 72. Calculate the empirical formula for this compound and then the molecular formula.

	Carbon	Hydrogen	oxygen
mass	***66.7***	***11.1***	***22.2***
Relative atomic mass	***12***	***1***	***16***
No of moles	***66.7/12 = 5.6***	***11.1/1 = 11.1***	***22.2/16=1.4***
Divide by smallest number	***5.6/1.4 = 4***	***11.1/1.4 = 8***	***1.4/1.4=1***
Empirical formula	***C_4H_8O***		

Total mass of the empirical formula = C x 4 = 12 X4 = 48

H x 8 = 1 x 8 = 8

O x 1 = 16 x 1 = 16

Total mass = 72

Because the mass of the empirical formula is the same as the mass of the molecular formula the molecule formula is also C_4H_8O

13. Write the word and balanced symbol equation, including state symbols, for the reaction between solid magnesium and oxygen gas to form solid magnesium oxide.

Word equation	***Magnesium***	***+***	***Oxygen***	→	***Magnesium oxide***
Balanced symbol equation	***2Mg (s)***	***+***	***O_2 (g)***	→	***2MgO (s)***

14. Complete the table below to show the risks and precautions you should take.

Risk	Precaution
Magnesium is flammable	***Wear safety glasses***
Magnesium burns with bright flame	***Do not look directly at the flame***

15. Write a brief method to explain how to carry out the experiment to find the empirical formula for magnesium oxide.

measure the mass of the magnesium and crucible before the experiment
Heat the crucible strongly until the magnesium has reacted fully
Carefully lift the lid of the crucible to allow oxygen in to help the magnesium to react
Once reaction has finished find the mass of the crucible and magnesium oxide after the experiment

16.What does the law of conservation of mass state?

that mass cannot be created or destroyed

17. Use the conservation of mass to explain what happens during a precipitation reaction.

The mass at the end of the reaction is the same as the mass at the start of the reaction

18. Use the conservation of mass to explain what happens during a reaction which produces a gas and is carried out in an open test tube.

The mass will be less after the experiment as the gas has escaped.

19. Calculate the mass of carbon dioxide you could obtain by adding hydrochloric acid to 15 g of calcium carbonate:

Calcium carbonate	+	Hydrochloric acid	→	Calcium chloride	+	water	+	Carbon dioxide
$CaCO_3$	+	2HCl	→	$CaCl_2$	+	H_2O	+	CO_2

Relative formula mass of calcium carbonate = 1 x Ca + 1 x C + 3 x O = 1x40 + 1x12 + 3x16 = 100
Relative formula mass of carbon dioxide = 1 x C + 2 X O = 1 x12 + 2x16 = 44

$$\frac{15}{100} = \frac{x}{44}$$

$$x = \frac{15 \times 44}{100} = 6.6g$$

20. Why is the mass of product limited by the reactant which is not in excess?

If there is not enough of a reactant then the chemical is not able to react fully

21. What is the equation used to calculate the concentration of a solution?

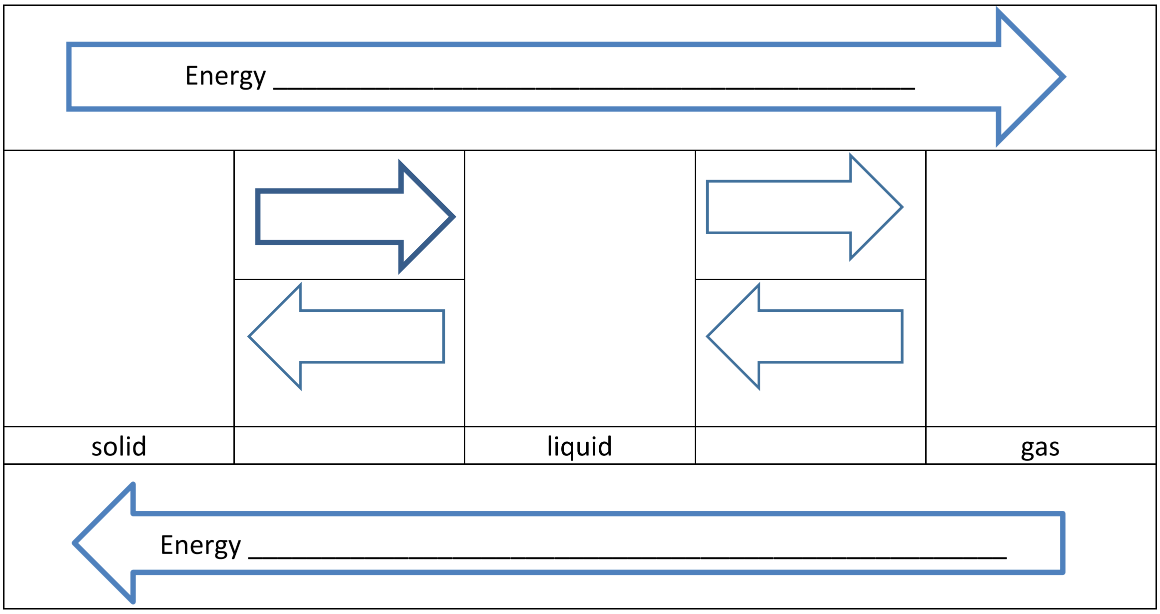
amount of solute = volume of solution x concentration

22.Calculate the concentration when 2.1g of of sodium hydrogencarbonate, $NaHCO_3$, in 250cm^3 of solution?

$$\text{Concentration} = \frac{\text{amount}}{\text{volume}} = \frac{2.1}{(250/1000)} = 8.4gdm^{-3}$$

Paper 3 – Topic 2 states of matter

1. Complete the diagram below to show the particle arrangement for solids, liquids and gases. Include the names for each change in state.



2. What is name for each change in state

   Solid to liquid ______________________

   Liquid to gas ________________________

   gas to liquid _________________________

   Liquid to solid ________________________

3. Complete the table below to show the state of matter for each substance at the temperature show.

Substance	Temperature (°C)	Melting point (°C)	Boiling point (°C)	Physical state
Iron	1647	1538	2862	
Oxygen	-230	-219	-183	
Sodium	22	98	883	
Nitrogen	-195	-210	-196	
Methane	-173	-182	-162	

4. Describe the meaning of the word pure.

   ____

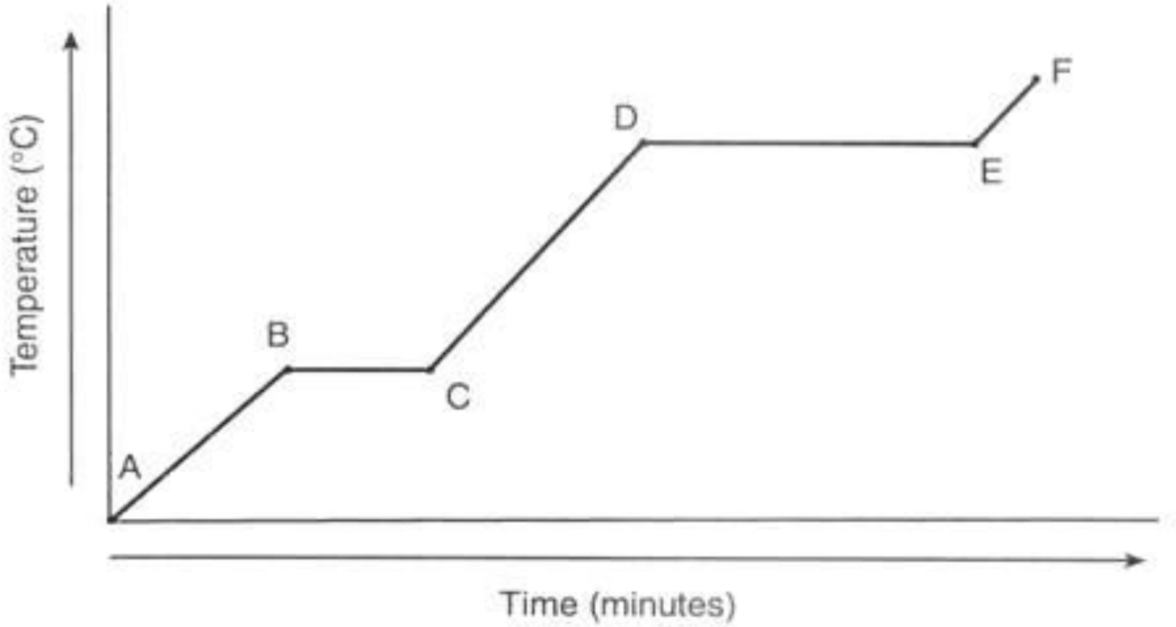
   ____

5. Describe what is meant by a mixture.

   ____

   ____

6. Label the heating curve to show when the substance is a solid, liquid and then gas.



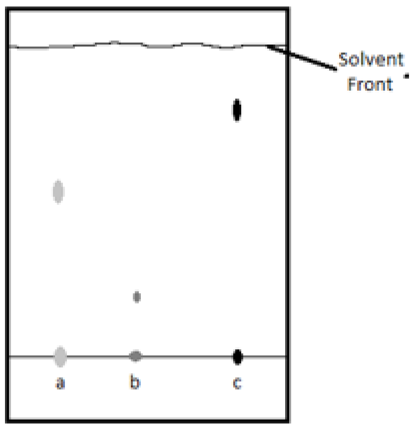
7. Add labels to the above diagram to show the melting point and boiling point for the substance.

8. Link each separation technique to the mixture it can be used to separate.

Separation technique		**Mixture**
Simple distillation		A dissolved solid where you do not want the liquid.
Fractional distillation		A mixture of soluble substances e.g. inks.
Filtration		A dissolved solid where you want to keep the liquid or 2 liquids with very different boiling points.
Crystallisation		A large sample of a mixture of liquids with similar boiling points.
Paper chromatography		An insoluble solid and a liquid.

9. Calculate the R_f value for sample a, b and c below. Show your working out.

Sample	R_f value
A	
B	
c	



10. Explain how you can tell is a substance is pure or a mixture from the chromatography results.

____

____

11. Describe the steps used in simple distillation.

____

____

____

____

12.How is fractional distillation different to simple distillation?

____

____

13.During filtration a residue and filtrate are produced. Explain the meaning of these two words.

Filtrate	
Residue	

14.Describe the steps used in crystallisation.

____

____

____

____

15.Explain the meaning of the following words which are used in chromatography.

Mobile phase	
Stationary phase	

16.Describe each of the following steps which are used to make water portable.

Sedimentation	______________________________ ______________________________ ______________________________
Filtration	______________________________ ______________________________ ______________________________
Chlorination	______________________________ ______________________________ ______________________________

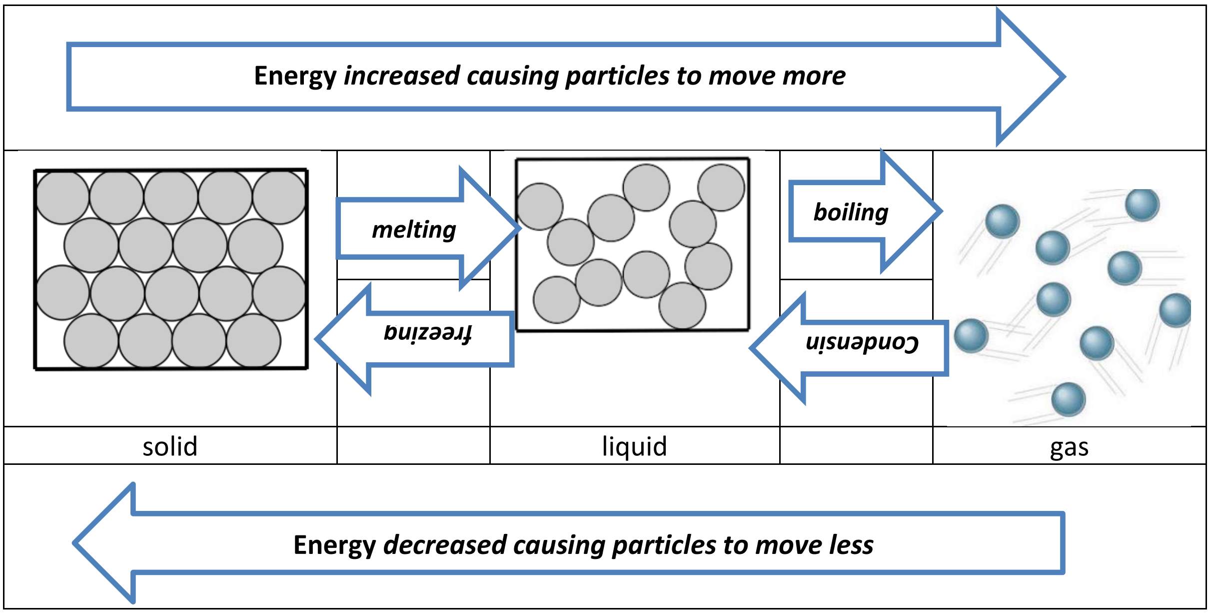
17. Describe how sea water is made portable.

____

____

Paper 3 – Topic 2 states of matter answers

1. Complete the diagram below to show the particle arrangement for solids, liquids and gases. Include the names for each change in state.



2. What is name for each change in state

Solid to liquid ***melting***

Liquid to gas ***boiling or evaporation***

gas to liquid ***condensation***

Liquid to solid ***freezing***

3. Complete the table below to show the state of matter for each substance at the temperature show.

Substance	Temperature (°C)	Melting point (°C)	Boiling point (°C)	Physical state
Iron	1647	1538	2862	***liquid***
Oxygen	-230	-219	-183	***Solid***
Sodium	22	98	883	***Solid***
Nitrogen	-195	-210	-196	***Gas***
Methane	-173	-182	-162	***liquid***

4. Describe the meaning of the word pure.
   a substance which contains only one type of particle.

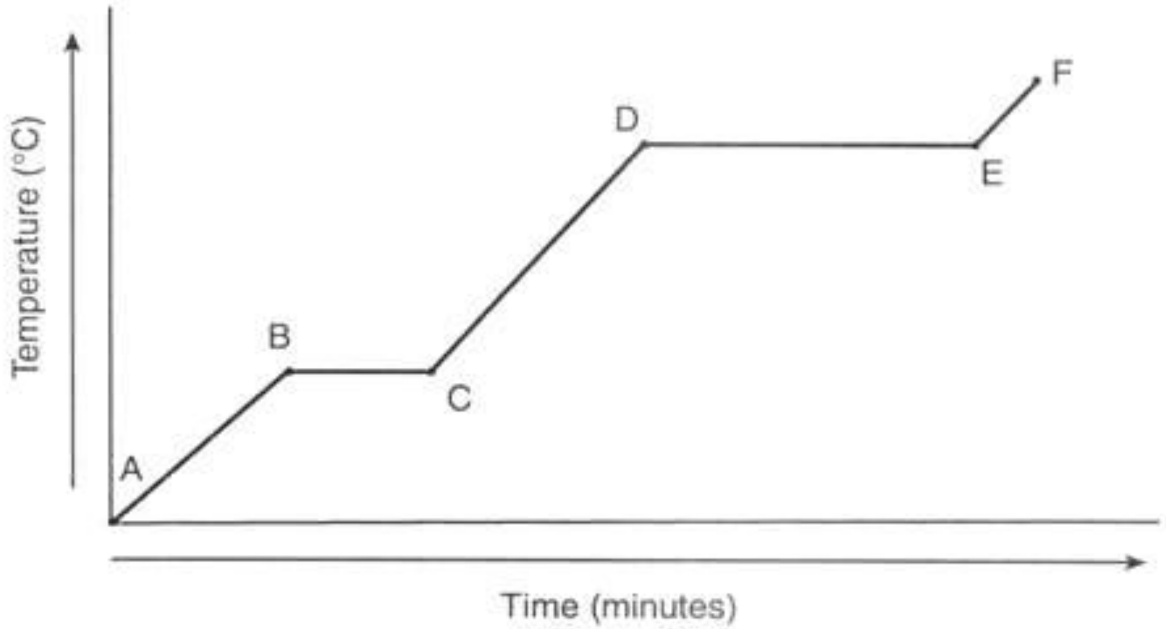
5. Describe what is meant by a mixture.
   a substance which contains two or more materials which are not chemically joined together.

6. Label the heating curve to show when the substance is a solid, liquid and then gas.

A to B ***solid state***

C to D ***Liquid state***

E to F ***Gas state***



7. Add labels to the above diagram to show the melting point and boiling point for the substance.

B to C ***melting point***

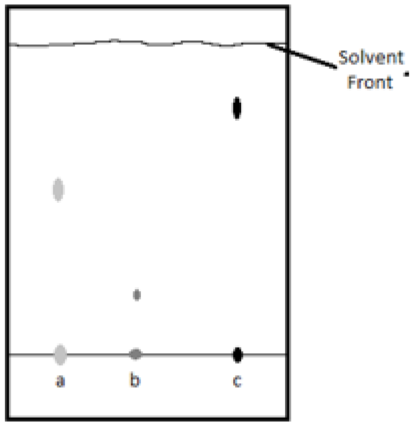
D to E ***boiling point***

8. Link each separation technique to the mixture it can be used to separate.

Separation technique		**Mixture**
Simple distillation		***A dissolved solid where you want to keep the liquid or 2 liquids with very different boiling points.***
Fractional distillation		***A large sample of a mixture of liquids with similar boiling points.***
Filtration		***An insoluble solid and a liquid.***
Crystallisation		***A dissolved solid where you do not want the liquid.***
Paper chromatography		***A mixture of soluble substances e.g. inks.***

9. Calculate the R_f value for sample a, b and c below. Show your working out.

Sample	R_f value
A	$\mathbf{R_f = \frac{2.4}{4} = 0.6}$
B	$\mathbf{R_f = \frac{0.8}{4} = 0.2}$
c	$\mathbf{R_f = \frac{3.1}{4} = 0.8}$



10. Explain how you can tell is a substance is pure or a mixture from the chromatography results.
    Pure – the substance will only produce one spot.
    Mixture – the substance will separate in to two or more spots.

11. Describe the steps used in simple distillation.
    Evaporation – the liquid with the lowest boiling point starts to evaporate and turn in to a gas.
    Condensation – The gas travels down the condenser, cools and turns in to a liquid.

12. How is fractional distillation different to simple distillation?
    An extra column is used which helps to improve the separation of the mixture.

13. During filtration a residue and filtrate are produced. Explain the meaning of these two words.

Filtrate	***This is the substance which passes through the filter paper and is collected in the conical flask.***
Residue	***This is the insoluble substance which remains in the filter paper.***

14. Describe the steps used in crystallisation.
    dissolving – the sample is dissolved to form a solution
    Heating – the solution is heated to evaporate the solution.
    Filter – filter the remaining solution to remove any insoluble substances.
    Crystallisation – leave the solution to evaporate to form crystals.

15. Explain the meaning of the following words which are used in chromatography.

Mobile phase	***This is the solution which the sample is dissolved in.***
Stationary phase	***The paper on which the sample travels.***

16.Describe each of the following steps which are used to make water portable.

Sedimentation	***The water is left so that any insoluble substances sink to the bottom and are removed.***
Filtration	***The water travels through beds of sand and gravel to remove any remaining insoluble substances.***
Chlorination	***Chlorine is added to kill microorganisms.***

17. Describe how sea water is made portable.
Distillation can be used.

Paper 3 – Topic 3 chemical change

1. Acids in solution contain _______________ ions.

2. Alkalis in solution contain _______________ ions.

3. The pH of a neutral solution is ______________.

4. Acids have a pH between _____ and _______. Strong acids have a pH of ______ and weak acids have a pH of ________.

5. Alkalis have a pH between ________ and _______. Strong alkalis have a pH of _________ and weak alkalis have a pH of ___________.

6. Complete the table below to show the colour change of indicators in acids and alkalis.

Indicator	**Colour in acidic solution**	**Colour in alkaline solution**
Litmus		
Methyl orange		
Phenolphthalein		

7. What happens to the hydrogen ion concentration as the pH of an acidic solution changes from pH 6 to pH 1?

____

8. What happens to the hydroxide ion concentration as the pH of an alkaline solution increases?

____

9. If the pH of a solution decreases by 1, by how much does the concentration of hydrogen ion increases?

____

10. What is the difference between a strong and weak acid?

____

____

11. What is the difference between a dilute and concentrated solution?

____

____

12. What is a base?

____

____

13. What is an alkali?

____

____

14. What is meant by a neutralisation reaction?

____

15.Complete the following chemical equations for the neutralisation reactions.

	Acid	+	**Base/alkali**	→	**Salt**	+	**water**
Word equation	Hydrochloric Acid	+	Sodium Hydroxide	→		+	
Balance symbol equation							
Word equation	Sulfuric Acid	+	Magnesium oxide	→		+	
Balance symbol equation							
Word equation	Nitric acid	+	Lithium hydroxide	→		+	
Balance symbol equation							

16. Complete the following neutralisation equations involving metal carbonate.

General equation	**Acid**	+	**Base/alkali**	→	**Salt**	+	**water**	+	**Carbon dioxide**
Word equation	Hydrochloric Acid	+	calcium Carbonate	→		+		+	
Balance symbol equation									
Word equation	Nitric Acid	+	Lithium Carbonate	→		+		+	
Balance symbol equation									
Word equation	Sulfuric Acid	+	Magnesium Carbonate	→		+		+	
Balance symbol equation									

17. What is the ionic equation for the neutralisation reaction?

	Acid	**+**	**Alkali**	→	**water**
Ionic equation		+		→	

18. Describe the chemical test for hydrogen gas.

____

____

19. Describe the chemical test for carbon dioxide.

____

____

20. What does the word excess mean in a chemical reaction?

____

____

21. Describe the steps to produce a soluble salt from an acid solution and an insoluble base.

____
____
____

22. Describe the steps to produce a soluble salt from an acid solution and an alkali solution (soluble).

____
____
____
____
____
____

23. Complete the table below to show the solubility of different substances.

Substance	**Soluble**	**Insoluble**
All nitrates		
Chlorides		
Sulfates		
Carbonates		
Hydroxides		

24. Explain what is meant by a precipitation reaction.

____
____

25. Complete the word and symbol equations below. Include state symbols.

Silver nitrate (aq)	+	Sodium chloride (aq)	→		+	
	+		→		+	
Iron sulfate (aq)	+	Sodium hydroxide (aq)	→		+	
	+		→		+	
barium chloride (aq)	+	Hydrochloric acid (aq)	→		+	
	+		→		+	
Copper sulfate (aq)	+	sodium hydroxide (aq)	→		+	
	+		→		+	

26. Describe how a pure, dry sample of an insoluble salt can be made.

Paper 3 – Topic 3 chemical change answers

1. Acids in solution contain ***hydrogen*** ions.

2. Alkalis in solution contain ***hydroxide*** ions.

3. The pH of a neutral solution is ***7.***

4. Acids have a pH between ***1*** and ***6***. Strong acids have a pH of ***1*** and weak acids have a pH of ***6***.

5. Alkalis have a pH between ***8*** and ***14***. Strong alkalis have a pH of ***14*** and weak alkalis have a pH of ***8***.

6. Complete the table below to show the colour change of indicators in acids and alkalis.

Indicator	**Colour in acidic solution**	**Colour in alkaline solution**
Litmus	***Red***	***Blue***
Methyl orange	***Red***	***yellow***
Phenolphthalein	***Colourless***	***pink***

7. What happens to the hydrogen ion concentration as the pH of an acidic solution changes from pH 6 to pH 1?
the hydrogen ion concentration increases

8. What happens to the hydroxide ion concentration as the pH of an alkaline solution increases?
The hydroxide ion concentration will increase

9. If the pH of a solution decreases by 1, by how much does the concentration of hydrogen ion increases?
the hydrogen ion concentration will increase by a factor of 10

10. What is the difference between a strong and weak acid?
Strong acids full dissociate into ions. Weak acids only partially dissociate

11. What is the difference between a dilute and concentrated solution?
a concentrated solution will contain a more particles in a certain volume than a dilute solution

12. What is a base?
A base is a substance that reacts with an acid to produce a salt and water

13. What is an alkali?
An alkali is a soluble base

14. What is meant by a neutralisation reaction?

It is the reaction between an acid and a base which forms a salt and water

15.Complete the following chemical equations for the neutralisation reactions.

	Acid	+	**Base/alkali**	→	**Salt**	+	**water**
Word equation	Hydrochloric Acid	+	Sodium Hydroxide	→	***Sodium Chloride***	+	***Water***
Balance symbol equation	***HCl***	+	***NaOH***	→	***NaCl***		***H_2O***
Word equation	Sulfuric Acid	+	Magnesium oxide	→	***Magnesium Sulfate***	+	***Water***
Balance symbol equation	***H_2SO_4***	+	***MgO***	→	***$MgSO_4$***		***H_2O***
Word equation	Nitric acid	+	Lithium hydroxide	→	***Lithium Nitrate***	+	***Water***
Balance symbol equation	***HNO_3***	+	***LiOH***	→	***$LiNO_3$***	+	***H_2O***

16. Complete the following neutralisation equations involving metal carbonate.

General equation	**Acid**	+	**Base/alkali**	→	**Salt**	+	**water**	+	**Carbon dioxide**
Word equation	Hydrochloric Acid	+	calcium Carbonate	→	***Calcium Chloride***	+	***Water***	+	***Carbon Dioxide***
Balance symbol equation	***2HCl***	+	***$CaCO_3$***	→	***$CaCl_2$***		***H_2O***		***CO_2***
Word equation	Nitric Acid	+	Lithium Carbonate	→	***Lithium Nitrate***	+	***Water***	+	***Carbon Dioxide***
Balance symbol equation	***$2HNO_3$***	+	***Li_2CO_3***	→	***$2LiNO_3$***	+	***H_2O***		***CO_2***
Word equation	Sulfuric Acid	+	Magnesium Carbonate	→	***Magnesium sulfate***	+	***Water***	+	***Carbon Dioxide***
Balance symbol equation	***H_2SO_4***	+	***$MgCO_3$***	→	***$MgSO_4$***	+	***H_2O***	+	***CO_2***

17. What is the ionic equation for the neutralisation reaction?

	Acid	**+**	**Alkali**	**→**	**water**
Ionic equation	***H^+ (aq)***	+	***OH^- (aq)***	→	***H_2O(l)***

18. Describe the chemical test for hydrogen gas.

Collect the hydrogen gas in a test tube
Put a lighted splint into the testtube.
If the gas is hydrogen, it will burn with a squeaky pop.

19. Describe the chemical test for carbon dioxide.

Bubble the gas through limewater
If the gas is carbon dioxide it will turn the limewater cloudy

20. What does the word excess mean in a chemical reaction?

Excess means that there is more of a chemical than you need to make the chemical reaction happen.

21. Describe the steps to produce a soluble salt from an acid solution and an insoluble base.

Add excess base to the acid solution
Filter to remove the excess, unreacted base
Leave the solution to evaporate to so the salt forms

22. Describe the steps to produce a soluble salt from an acid solution and an alkali solution (soluble).

A titration must be used
Acid is placed in a burette and a pipette is used to accurately measure the correct amount of base
Indicator is added to the base
The acid is added to the base until the base is neutralised. The indicator tells us when this has happened
Record the volume of acid added.
Repeat the experiment adding the same amount of acid but this time do not add indicator.
Leave the neutralised solution to evaporate to form the salt.

23. Complete the table below to show the solubility of different substances.

Substance	**Soluble**	**Insoluble**
All nitrates	***All nitrates***	
Chlorides	***Most chlorides***	***Silver and lead chloride***
Sulfates	***Most sulfates***	***Lead, barium and calcium sulfate***
Carbonates	***Sodium, potassium and ammonium carbonates***	***Most carbonates***
Hydroxides	***Sodium, potassium and ammonium hydroxides***	***Most hydroxides***

24. Explain what is meant by a precipitation reaction.

A precipitation reaction involves reacting two solutions together to form an insoluble product

25. Complete the word and symbol equations below. Include state symbols.

Silver nitrate (aq)	+	Sodium chloride (aq)	→	***Silver Chloride (s)***	+	***Sodium nitrate (aq)***
$AgNO_3$ (aq)	+	***NaCl (aq)***	→	***AgCl (s)***	+	***$NaNO_3$ (aq)***
Iron sulfate (aq)	+	Sodium hydroxide (aq)	→	***Iron hydroxide (s)***	+	***Sodium sulfate (aq)***
$FeSO_4$ (aq)	+	***NaOH (aq)***	→	***FeOH (s)***	+	***$NaSO_4$ (aq)***
Copper sulfate (aq)	+	sodium hydroxide (aq)	→	***Copper hydroxide (s)***	+	***Sodium Sulfate (aq)***
$CuSO_4$ (aq)	+	***2NaOH (aq)***	→	***$Cu(OH)_2$ (s)***	+	***Na_2SO_4 (aq)***

26. Describe how a pure, dry sample of an insoluble salt can be made.

Carry out the precipitation reaction
Filter to remove excess, unreacted solid
Wash solid using distilled water
Leave the solid to dry

Paper 3 – Topic 3 electrolytic processes

1. What is an electrolyte?

____

2. What is meant by electrolysis?

____

3. What type of electrical energy is used to carry out electrolysis?

____

4. Electrolysis causes compounds to decompose. What is meant by decompose?

____

5. Explain why solid ionic compounds do not conduct electricity.

____

____

6. Explain why dissolved or molten ionic compounds will conduct electricity.

____

____

7. What is meant by an anion and a cation?

____

____

8. What happens to anions during electrolysis?

____

____

9. What happens to cations during electrolysis?

____

____

10. Give the name of a material used as an inert electrode during electrolysis.

____

11. Why should inert electrodes be used during electrolysis?

____

____

12. Complete the table below to show the products formed during the electrolysis of the compound shown.

Electrolyte		Cathode	Anode
Copper chloride solution, $CuCl_2$ (aq)	Half equation		
	Type of reaction		
	Product formed		
Sodium chloride solution, $NaCl$ (aq)	Half equation		
	Type of reaction		
	Product formed		
Sodium sulfate solution, Na_2SO_4 (aq)	Half equation		
	Type of reaction		
	Product formed		
Water acidified with sulfuric acid	Half equation		
	Type of reaction		
	Product formed		
Molten lead bromide, $PbBr_2$ (l)	Half equation		
	Type of reaction		
	Product formed		

13. What is meant by oxidation?

____

14. What is meant by reduction?

____

15. Complete the table below for the electrolysis of copper sulfate solution using copper electrodes.

Electrolyte		Cathode	Anode
Copper sufate solution, $CuSO_4$ (aq)	Half equation		
	Type of reaction		
	Product formed		

16. Explain why happens to the mass of the cathode and anode during the electrolysis of copper sulfate using copper electrodes.
 Cathode ____
 ____
 Anode ____
 ____

Paper 3 – Topic 3 electrolytic processes answers

1. What is an electrolyte?
An ionic compound that is either dissolved in solution of in a molten state.

2. What is meant by electrolysis?
Using a direct current (D.C.) supply to break down (decompose) an electrolyte into simpler substances.

3. What type of electrical energy is used to carry out electrolysis?
Direct current

4. Electrolysis causes compounds to decompose. What is meant by decompose?
Break down the substances into simpler substances.

5. Explain why solid ionic compounds do not conduct electricity.
The ions in a solid or held in place in the lattice by strong electrostatic forces which means that they cannot move.

6. Explain why dissolved or molten ionic compounds will conduct electricity.
When the ionic compound is dissolved or molten the ions are free to move because the electrostatic forces have been broken

7. What is meant by an anion and a cation?
Anion – a negatively charged ion
Cation – a positively charged ion

8. What happens to anions during electrolysis?
Anions migrate to the anode (positive electrode)

9. What happens to cations during electrolysis?
Cations migrate to the cathode (negative electrode)

10. Give the name of a material used as an inert electrode during electrolysis.
carbon

11. Why should inert electrodes be used during electrolysis?
So that the electrodes do not react with the electrolyte during electrolysis

12. Complete the table below to show the products formed during the electrolysis of the compound shown.

Electrolyte		Cathode	Anode
Copper chloride solution, $CuCl_2$ (aq)	Half equation	***Cu^{2+} (aq) + $2e^-$ $\rightarrow$ Cu (s)***	***$2Cl^-$ (aq) $\rightarrow$ Cl_2 (g) + $2e^-$***
	Type of reaction	***reduction***	***Oxidation***
	Product formed	***Copper***	***Chlorine***
Sodium chloride solution, NaCl (aq)	Half equation	***$2H^+$ (aq) + $2e^-$ $\rightarrow$ H_2 (g)***	***$2Cl^-$ (aq) $\rightarrow$ Cl_2 (g) + $2e^-$***
	Type of reaction	***reduction***	***Oxidation***
	Product formed	***Hydrogen***	***Chlorine***
Sodium sulfate solution, Na_2SO_4 (aq)	Half equation	***$2H^+$ (aq) + $2e^-$ $\rightarrow$ H_2 (g)***	***$4OH^-$ (aq) $\rightarrow$ $2H_2O$ (l) + O_2 (g) + $4e^-$***
	Type of reaction	***reduction***	***Oxidation***
	Product formed	***Hydrogen***	***Oxygen***
Water acidified with sulfuric acid	Half equation	***$2H^+$ (aq) + $2e^-$ $\rightarrow$ H_2 (g***	***$2O^{2-}$ (aq) $\rightarrow$ O_2 (g) + $4e^-$***
	Type of reaction	***reduction***	***Oxidation***
	Product formed	***Hydrogen***	***Oxygen***
Molten lead bromide, $PbBr_2$ (l)	Half equation	***Pb^{2+} (l) + $2e^-$ $\rightarrow$ Pb (s)***	***$2Br^-$ (l) $\rightarrow$ Br_2 (g) + $2e^-$***
	Type of reaction	***reduction***	***Oxidation***
	Product formed	***Lead***	***Bromine***

13. What is meant by oxidation?

Oxidation is loss of electrons

14. What is meant by reduction?

Reduction is gaining electrons

15.Complete the table below for the electrolysis of copper sulfate solution using copper electrodes.

Electrolyte		Cathode (pure cooper)	Anode (impure copper)
Copper sufate solution, $CuSO_4$ (aq)	Half equation	***Cu^{2+} (aq) + $2e^-$ $\rightarrow$ Cu (s)***	***Cu (s) $\rightarrow$ Cu^{2+} (aq) + $2e^-$***
	Type of reaction	***reduction***	**Oxidation**
	Effect on electrode	***Gains mass***	**Loses mass**

16. Explain why happens to the mass of the cathode and anode during the electrolysis of copper sulfate using copper electrodes.
   Cathode ***mass increases as new copper is formed***
   Anode ***loses mass***

Paper 3 – Topic 3 obtaining and using metals

1. What is formed when a metal reacts with water?

____

2. Write a balanced equation for the reaction of sodium with water.

Sodium	+	Water	→		+	
	+		→		+	

3. What is formed when a metal reacts with an acid?

____

4. Write a word and balanced symbol equation for the reaction between magnesium and hydrochloric acid.

Magnesium	+	Hydrochloric acid	→		+	
	+		→		+	

5. Which metal is the most reactive metal in the reactivity series?

____

6. Why is hydrogen included in the reactivity series of metals?

____

7. What is meant by a displacement reaction?

____

____

8. What is an ore?

____

____

9. Why is gold found in the Earth's crust?

____

____

10. What happens during oxidation?

____

11. What happens during reduction?

____

12. Explain how iron is extracted from its ore.

____

____

____

13. Explain how aluminium is extracted from its ore.

14. Why is electrolysis used to extract aluminium but not iron.

15. Explain what is meant by phytoextraction.

16. Explain what is meant by bioleaching.

17. What is meant by corrosion?

18. Which metals in the reactivity series are less likely to undergo corrosion? Explain your answer.

19. Give some advantages of recycling metals.

20. Explain what is meant by a life time assessment for a product.

Paper 3 – Topic 3 obtaining and using metals answers

1. What is formed when a metal reacts with water?
   hydrogen gas and a metal hydroxide

2. Write a balanced equation for the reaction of sodium with water.

Sodium	+	***Water***	→	***Sodium hydroxide***	+	***Hydrogen***
Na	+	***H_2O***	→	***NaOH***	+	***H_2***

3. What is formed when a metal reacts with an acid?
   a metal salt and hydrogen gas

4. Write a word and balanced symbol equation for the reaction between magnesium and hydrochloric acid.

Magnesium	+	Hydrochloric acid	→	***Magnesium Chloride***	+	***Hydrogen***
Mg	+	***HCl***	→	***$MgCl_2$***	+	***H_2***

5. Which metal is the most reactive metal in the reactivity series?
   Potassium

6. Why is hydrogen included in the reactivity series of metals?
   as a reference point. Metals below hydrogen do not react with water or acids

7. What is meant by a displacement reaction?
   a reaction in which a more reactive element displaces a less reactive element from its compound.

8. What is an ore?
   Ores are rocks that are obtained from the Earth's crust that contain metals

9. Why is gold found in the Earth's crust?
   Gold is unreactive so it does not combine with oxygen

10. What happens during oxidation?
    An element gains oxygen

11. What happens during reduction?
    Oxygen is removed from a compound

12. Explain how iron is extracted from its ore.
    Iron ore is haematite and it contains iron oxide
    Iron oxide is heated in a blast furnace with carbon
    Carbon is more reactive than iron
    The carbon displaces the iron from the iron oxide forming iron
    The iron has been reduced by the carbon

13. Explain how aluminium is extracted from its ore.
Aluminium ore is bauxite and it contains aluminium oxide
Electrolysis is used to extract the aluminium from the ore
The aluminium oxide is reduced to form aluminium metal

14. Why is electrolysis used to extract aluminium but not iron.
Aluminium is more reactive than iron
Aluminium would not be reduced by carbon

15. Explain what is meant by phytoextraction.
Plants are grown that absorb metal compounds
The plants are then burnt to form ash
The metal can be extracted from the ash

16. Explain what is meant by bioleaching.
Bacteria is grown on metal ore
Bacteria produce a solution that contains the metal ions
The metal can be produced from the metal ions

17. What is meant by corrosion?
Reactive metals will react with oxygen
This is an oxidation reaction

18. Which metals in the reactivity series are less likely to undergo corrosion? Explain your answer.
The metals at the bottom of the reactivity series because they are less reactive

19. Give some advantages of recycling metals.
Natural reserves in Earth's crust will last longer
Mining damages the environment e.g. noise and dust pollution
Recycling uses less energy than extracting new metal from its ore
Less waste in landfill

20. Explain what is meant by a life time assessment for a product.
This considers the effect on the environment of obtaining the raw materials, manufacturing the product, using the product and disposing of the product when it is no longer useful

Paper 3 – Topic 4 reversible reactions and equilibria

1. What is meant by a reversible reaction?

____

____

2. What is the symbol used to show that a reaction is reversible?

____

3. What is meant by a dynamic equilibrium?

____

____

4. Complete the word and balanced symbol equation for the Haber process.

	+		→	Ammonia
	+		→	

5. Where is the nitrogen obtained from?

____

6. Where is the hydrogen obtained from?

____

7. What three conditions are used during the haber process?

____

____

____

8. Does the Haber process favour high or low pressure? Explain your answer.

____

____

9. Does the Haber process favour high or low temperatures? Explain your answer.

____

____

10. Why is a catalyst added to the reaction?

____

____

11. Why are compromise conditions used in the process?

____

____

Paper 3 – topic 4 reversible reactions and equilibria answers

1. What is meant by a reversible reaction?

A reaction which can work in both the forward and backward direction

2. What is the symbol used to show that a reaction is reversible?

$\rightleftharpoons$

3. What is meant by a dynamic equilibrium?

both the forward and backward reactions take place at the same rate
the concentration of reactants and products remains constant

4. Complete the word and balanced symbol equation for the Haber process.

Hydrogen	+	***Nitrogen***	$\rightarrow$	Ammonia
$3H_2$	+	***N_2***	$\rightarrow$	**$2NH_3$**

5. Where is the nitrogen obtained from?

from air

6. Where is the hydrogen obtained from?

from natural gas

7. What three conditions are used during the haber process?

temperature 450oC
pressure 200 atmosphere
Iron catalyst

8. Does the Haber process favour high or low pressure? Explain your answer.

There are 4 moles of reactant gases and 2 moles of product gases in the equation
The reaction favours an increase in pressure
When the pressure is increased the equilibrium position moves to the right to try and reduce the pressure (there are less moles of gas)
this means more ammonia is formed

9. Does the Haber process favour high or low temperatures? Explain your answer.

The forward direction is exothermic which means it produces heat
The reaction favours lower temperatures
When the temperature is lowered the equilibrium position moves to the right to try and increase the temperature
this means more ammonia is formed

10. Why is a catalyst added to the reaction?

The catalyst sets the equilibrium up much more quickly

11. Why are compromise conditions used in the process?

High pressure – high pressures are expensive to generate. So a compromise is made to balance pressure and cost. The compromise is 200 atmospheres.

Low temperature – a low temperature will mean the reaction will have a slow rate of reaction. A compromise is made so that the reaction is not too slow. The compromise is 450^{o}C.

Paper 3 Topic 5 Triple chemistry only

1. Transition metals have the following properties:

i) ___________________________________

ii) ___________________________________

iii) __________________________________

iv) __________________________________

2. What is the oxidation of metals called?

____

3. Which two conditions are needed for iron to rust?

____

4. State 3 ways in which the rusting of iron can be prevented.

i) ___________________________________

ii) ___________________________________

iii) ___________________________________

5. Name the process which is used to improve the appearance and/or the resistance to corrosion of metals.

____

6. Describe the process of electroplating.

i) ____

ii) ____

iii) ____

iv) ____

7. Draw a labelled diagram to show the particle arrangement in an alloy.

8. Explain why alloys are often stronger than the pure metal.

____

____

____

9. What is the advantage of making the following iron alloys?

Alloy	Advantage
Stainless steel – iron with chromium/nickel	
Steel – iron with carbon	

10. Give the property and use for each material list below.

Material	Property	Uses
Aluminium		
Copper		
Gold		
Magnalium (5% magnesium)		
Brass		

11. What is the equation used to calculate concentration?

____

____

12. What are the units for concentration?

____

13. Calculate the concentration of 45g of sodium chloride in 350cm^3 of water in gdm^{-3}.

____

____

____

14. Calculate the concentration of 25g of NaOH in 475cm^3 in moldm^{-3}.

____

____

____

15. Convert the concentration of 0.05moldm^{-3} sodium carbonate, Na_2CO_3, in to gdm^{-3}.

____

____

____

16. What is the concentration in $mol dm^{-3}$ of some potassium hydroxide, KOH, solution with a concentration of 11.2 $g dm^{-3}$ 0.200$mol dm^{-3}$?

____

____

____

17. Describe the steps to carry out a titration.

i) ____

ii) ____

iii) ____

iv) ____

v) ____

vi) ____

vii) ____

18. What is the ionic equation that takes place during the neutralisation reaction in a titration?

____

19. What is the equation used to calculate the number of moles in a solution?

____

____

20. What is the equation used to calculate the percentage yield for a reaction?

____

____

21. What is the actual yield?

____

22. What is the theoretical yield?

____

23. Why is the actual yield always less than the theoretical yield?

i) ____________________________________

ii) ____________________________________

iii) ____________________________________

24. What is the equation used to calculate the atom economy for a reaction?

____

____

25. What is the definition for atom economy?

____

____

26. What factors have to be considered when choosing a reaction pathway?

i) ________________________

ii) _______________________

iii) ______________________

iv) ______________________

v) _______________________

27. What is the molar volume of a gas?

____

____

28. What is the Haber process?

____

____

29. Why are compromise conditions used for the Haber process?

Temperature 450^{o}C – ____

____

Pressure 200 atms – ____

____

30. Complete the table below to show how changing conditions effects the reversible reaction.

Type of reaction	Condition	How does it effect the equilibrium position?	Effect on product yield
Endothermic in the forward direction	Lowering the temperature		
	Increasing the temperature		
Exothermic in the forward direction	Lowering the temperature		
	Increasing the temperature		
More moles on the right	Lowering the pressure		

	Increasing the pressure		
More moles on the left	Lowering the pressure		
	Increasing the pressure		

31. Complete the table below to show how the rate of attainment of equilibrium is affected by each condition.

Condition	**Effect on attainment of equilibrium**
Increasing temperature	
Decreasing temperature	
Increasing pressure	
Decreasing pressure	
Increasing concentration	
Decreasing concentration	
Addition of a catalyst	

32. Which chemicals compounds do fertilisers commonly contain?

____

33. Why do fertilisers contain potassium, nitrogen and phosphorus compounds?

____

34. Complete the equation for the reaction below:

Word equation	Ammonia	+	Nitric acid	→	
Balanced symbol equation		+		→	

35. What are the differences between the laboratory preparation and industrial preparation of ammonium sulfate?

Laboratory	____
Industrial	____

36. What does a fuel cell do?

____
____

37. How long will a fuel cell produce a voltage for?

____
____

38. What are the strengths and weaknesses of fuel cells?

Strengths	Weaknesses

Paper 3 Topic 5 Triple chemistry only answers

1. Transition metals have the following properties:

i) high melting points

ii) high density

iii) they form coloured compounds

iv) act as catalysts e.g. iron in the Haber process

2. What is the oxidation of metals called?

Corrosion of metals.

3. Which two conditions are needed for iron to rust?

The presence of oxygen and water.

4. State 3 ways in which the rusting of iron can be prevented.

i) application of oil to the surface of the metal

ii) application of paint to the surface of the metal

iii) sacrificial protection – using a more reactive metal

5. Name the process which is used to improve the appearance and/or the resistance to corrosion of metals.

Electroplating

6. Describe the process of electroplating.

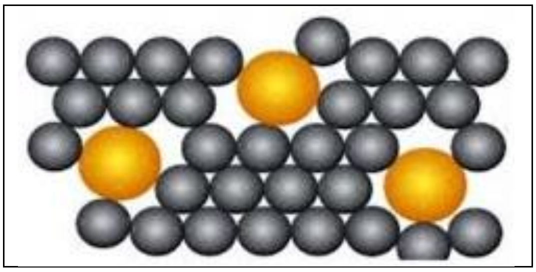
i) place the object to be electroplated at the cathode

ii) the electrolyte should contain the metal ions which will cover the object

iii) Turn the circuit on. The cations migrate to the cathode and are reduced.

iv) The object is covered in a thin layer of the new metal.

7. Draw a labelled diagram to show the particle arrangement in an alloy.



8. Explain why alloys are often stronger than the pure metal.

The different sized ions disrupt the layers of the metal.

This means that the layers can no longer slide.

This makes the alloy stronger.

9. What is the advantage of making the following iron alloys?

Alloy	Advantage
Stainless steel – iron with chromium/nickel	***Stainless steel is resistant to corrosion***
Steel – iron with carbon	***The steel is stronger than the original iron***

10. Give the property and use for each material list below.

Material	Property	Uses
Aluminium	***Low density, resistant to corrosion***	***Cooking foil***
Copper	***Good conductor of heat and electricity, does not react with water***	***Electrical wires, water pipes***
Gold	***Does not corrode, malleable***	***Jewellery***
Magnalium (5% magnesium)	***Lighter and more workable than pure aluminium***	***Aircraft parts, scientific equipment***
Brass	***Good conductor, stronger than pure copper***	***Musical instruments, plumbing and electrical components***

11. What is the equation used to calculate concentration?

$$\textbf{Concentration} = \frac{\textbf{Amount (g or mol)}}{\textbf{Volume (dm}^3\textbf{)}}$$

12. What are the units for concentration?

gdm^{-3} or $moldm^{-3}$

13. Calculate the concentration of 45g of sodium chloride in 350cm^3 of water in gdm^{-3}.

$$\textbf{Concentration} = \frac{\textbf{amount}}{\textbf{Volume}} = \frac{\textbf{45}}{\textbf{(350/1000)}} = \textbf{129gdm}^{-3}$$

Do not forget to divide the volume by 1000cm^3 to convert the volume to dm^3

14. Calculate the concentration of 25g of NaOH in 475cm^3 in moldm^{-3}.

$$\textbf{No of moles of NaOH} = \frac{\textbf{mass}}{\textbf{mass of 1 mole}} = \frac{\textbf{25}}{\textbf{(23+16+1)}} = \textbf{0.625 moles}$$

$$\textbf{Concentration} = \frac{\textbf{amount}}{\textbf{Volume}} = \frac{\textbf{0.625}}{\textbf{(475/1000)}} = \textbf{1.32 moldm}^{-3}$$

Do not forget to divide the volume by 1000cm^3 to convert the volume to dm^3

15. Convert the concentration of $0.05 moldm^{-3}$ sodium carbonate, Na_2CO_3, in to gdm^{-3}.

Mass = no of moles of Na_2CO_3 x mass of one mole = 0.05 x ((23x2)+12+(16x3)) = 5.3 gdm^{-3}

16. What is the concentration in $moldm^{-3}$ of some potassium hydroxide, KOH, solution with a concentration of 11.2 gdm^{-3} 0.200$moldm^{-3}$?

No of moles of KOH = $\dfrac{\textbf{mass}}{\textbf{Mass of 1 mole}} = \dfrac{\textbf{11.2}}{\textbf{(39+16+1)}} = \textbf{0.2 moldm}^{-3}$

17. Describe the steps to carry out a titration.

i) Fill the burette with hydrochloric acid.

ii) Use a pipette to measure 25cm3 of sodium hydroxide and add this to a conical flask.

iii) Add a few drops of phenolphthalein indicator to the sodium hydroxide.

iv) Place the conical flask on a white tile under the burette.

v) Add the hydrochloric acid to the conical flask whilst swirling the contents of the flask.

vi) Stop adding the acid once the indicator has change colour from pink to clear.

vii) Repeat the process until you achieve 2 further results that are concordant.

18. What is the ionic equation that takes place during the neutralisation reaction in a titration?

$$H^+ + OH^- \rightarrow H_20$$

19. What is the equation used to calculate the number of moles in a solution?

$$\textbf{Number of moles} = \frac{\textbf{volume (cm}^3\textbf{)}}{\textbf{1000}} \textbf{ x concentration (moldm}^{-3}\textbf{)}$$

20. What is the equation used to calculate the percentage yield for a reaction?

$$\textbf{Percentage yield} = \frac{\textbf{actual yield}}{\textbf{Theoretical yield}} \textbf{ x 100\%}$$

21. What is the actual yield?

The amount of product produced during a chemical reaction.

22. What is the theoretical yield?

The amount of product produced based on the balanced chemical equation.

23. Why is the actual yield always less than the theoretical yield?

i) Incomplete reaction

ii) Loss of chemicals during the reaction

iii) unwanted side reactions

24. What is the equation used to calculate the atom economy for a reaction?

$$\textit{Atom economy} = \frac{\textit{mass of desired product}}{\textit{Total mass of products}} \times 100$$

25. What is the definition for atom economy?

The atom economy measures how many reactant atoms form a desired product.

26. What factors have to be considered when choosing a reaction pathway?

i) atom economy
ii) yield
iii) rate
iv) equilibrium position
v) usefulness of by products

27. What is the molar volume of a gas?

The volume occupied by one mole of molecules of any gas at room temperature and pressure.

28. What is the Haber process?

A reversible reaction between nitrogen and hydrogen to form ammonia.

29. Why are compromise conditions used for the Haber process?

Temperature 450^{o}C – ***reaction would favour low temperature as it is exothermic in the forward direction. A low temperature would cause the rate of reaction to be too slow.***
Pressure 200 atms – ***reaction favours high pressure as there are less moles on the right. A high pressure would be expensive as pipes would need to be reinforced to withstand the pressure.***

30. Complete the table below to show how changing conditions effects the reversible reaction.

Type of reaction	Condition	How does it effect the equilibrium position?	Effect on product yield
Endothermic in the forward direction	Lowering the temperature	***Equilibrium shifts to the left to increase the temperature***	***Product decreases***
	Increasing the temperature	***Equilibrium shifts to the right to decrease the temperature***	***Product increases***
Exothermic in the forward direction	Lowering the temperature	***Equilibrium shifts to the right to increase the temperature***	***Product increases***
	Increasing the temperature	***Equilibrium shifts to the left to decrease the temperature***	***Product decreases***

More moles on the right	Lowering the pressure	***Equilibrium moves to the right to increase the pressure***	***Product increases***
	Increasing the pressure	***Equilibrium moves to the left to decrease the pressure***	***Product decreases***
More moles on the left	Lowering the pressure	***Equilibrium moves to the left to increase the pressure.***	***Product decreases***
	Increasing the pressure	***Equilibrium moves to the right to decrease the pressure***	***Product increases***

31. Complete the table below to show how the rate of attainment of equilibrium is affected by each condition.

Condition	**Effect on attainment of equilibrium**
Increasing temperature	***Quicker attainment of the equilibrium***
Decreasing temperature	***Slower attainment of the equilibrium***
Increasing pressure	***Quicker attainment of the equilibrium***
Decreasing pressure	***Slower attainment of the equilibrium***
Increasing concentration	***Quicker attainment of the equilibrium***
Decreasing concentration	***Slower attainment of the equilibrium***
Addition of a catalyst	***Quicker attainment of the equilibrium***

32. Which chemicals compounds do fertilisers commonly contain?

Nitrogen, phosphorus and potassium compounds.

33. Why do fertilisers contain potassium, nitrogen and phosphorus compounds?

They promote plant growth.

34. Complete the equation for the reaction below:

Word equation	Ammonia	+	Nitric acid	→	***Ammonium nitrat***e
Balanced symbol equation	NH_3	**+**	HNO_3	**→**	NH_4NO_3

35. What are the differences between the laboratory preparation and industrial preparation of ammonium sulfate?

Laboratory	***Ammonia is reacted directly with sulfuric acid. This is a small scale batch process. Slow and only small amounts produced.***
Industrial	***The process starts with the individual elements and involves many stages to produce the ammonia and sulfuric acid. Once made the ammonia and sulfuric acid are reacted together to make ammonium sulfate. This is a large scale continuous process which is quicker and produces larger quantities.***

36. What does a fuel cell do?

A fuel cell reacts hydrogen and oxygen together to produce a voltage and water is the only product.

37. How long will a fuel cell produce a voltage for?

Until one of its reactants is used up.

38. What are the strengths and weaknesses of fuel cells?

Strengths	Weaknesses
Do not produce greenhouse gases e.g. carbon dioxide	***Cannot be used for large scale energy production***
Do not produce harmful products e.g. carbon monoxide, sulfur dioxide, nitrogen oxides, soot (carbon).	***Hydrogen needs much larger storage volume compared to fossil fuels as it is a gas.***
More efficient than fossil fuel power stations or batteries.	***Safe storage of hydrogen under high pressure (reduces volume needed) is a problem.***
Renewable (as long as hydrogen is available from the water produced)	***Electrolysis is need to produce hydrogen from water. This is expensive.***
Do not have many moving parts to go wrong.	***Disposal of used fuel cells is a problem. They may contain toxic metal compounds.***

Paper 4 – Topic 6 groups in the periodic table

1. The elements in group 1 of the periodic table are called __________________. The elements in group 7 are called ____________ and the elements in group 0 are called the ________________.

2. Draw a labelled diagram for the structure of particles in metal.

3. Use the structure of metals to explain the properties shown in the table.

Property	Explanation
Will conduct electricity	
Have high melting points	

4. Group 1 metals can be cut with a knife. This means that they are _____________. The heat released as they react causes group 1 metals to form a molten ball because they have relatively __________ melting points.

5. Describe what you see when a group 1 metal reactions with water.

____

____

6. What is the pattern in reactivity for the group 1 metals and explain this pattern?

____

____

____

7. Complete the table below for the group 7 elements.

Element	Colour	State at room temperature
Fluorine		
Chlorine		
Bromine		
Iodine		

8. What is the pattern in melting and boiling point of the group 7 halogens?

____

____

9. Describe the chemical test for chlorine.

____

____

10. Complete the word and symbol equations below for the reaction between a metal and a halogen.

Aluminium	+	Chlorine	→	
Iron	+	Bromine	→	

11. Complete the word and symbol equations below for the reaction between a metal and a halogen.

Hydrogen	+	Chlorine	→	

12. What is formed when the product of the reaction shown in question 10 dissolves in water?

____

13. What is the pattern in reactivity for the elements in group 7 and explain it?

____

____

____

14. Complete the equations below to show the displacement reactions which take place between the halogens.

1	Chlorine	+	Potassium Iodide	→		+	
2	Iodine	+	Potassium Chloride	→			

15. What is meant by reduction?

____

____

16. What is meant by oxidation?

____

____

17. For reaction 1 in question 13 identify which substances have been oxidised and reduced.
Equation 1 - Oxidised ____________________ reduced ______________________

18. Explain why the noble gases are chemically inert.

____

____

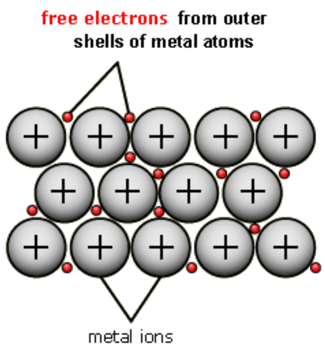
19. Complete the table below to show the properties and uses for noble gases.

Property	Use

Paper 4 – Topic 6 groups in the periodic table answers

1. The elements in group 1 of the periodic table are called ***alkali metals***. The elements in group 7 are called ***halogens*** and the elements in group 0 are called the ***noble gases.***

2. Draw a labelled diagram for the structure of particles in a metal.



3. Use the structure of metals to explain the properties shown in the table.

Property	Explanation
Will conduct electricity	***The delocalised electrons are able to move about to carry the electrical current***
Have high melting points	***There are strong electrostatic attractions in the structure***

4. Group 1 metals can be cut with a knife. This means that they are ***soft***.. The heat released as they react causes group 1 metals to form a molten ball because they have relatively ***low*** melting points.

5. Describe what you see when a group 1 metal reactions with water.
they float on the surface, fizz as they produce a gas (hydrogen) and form a molten ball.
If indicator is added the water will turn purple as the alkali is made (a metal hydroxide)

6. What is the pattern in reactivity for the group 1 metals and explain this pattern?
as you go down the group the metals become more reactive
all group 1 metals need to lose an electron to become stable
the atoms become bigger as you go down the group meaning that the electrons are further away from the positive nucleus and are therefore easier to lose.

7. Complete the table below for the group 7 elements.

Element	Colour	State at room temperature
Fluorine	***Yellow***	***Gas***
Chlorine	***Green***	***Gas***
Bromine	***Orange***	***Liquid***
Iodine	***Dark grey***	***Solid***

8. What is the pattern in melting and boiling point of the group 7 halogens?
As you go down the group the melting and boiling point increases.

9. Describe the chemical test for chlorine.
Let the gas pass over a piece of damp litmus. If the gas is chlorine the litmus will bleach (turn white).

10. Complete the word and symbol equations below for the reaction between a metal and a halogen.

Aluminium	+	Chlorine	→	***Aluminium chloride***
2Al	***+***	***$3Cl_2$***	→	***$2AlCl_3$***
Iron	+	Bromine	→	***Iron Bromide***
2Fe	***+***	***$3Br_2$***		***$2FeBr_3$***

11. Complete the word and symbol equations below for the reaction between a metal and a halogen.

Hydrogen	+	Chlorine	→	***Hydrogen chloride***
H_2		***Cl_2***	→	***2HCl***

12. What is formed when the product of the reaction shown in question 9 dissolves in water?
an acidic solution

13. What is the pattern in reactivity for the elements in group 7 and explain it?
as you go up the group they become more reactive
they all need to gain one electron to become stable
the atoms become smaller as you go up the group making it easier for them to gain an electron (the electron would be added nearer to the positive nucleus)

14. Complete the equations below to show the displacement reactions which take place between the halogens.

Chlorine	+	Potassium Iodide	→	***Potassium Chloride***	***+***	***Iodine***
Cl_2	+	***2KI***	→	***2KCl***	+	***I_2***
Iodine	+	Potassium Chloride	→	***No reaction***		

15. What is meant by reduction?
Reduction is the gaining of electrons

16. What is meant by oxidation?
Oxidation is the loss of electrons

17. For reaction 1 in question 13 identify which substances have been oxidised and reduced.
Equation 1 - Oxidised ***Iodide ion to iodine*** reduced ***chlorine to chloride ion***

18. Explain why the noble gases are chemically inert.

they all have full outer energy which means that they are stable

they do not need to react to gain or lose electrons

19. Complete the table below to show the properties and uses for noble gases.

Property	Use
Inert	***Provide an atmosphere for welding and filament lamps to stop the hot metals reacting***
Low density	***Filling a balloons so that they float***
Non- flammable	***As above***

Paper 4 – Topic 7 rates of reaction

1. When particles collide they need to have enough ________________ and the correct _______________ for the collision to be _________________. If the particles do not have enough energy then they will not react.

2. Explain how increasing the following causes the rate of reaction to increase

a) Increasing concentration or pressure in gases

____

____

____

b) increasing temperature

____

____

____

c) Increasing surface area

____

____

____

d) Adding a catalyst

____

____

____

3. What is a catalyst?

____

____

____

4. What is an enzyme?

____

____

5. Where are enzymes used?

____

6. What is meant by an exothermic reaction?

____

7. What is meant by an endothermic reaction?

____

8. When bonds are broken energy is _________________. This means that it is an ____________________ process.

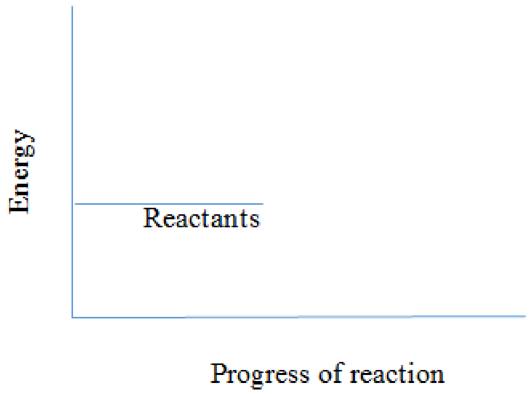
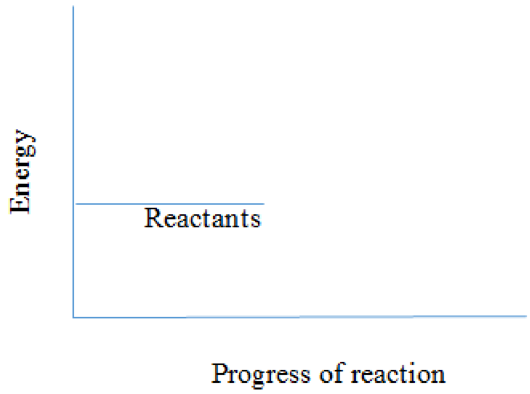
9. When bonds are made energy is __________________. This means that it is an ____________________ process.

10. Using the information in questions 8 and 9 to explain why a reaction would be
a) endothermic overall.

b) exothermic overall.

12. What is meant by activation energy?

13. Complete the two diagrams to show the energy changes for endothermic and exothermic reactions. Clearly label the activation energy in each diagram.

Exothermic reaction	Endothermic reaction
	

14. Calculate the energy change for the combustion of methane. Complete the word and balanced symbol equations, including a diagram of the molecules involved.

Word equation	Methane	+	Oxygen	→	________ ________	+	________ ________
Symbol equation		+		→		+	
Diagram of molecules		+		→		+	

Bond energy values

Type of bond	Amount of energy to break or make the bond (KJ mol^{-1})
C-H	413
O=O	495
C=O	745
O-H	467

Paper 4 – Topic 7 rates of reaction answers

1. When particles collide they need to have enough ***energy*** and the correct ***orientation*** for the collision to be ***successful***. If the particles do not have enough energy then they will not react.

2. Explain how increasing the following causes the rate of reaction to increase

a) Increasing concentration or pressure in gases

this means that there are more particles which increases the chance of successful collisions causing the rate of reaction to increase

b) increasing temperature

this gives the particles more energy which means that they move around more which increases the chance of successful collisions causing the rate of reaction to increase

c) Increasing surface area

this means that there is more space available for successful collisions to happen causing the rate of reaction to increase

d) Adding a catalyst

a catalysts lowers the amount of energy needed for particles to react. This means that there are more particles with enough energy to react which increases the chance of successful collisions causing the rate of reaction to increase.
Catalysts are not used up during the reaction.

3. What is a catalyst?

A catalyst is a substance that speeds up the rate of a reaction by lowering the activation energy
It is not chemical changed or used up during the reaction – its mass stays the same

4. What is an enzyme?

An enzyme is a biological catalyst

5. Where are enzymes used?

Enzymes are used in the production of alcoholic drinks

6. What is meant by an exothermic reaction?

a chemical reaction which releases energy to the surroundings causing them to become hotter

7. What is meant by an endothermic reaction?

a chemical reaction which takes in energy from the surroundings causing them to become cooler

8. When bonds are broken energy is ***taken in***. This means that it is an ***endothermic*** process.

9. When bonds are made energy is ***released***. This means that it is an ***exothermic*** process.

10. Using the information in questions 8and 9 to explain why a reaction would be

a) endothermic overall.

more energy is taken in to break bonds than is released as bonds are made

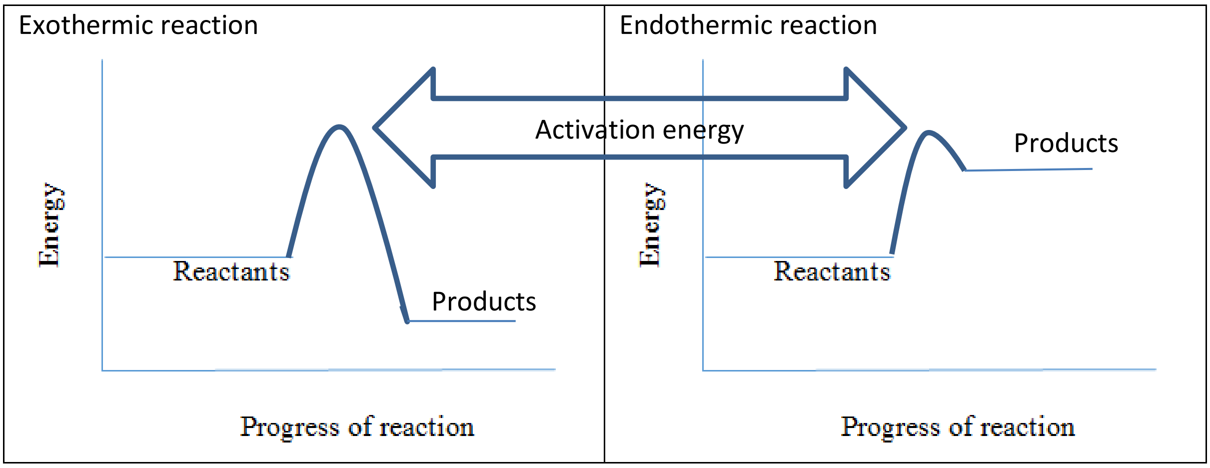
b) exothermic overall.

more energy is released as bonds are made than is taken in to break bonds

12. What is meant by activation energy?

Activation energy is the minimum amount of energy required for a successful reaction to happen when particles collide

13. Complete the two diagrams to show the energy changes for endothermic and exothermic reactions. Clearly label the activation energy in each diagram.



14. Calculate the energy change for the combustion of methane. Complete the word and balanced symbol equations, including a diagram of the molecules involved.

Word equation	Methane	+	Oxygen	→	***Carbon Dioxide***	+	***Water***
Symbol equation	***CH_4***	+	***$2O_2$***	→	***CO_2***	+	***$2H_2O$***
Diagram of molecules	H–C–H with H above and below C	+	O=O O=O	→	O=C=O	+	H–O–H H–O–H

Bond energy values

Type of bond	Amount of energy to break or make the bond (KJ mol^{-1})
C-H	413
O=O	495
C=O	745
O-H	467

Bonds broken				*Bonds made*		
C-H x 4	*4x 413 =*	*1652*		*C=O x 2*	*2 x 745 =*	*1490*
O=O x2	*2x 495 =*	*990*		*O-H x 4*	*4 x 467 =*	*1868*
	Total	***2642 KJmol^{-1}***			***Total***	***3358 KJmol^{-1}***
Total energy change		*=Bonds broken – bonds made*				
		= 2642 – 3358 = ***-716 KJmol^{-1}***				
This means that this reaction is exothermic. More energy is released as bonds are made than is taken in when bonds are broken.						

Paper 4 – topic 8 fuels and Earth science

1. What does hydrocarbon mean?

____

2. Crude oil is a mixture of ____________________.

3. Why is crude oil called a finite resource?

____

____

4. What is the name of the process which is used to separate crude oil into more useful products?

____

5. Explain how the process in number 4 is used to separate the fractions in crude oil.

____

____

____

____

____

6. Give a use for each of the following fractions

Fraction	Use
Gases	
Petrol	
kerosene	
Diesel oil	
Fuel oil	
bitumen	

7. Complete the following table by adding arrows to show the trend in properties

	Boiling point	Ease of ignition	viscosity
Short hydrocarbon chain Long hydrocarbon chain			

8. Explain what the term homologous series means.

____

____

9. Explain what an alkane is.

____

____

10. What is meant by complete combustion?

____

____

11. Complete the equation below for the complete combustion of methane

Methane	+	__________	→	__________ __________	+	__________ __________

12. How do you test for carbon dioxide and what is a positive result?

____

13. What two chemicals does the incomplete combustion of a hydrocarbon produce? Why does this happen?

____

____

14. Explain why one of the chemicals in question 13 is toxic.

____

____

15. Why is sulfur dioxide formed when some hydrocarbons are burnt?

____

16. How does sulfur dioxide form acid rain and what problems does acid rain cause?

____

____

17. Explain how oxides of nitrogen are formed.

____

____

18. What are the advantages and disadvantages of using hydrogen as a fuel in cars?

Advantages	Disadvantages

19. Write the word and balanced symbol equation for the combustion of hydrogen.

Hydrogen	+	Oxygen	→	
	+		→	

20. What is meant by a non-renewable fuel?

____

____

21. List three fuels formed from crude oil which are non-renewable.

a) ___________________ b) _____________________ c) ______________________

22. Which non-renewable fuel is formed from natural gas?

____

23. What is meant by the term cracking?

____

____

24. Why do we need to use cracking?

____

25. Decane $C_{10}H_{22}$ can be cracked to produce a molecule of propane, C_3H_8, and one molecule of an alkane X. The formula of alkane molecule X is:

A C_7H_{14}

B C_7H_{16}

C C_8H_{16}

D $C_{13}H_{28}$

26. The earths early atmosphere was produced by __________________________

27. Which gases did the Earth's early atmosphere contain?

Gas	Amount

28. What caused the oceans to form?

____

____

29. Explain how the amount of carbon dioxide in the atmosphere was reduced by the oceans?

____

30. What effect did primitive plants have on the atmosphere and what process did they use to do this?

____

____

31. Describe the chemical test for oxygen.

____

____

32. Which gases in the Earth's atmosphere keep the Earth warm?

Gases

33. Explain how the gases in question 32 keep the Earth warm.

____

____

34. What does the term greenhouse effect mean?

____

____

35. Which human activities do scientists believe are causing climate change?
1. ____________________________
2. ____________________________

36. List and explain activities that can be used to reduce the effects of climate change.

Activities	Explanation

Paper 4 – topic 8 fuels and Earth science answers

1. What does hydrocarbon mean?

a compound which contains carbon and hydrogen only

2. Crude oil is a mixture of ***hydrocarbons***

3. Why is crude oil called a finite resource?

Crude oil is formed from fossils in a process that takes millions of years. This means that it cannot be replaced and it will run out.

4. What is the name of the process which is used to separate crude oil into more useful products?

fractional distillation (1)

5. Explain how the process in number 4 is used to separate the fractions in crude oil.

The crude oil mixture is heated

The shorter hydrocarbon molecules have a lower boiling point and rise high up the column. The longer hydrocarbons have higher boiling points and condense towards the bottom of the column. The fractions condense at different points in the column allowing them to be separated.

6. Give a use for each of the following fractions

Fraction	Use
Gases	***Heating in houses/cooking***
Petrol	***Fuel for cars***
kerosene	***Fuel for aeroplanes***
Diesel oil	***Fuel for cars/trains***
Fuel oil	***Fuel for large ships and power stations***
bitumen	***Road surfaces/roofs***

7. Complete the following table by adding arrows to show the trend in properties

	Boiling point	Ease of ignition	viscosity
Short hydrocarbon chain long hydrocarbon chain	↑	↑	↓

8. Explain what the term homologous series means.

A family of molecules that have the same general formula, increase by a CH_2 group, have an increase in boiling point and similar chemical properties.

9. Explain what an alkane is.

Alkane are an homologous series that contain carbon carbon single bonds.

10. What is meant by complete combustion?
burning a hydrocarbon with enough oxygen to produce carbon dioxide and water, releasing energy

11. Complete the equation below for the complete combustion of methane

Methane	+	***oxygen***	→	***Carbon dioxide***	+	***water***

12. How do you test for carbon dioxide and what is a positive result?
Collect the gas and bubble it through limewater. If the gas turns cloudy then it is carbon dioxide.

13. What two chemicals does the incomplete combustion of a hydrocarbon produce? Why does this happen?
carbon (soot) and carbon monoxide are formed. This happens when the fuel is burnt with a limited supply of oxygen.

14. Explain why one of the chemicals in question 13 is toxic.
carbon monoxide will bond to red blood cells
stops them carrying oxygen
causes suffocation

15. Why is sulfur dioxide formed when some hydrocarbons are burnt?
The fuel contains sulfur as an impurity which reacts with oxygen to form sulfur dioxide

16. How does sulfur dioxide form acid rain and what problems does acid rain cause?
it dissolves in water in the clouds
can erode buiildings/statues

17. Explain how oxides of nitrogen are formed.
Fuels burn in engines. The high temperatures produced causes the nitrogen and oxygen to react together forming oxides of nitrogen

18. What are the advantages and disadvantages of using hydrogen as a fuel in cars?

Advantages	Disadvantages
Does not produce carbon dioxide when burnt	***It is a gas so it is difficult to store***
It is easy to burn	***Limited availability for refuelling***

19. Write the word and balanced symbol equation for the combustion of hydrogen.

Hydrogen	+	Oxygen	→	***Water***
$2H_2$	+	O_2	→	$2H_2O$

20. What is meant by a non-renewable fuel?
A fuel that once used cannot be replaced because it takes too long to replace the fuel e.g. fossil fuels.

21. List three fuels formed from crude oil which are non-renewable.
a) ***petrol*** b) ***kerosene*** c) ***diesel oil***

22. Which non-renewable fuel is formed from natural gas?
methane

23. What is meant by the term cracking?
splitting long chain hydrocarbons into shorted hydrocarbons

24. Why do we need to use cracking?
not enough short chain hydrocarbons available to meet the demands for fuels
need to split long chains to make more short chains

25. Decane $C_{10}H_{22}$ can be cracked to produce a molecule of propane, C_3H_8, and one molecule of an alkane X. The formula of alkane molecule X is:

A ***C_7H_{14}***

26. The earths early atmosphere was produced by ***volcanic activity***

27. Which gases did the Earth's early atmosphere contain?

Gas	Amount
Carbon dioxide	***Large amounts***
Water vapour	
Other gases e.g.methane, ammonia	***Small amounts***
No oxygen	

28. What caused the oceans to form?
The water vapour cooled and condensed to form oceans

29. Explain how the amount of carbon dioxide in the atmosphere was reduced by the oceans?
The carbon dioxide gas dissolved in the oceans

30. What effect did primitive plants have on the atmosphere and what process did they use to do this?
They absorbed carbon dioxide and produced oxygen using photosynthesis

31. Describe the chemical test for oxygen.
Place a glowing splint in to the gas and if it relights then the gas is oxygen

32. Which gases in the Earth's atmosphere keep the Earth warm?

Gases
Carbon dioxide
Methane
Water vapour

33. Explain how the gases in question 32 keep the Earth warm.

The gases absorb heat radiation and release it keeping the Earth warm

34. What does the term greenhouse effect mean?

This describes the process which causes the Earth to become warm as the atmosphere causes heat to be trapped.

35. Which human activities do scientists believe are causing climate change?

1. ***Burning fossil fuels***
2. ***Livestock farming***

36. List and explain activities that can be used to reduce the effects of climate change.

Activities	Explanation
Use renewable energy resources	***Reduced greenhouse gas emissions***
Flood defences	***Prevent damage as water levels rise***

Paper 4 – Topic 9 Triple chemistry only

1. Why must the qualitative test for an ion be unique?

____

2. Describe how to carry out a flame test.

i) ____

ii) ____

iii) ____

iv) ____

v) ____

vi) ____

3. Complete the table below to show the colour of the cations during a flame test.

Cation	**Flame test colour**
Lithium, Li^{+}	
Sodium, Na_{+}	
Potassium, K_{+}	
Calcium, Ca^{2+}	
Copper, Cu^{2+}	

4. What is a precipitate?

____

5. Complete the table below to show the observations for the tests using sodium hydroxide. Complete an equation for each reaction.

Cation in solution	**Description of the test**	**Observation**	**Chemical equation for the reaction**
Aluminium, Al^{3+}			
Calcium, Ca^{2+}			
Copper, Cu^{2+}			
Iron (II), Fe^{2+}			
Iron (III) Fe^{3+}			
Ammonium, NH_4^{+}			

6. Describe the chemical test for ammonia.

____

7. Describe the test for carbonate ions, CO_3^{2-}.

____

____

8. Describe the test for sulfate ions, SO_4^{2-}.

____

____

9. Give the ionic equation for the test for a sulfate ion.

____

10. Describe the test for halide ions.

____

____

11. Complete the table below to show the results and equations for halide ions.

Halide ion	Observation	Equations
Chloride, Cl^-		
Bromide, Br^-		
Iodide, I^-		

12. What are the advantages of instrumental qualitative tests?

i) ____

ii) ____

iii) ____

13. Describe two ways that a flame photometer is used to identify ions.

i) ____

ii) ____

14. Complete the table below to show the formula and molecular structure for each alkane.

Alkanes	Formula	Molecular structure
Methane		
Ethane		
Propane		
Butane		

15. Why are alkanes saturated hydrocarbons?

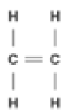
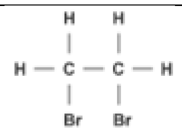
16. Complete the table below to show the formula and molecular structure for each alkene.

Alkenes	**Formula**	**Molecular structure**
Ethene		
Propene		
But-1-ene		
But-2-ene		

17. Why are alkenes unsaturated hydrocarbons?

18. Describe the test to identify alkanes and alkenes.

19. Complete the table below to show the reactions of bromine water with ethene.

Ethene	+	Bromine	→	1, 2 – dibromoethane
C_2H_4	+	Br_2	→	CH_3BrCH_3Br
	+	Br_2	→	

20. What type of reaction is combustion?

21. What are the products of the complete combustion (oxidation) of alkanes and alkenes?

22. Complete the reactions below to produce balanced equations for the combustion of alkanes and alkenes.

Methane	+	Oxygen	→		+	
	+		→		+	
Ethene	+	Oxygen	→		+	
	+		→		+	
Ethane	+	Oxygen	→		+	
	+		→		+	

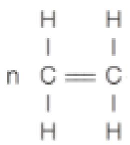
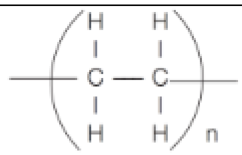
23. What is a polymer?

____

24. What is a monomer?

____

25. Complete the equation below to show the formation of poly(ethene) from ethene.

	
Ethene	Poly(ethene)

26. What is needed to make additional polymers?

____

27. Draw the monomers and polymers below.

Monomer	**Polymer structure**
Propene	Poly(propene)
Chloroethene	Poly(chloroethene) (PVC)
Tetrafluoroethene	Poly(tetrafluoroehene) (PTFE)

28. Complete the table below to show the properties and uses of each polymer.

Polymer	Properties	uses
Poly(ethene)		
Poly(propene)		
Poly(chloroethene) (PVC)		
Poly(tetrafluoroethene) (PTFE)		

29. What is condensation polymerisation?

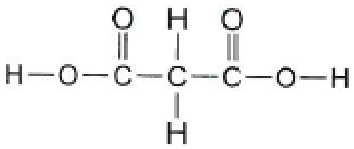
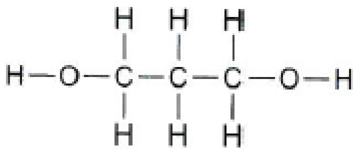
____

____

30. What type of link is used to join condensation polymers?

____

31. Complete the table below to show how a polyester is formed.

	
Dicarboxylic acid monomer	Diol monomer
Polyester repeating unit	

32. What are the problems associated with the use of polymers?

i) ___________________________________

ii) ___________________________________

iii) ___________________________________

iv) ___________________________________

33. What are the advantages and disadvantages of recycling polymers?

Advantages	Disadvantages

34. Complete the table for the naturally occurring polymers.

Polymer	**Monomer units**
DNA	
Starch	
Proteins	

35. Complete the table below to show the formula and molecular structure for each alcohol.

Alcohols	**Formula**	**Molecular structure**
Methanol		
Ethanol		
Propan-1-ol		
Butan-1-ol		

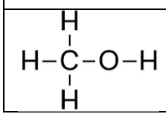
36. What is the functional group an alcohol?

____

37. When alcohols are dehydrated what do they form?

____

38. Complete the equation below for the dehydration of ethanol.

Ethanol	$\rightarrow$		+	
C_2H_5OH	$\rightarrow$		+	
	$\rightarrow$		+	

39. Why do we burn alcohols?

____

40. Complete the table below to show the formula and molecular structure for each carboxylic acid.

Carboxylic acids	**Formula**	**Molecular structure**
Methanoic acid		
Ethanoic acid		
Propanoic acid		
Butanoic acid		

41. What is the functional group of a carboxylic acid?

____

42. How do solutions of carboxylic acids behave?

____

43. What does ethanol, CH_3CH_2OH, produce when it is oxidised?

____

44. What do molecules within a homologous series have in common?

i) ______________________________

ii) ______________________________

iii) ______________________________

iv) ______________________________

v) ______________________________

45. How is ethanol produced by fermentation?

____

46. Complete the equation for the production of ethanol during fermentation.

Glucose	→		+	
$C_6H_{12}O_6$	→		+	

47. How do we produce a concentration solution of alcohol from fermentation?

____

48. What size are nanoparticles?

____

49. How does the size of nanoparticles compare to atoms and molecules?

____

50. What key property do nanoparticles have that make them useful?

____

51. Give some uses of nanoparticles.

____

52. What are the risks in using nanoparticles?

i) ____

ii) ____

iii) ____

53. Complete the table below to show the properties of the materials.

Bulk material	**Properties**	**Uses**
Glass		
Clay		
Polymers		
Concrete/laminates		

Paper 4 – topic 9 triple chemistry answers

1. Why must the qualitative test for an ion be unique?

So the presence of a certain ion can be clearly identified.

2. Describe how to carry out a flame test.

i) Dip the wire loop in hydrochloric acid.
ii) Place the loop into the Bunsen burner flame.
iii) Place the loop into distilled water.
iv) Dip the wire in to the salt to be tested.
v) Place the loop into the pale blue Bunsen burner flame.
vi) observe the colour produced.

3. Complete the table below to show the colour of the cations during a flame test.

Cation	Flame test colour
Lithium, Li^+	***Red***
Sodium, Na_+	***yellow***
Potassium, K_+	***Lilac***
Calcium, Ca^{2+}	***Orange-red***
Copper, Cu^{2+}	***Blue-green***

4. What is a precipitate?

An insoluble solid formed during a reaction.

5. Complete the table below to show the observations for the tests using sodium hydroxide. Complete an equation for each reaction.

Cation in solution	Description of the test	Observation	Chemical equation for the reaction
Aluminium, Al^{3+}	*Add a few drops of sodium hydroxide. Continue to add sodium hydroxide once the precipitate has formed*	*A white precipitate forms which dissolves when excess sodium hydroxide is added to form a colourless solution*	$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$ $Al(OH)_3(s) + OH^-(aq) \rightarrow Al(OH)_4^-(aq)$
Calcium, Ca^{2+}	*Add a few drops of sodium hydroxide to the test solution*	*A white precipitate of calcium hydroxide forms*	$Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$
Copper, Cu^{2+}		*A blue precipitate of copper (II) hydroxide forms*	$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
Iron (II), Fe^{2+}		*A green precipitate of iron (II) hydroxide forms*	$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$
Iron (III) Fe^{3+}		*A brown precipitate of iron (III) hydroxide forms*	$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
Ammonium, NH_4^+	*Add Sodium hydroxide. Gently heat the solution holding a piece of*	*No precipitate forms. Ammonia gas is produced when heated which turns*	$NH_4^+(aq) + 3OH^-(aq) \rightarrow NH_3(g) + H_2O(l)$

	damp red litmus over the top of the test tube.	*the red litmus paper blue.*	

6. Describe the chemical test for ammonia.

Add litmus paper to the gas. It will turn from red to blue.

7. Describe the test for carbonate ions, CO_3^{2-}.

Add hydrochloric acid to the sample.

Bubble any gas produced through limewater.

If the sample if a carbonate it will turn the limewater from clear to cloudy.

8. Describe the test for sulfate ions, SO_4^{2-}.

Add a few drops of hydrochloric acid to the sample followed by barium chloride.

If sulfate ions are present a white precipitate forms.

9. Give the ionic equation for the test for a sulfate ion.

$$Ba^{2+}\ (aq) + SO_4^{2-}\ (aq) \rightarrow BaSO_4\ (s)$$

10. Describe the test for halide ions.

Add dilute nitric acid to the sample followed by silver nitrate. A precipitate forms if a halide ion is present.

11. Complete the table below to show the results and equations for halide ions.

Halide ion	Observation	Equations
Chloride, Cl^-	***White precipitate forms***	$Ag^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s)$
Bromide, Br^-	***Cream precipitate forms***	$Ag^+\ (aq) + Br^-\ (aq) \rightarrow AgBr\ (s)$
Iodide, I^-	***Yellow precipitate forms***	$Ag^+\ (aq) + I^-\ (aq) \rightarrow AgI\ (s)$

12. What are the advantages of instrumental qualitative tests?

i) improved sensitivity

ii) improved accuracy

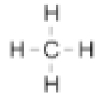
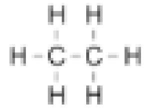
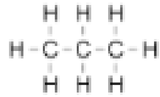
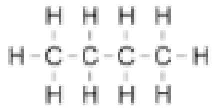
iii) improved speed

13. Describe two ways that a flame photometer is used to identify ions.

i) to find the concentration of an ion using a calibration curve.

ii) by comparison of data against reference data.

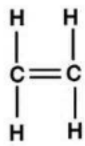
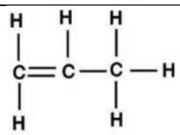
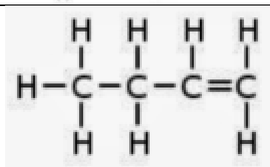
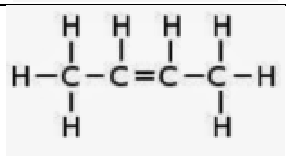
14. Complete the table below to show the formula and molecular structure for each alkane.

Alkanes	Formula	Molecular structure
Methane	***CH_4***	
Ethane	***C_2H_6***	
Propane	***C_3H_8***	
Butane	***C_4H_{10}***	

15. Why are alkanes saturated hydrocarbons?

The contain carbon-carbon single bonds (C-C) and carbon and hydrogen atoms only.

16. Complete the table below to show the formula and molecular structure for each alkene.

Alkenes	Formula	Molecular structure
Ethene	***C_2H_4***	
Propene	***C_3H_6***	
But-1-ene	***C_4H_8***	
But-2-ene		

17. Why are alkenes unsaturated hydrocarbons?

The contain carbon=carbon double bonds (C=C) and carbon and hydrogen atoms only.

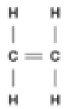
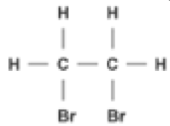
18. Describe the test to identify alkanes and alkenes.

Add bromine water to the sample.

In alkenes the bromine water changes from yellow to colourless.

In alkanes it remains yellow.

19. Complete the table below to show the reactions of bromine water with ethene.

Ethene	+	Bromine	→	1, 2 – dibromoethane
C_2H_4	**+**	***Br_2***	**→**	***CH_3BrCH_3Br***
	**+**	Br_2	**→**	

20. What type of reaction is combustion?

It is an oxidation reaction.

21. What are the products of the complete combustion (oxidation) of alkanes and alkenes?

Carbon dioxide and water.

22. Complete the reactions below to produce balanced equations for the combustion of alkanes and alkenes.

Methane	+	Oxygen	→	***Carbon dioxide***	+	***Water***
CH_4	**+**	***$2O_2$***	**→**	***CO_2***	**+**	***$2H_2O$***
Ethene	+	Oxygen	→	***Carbon dioxide***	+	***Water***
C_2H_4	**+**	***$3O_2$***	**→**	***$2CO_2$***	**+**	***$2H_2O$***
Ethane	+	Oxygen	→	***Carbon dioxide***	+	***Water***
$2C_2H_6$	**+**	***$7O_2$***	**→**	***$4CO_2$***	**+**	***$6H_2O$***

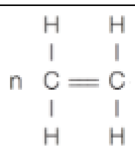
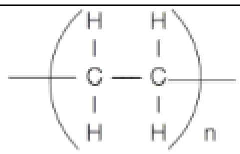
23. What is a polymer?

A substance of high average relative molecular mass made up of small repeating units.

24. What is a monomer?

The repeating unit which is joined together.

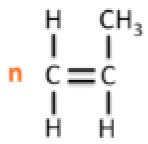
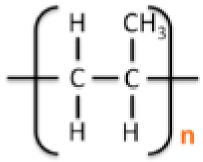
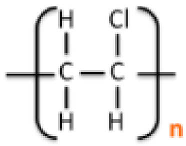
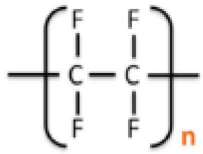
25.Complete the equation below to show the formation of polyethene from ethene.

	
Ethene	Poly(ethene)

26. What is needed to make additional polymers?

A monomer containing C=C bonds.

27. Draw the monomers and polymers below.

Monomer	Polymer structure
 Propene	 Poly(propene)
 Chloroethene	 Poly(chloroethene) (PVC)
 Tetrafluoroethene	 Poly(tetrafluoroethene) (PTFE)

28. Complete the table below to show the properties and uses of each polymer.

Polymer	**Properties**	**uses**
Poly(ethene)	***Cheap, flexible, good insulator***	***Plastic bags and bottles, cling film***
Poly(propene)	***Flexible, does not shatter***	***Buckets, crates, rope and carpet***
Poly(chloroethene) (PVC)	***Tough, good insulator***	***Window frames, gutters, pipes, electrical wire insulation***
Poly(tetrafluoroethene) (PTFE)	***Tough, slippery, non-toxic***	***Non-stick coating on pans, stain proof clothing***

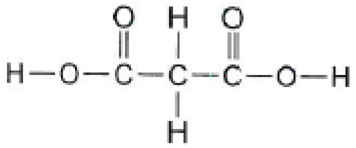
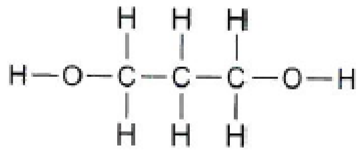
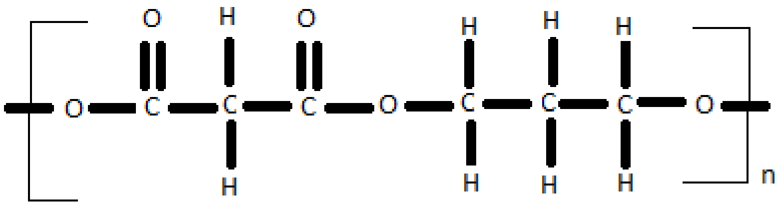
29. What is condensation polymerisation?

The joining together of monomers to release water. Each monomer must contain two functional groups e.g. dicarboxylic acids joining diols.

30. What type of link is used to join condensation polymers?

An ester link.

31. Complete the table below to show how a polyester is formed.

	
Dicarboxylic acid monomer	Diol monomer
	
Polyester repeating unit	

32. What are the problems associated with the use of polymers?

i) availability of starting materials.

ii) persistence in landfill sites, they are non-biodegradable.

iii) gases produced during their disposal by combustion.

iv) requirement to sort polymers so that they can be melted and reformed into a new product.

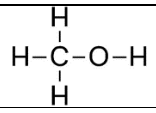
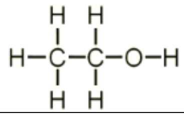
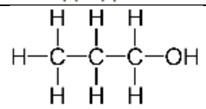
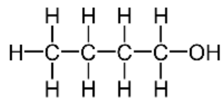
33. What are the advantages and disadvantages of recycling polymers?

Advantages	**Disadvantages**
Prevents disposal to landfill or incinerator	Collecting the polymers
Prolongs a finite resource	Sorting the polymers

34. Complete the table for the naturally occurring polymers.

Polymer	**Monomer units**
DNA	***Four different nucleotide***
Starch	***Sugars***
Proteins	***Amino acids***

35. Complete the table below to show the formula and molecular structure for each alcohol.

Alcohols	Formula	Molecular structure
Methanol	***CH_3OH***	
Ethanol	***C_2H_5OH***	
Propan-1-ol	***C_3H_7OH***	
Butan-1-ol	***C_4H_9OH***	

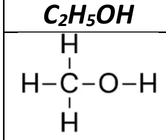
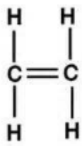
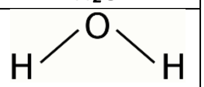
36. What is the functional group an alcohol?

The –OH group.

37. When alcohols are dehydrated what do they form?

An alkene

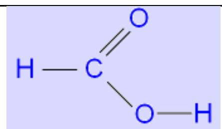
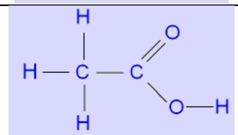
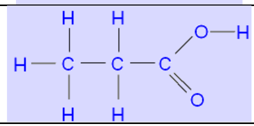
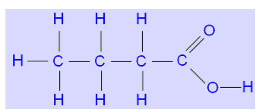
38. Complete the equation below for the dehydration of ethanol.

Ethanol	→	Ethene	+	Water
C_2H_5OH	**→**	***C_2H_4***	**+**	***H_2O***
	→		+	

39. Why do we burn alcohols?

The combustion of alcohols releases energy.

40. Complete the table below to show the formula and molecular structure for each carboxylic acid.

Carboxylic acids	Formula	Molecular structure
Methanoic acid	***HCOOH***	
Ethanoic acid	***CH_3COOH***	
Propanoic acid	***C_2H_5COOH***	
Butanoic acid	***C_3H_7COOH***	

41. What is the functional group of a carboxylic acid?
-COOH

42. How do solutions of carboxylic acids behave?
They have the properties of typical acids.

43. What does ethanol, CH_3CH_2OH, produce when it is oxidised?
Ethanoic acid, CH_3COOH

44. What do molecules within a homologous series have in common?
i) The same general formula
ii) The same functional group
iii) Similar chemical properties
iv) a trend in their physical properties
v) increase by $-CH_2$ group each time.

45. How is ethanol produced by fermentation?
Yeast contains enzymes. The yeast is added to a solution of carbohydrate and ethanol is produced.

46. Complete the equation for the production of ethanol during fermentation.

Glucose	→	***Ethanol***	***+***	***Carbon dioxide***
$C_6H_{12}O_6$	→	***C_2H_5OH***	***+***	***CO_2***

47. How do we produce a concentration solution of alcohol from fermentation?
Carry out fractional distillation of the fermentation mixture.

48. What size are nanoparticles?
Between 1 and 100nm in diameter

49. How does the size of nanoparticles compare to atoms and molecules?
Nanoparticles are bigger than atoms and molecules.

50. What key property do nanoparticles have that make them useful?
They have a large surface area to volume ratio.

51. Give some uses of nanoparticles.
Catalysts, lubricant coating, electrical conductors, surgical antibacterial materials.

52. What are the risks in using nanoparticles?
i) They can be breathed in or could pass between cell-surface membranes.
ii) They could carry toxic substances on their surfaces.
iii) They could catalyse harmful reactions.

53. Complete the table below to show the properties of the materials.

Bulk material	Properties	Uses
Glass	***Unreactive, brittle, transparent***	***Windows***
Clay	***Can be shaped***	***Bricks, porcelain, china, tiles***
Polymers	***Moulded, strong, unreactive, poor conductors of heat and electricity***	***Insulation, window frames, water pipes, waterproof flooring***
Concrete/laminates	***Mixture of 2 or more materials, strong,***	***Building materials,***