CfE Higher Chemistry Summary Notes

Unit 1

**Section 1A – Controlling the Rate**

1. The rate of a reaction can be followed by measuring changes in concentration, mass

or volume.

Change in …………………………

Average Rate =

………………………… taken

Units are mol l-1 s-1 (for concentration change), ………………. (for mass change), …………………… (for volume change)

2. The rate of a reaction is proportional to 1/........................................ (units: s-1).

3. Reaction rate increases as the concentration of a reactant.........................................

As the concentration of a reactant increases the number of collisions between reactant

particles ............................................ .

4. Reaction rate increases as the particle size of solid reactants .........................................

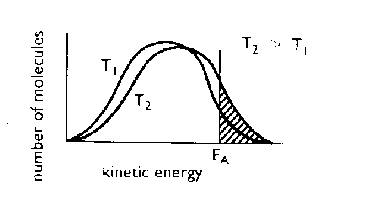
The smaller the particle size the ....................................... the surface area of the solid.

The bigger the surface area the ................................ collisions there are between reactant

particles.

5. Reaction rate increases as the temperature at which the reaction is carried out ...............................

Temperature is a measure of the average ............................................. energy of the particles of a substance. The higher the kinetic energy of the particles the ............................................ the number of collisions between reactant particles.



6. The .......................................... energy (EA) is the minimum kinetic energy required

by colliding particles before reaction will occur. Increasing the temperature

......................................... the number of particles which possess the

.................................................. energy (EA) and so increases the rate of the reaction.

A ............................... increase in temperature can give a large increase in reaction rate.

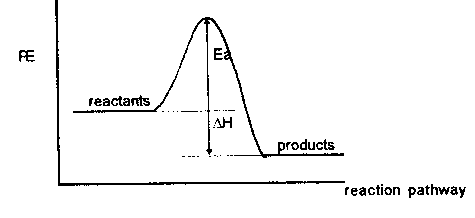
7. With some chemical reactions light energy can be used to increase the number of particles with the necessary activation energy. One such reaction is ........................................................... in which green plants use light energy to make starch.

8. In an ............................................... reaction heat energy is released to the surroundings.

In an ............................................... reaction heat energy is absorbed from the surroundings.

9. In an exothermic reaction the reactants have ........................... stored energy than the products.

In an endothermic reaction the reactants have ........................... stored energy than the products.



10. A potential energy diagram can be used to show the energy stored in both

reactants and products. The enthalpy change (H) for the reaction is given

by Hproducts - .............................................. The enthalpy change has a negative

value in an ............................................... reaction and a positive value in an

.................................................... reaction.

*exothermic reaction*

11. The activated ................................... is an unstable arrangement of atoms formed at the maximum of the potential energy barrier during a reaction .

12. The ........................................... ................................... (EA) is the minimum energy required by the reactant particles in order to form the activated complex.

13. A catalyst speeds up a reaction by providing an alternative reaction pathway with a .............................activation energy.

**Section 1B– Periodicity**

1. The Modern day Periodic Table is based on the work of the Russian chemist .............................................who arranged the elements in order of increasing .................................... mass and left spaces for yet to be .................................................... elements.

The present day Periodic Table actually arranges the elements in order of increasing atomic number(the number of .............................. in the nucleus of an atom).

2. The covalent radius of the elements ................................................ from left to right across a period and .................................................. from top to bottom within a group.

3. The .................................................. in covalent radius from left to right across a group is due to an................................................. in the number of protons (positive charges) in the nucleus attracting the electron energy levels ......................................... to the nucleus.

4. The .................................................. in covalent radius from top to bottom in a group is due to an.............................................. in the number of occupied electron energy levels as we go down the group.

Each new energy level is ..................................... away from the nucleus making the atom .............................in size.

5. The first ........................................................ energy is the amount of energy that has to be supplied to remove one mole of electrons from one mole of neutral atoms in the gas state.

ie for sodium ................ + ........... H= .................... kJ mol-1 (data booklet page 11)

6. The second ...................................... energy is the amount of energy to remove one mole of electrons from one mole of +1 ions in the gas sate.

ie for sodium  ................... + ...................... H= .................. kJ mol-1 (data booklet page 11)

7. The second ionisation energy is higher than the first due to the electron having to be removed from a ......................... ion rather from a .......................... atom.

8. The first ionisation energy ..................................... from left to right across a group due to the increasing number of ................................. in the nucleus attracting the electrons more ...........................

This increased attraction means that more energy has to be supplied to remove an electron from the atom.

9. The first ionisation energy ............................... from top to bottom within a group due to the

increasing number of occupied electron ................................... .......................................... Each new energy level that is occupied is ........................... away from the nucleus. Each new energy level provides additional screening and so it requires ............................. energy to remove an electron.

10. Atoms of different atoms have different attractions for bonding electrons. The

.................................................................. value is a measure of the attraction that an atom has for electrons it shares with other atoms.

11. The higher the electronegativity value of an atom the more ...................................... it attracts bonding electrons towards itself.

12. Electronegativity values ..................................... from left to right across a period. This is due to an increase in the number of ................................ in the nucleus of the atoms attracting the shared electrons in the bond more strongly.

13. Electronegativity value ...................................... from top to bottom within a group. This is due to the greater ..................................... effect on more occupied electron energy levels.

**Section 1C – Structure & Bonding**

**a. Elements**

1. Metallic bonding is an electrostatic attraction between the ................................charged nucleus and the delocalised outer ..................................................

2. A metallic structure consists of a giant .................................... of positive ions surrounded by

...................................................... outer electrons.

**- - - - - - -**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

**+**

lattice of .........................

charged nuclei

**- - - - - -**

delocalised outer

...............................

**-**

3. Atoms of non-metal elements bond together by sharing electrons – this type of bond is known as a ......................................... bond. In a ..................................... bond the atoms are held together by an electrostatic attraction between the ........................................ charged shared electrons and the ............................................ charged nuclei.

4. A covalent molecular structure consists of discrete molecules with ............................... bonds

between the atoms inside the molecule and weak ...................... ..................... ............................. forces between the molecules.

**H**

covalent bond

van der Waals force

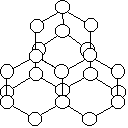
**H**

**H**

eg

**H**

5. A covalent network structure consists of a giant ........................................ of covalently bonded atoms.



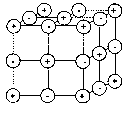
covalent bond

eg diamond

6, Ionic bonding usually occurs between metal and ........................................ atoms.

Ionic bonding is an electrostatic attraction between the .......................................... charged metal ions and the ........................................... charged non-metal ions. Ionic bonding occurs when there is a …………………… difference in electronegativity.

7. An ionic structure consists of a giant .................................... of oppositely ...................................... ions.



eg sodium chloride

8. Pure covalent bonding and ionic bonding can be considered as being at opposite ends of a bonding ……………………………… with …………….. covalent bonding lying between these two extremes.

9. Intermolecular forces, known as ………… …………… ………………… forces, occur between all molecules.

10. …………………… …………………….forces are weak forces of attraction that exists between all atoms and molecules. .................... ................... forces are much .................................. than all other types of bonding.

11. …………………… ………………… forces are caused by small electrostatic attractions between ...................................... dipoles created in atoms by an uneven distribution of electrons in the electron cloud round an atom.

12. London dispersion forces ................................... in strength as we go down a group.

Bigger atoms have more electrons and so the size of the temporary dipoles ..................................... as we go down the group.

13. A molecule is described as being polar if the it has a ............................. dipole.

eg hydrogen chloride +H-Cl-.

14. The electrostatic attraction between molecules with permanent dipoles are .................................... than the weak London dispersion forces between molecules with .................................................. dipoles.

15. Symmetrical molecules with polar bonds are themselves ................................................ – the symmetry cancels out the polarity of the bonds.

For example, +O=C+=O- is a non-polar molecule as it has a ........................................ shape.

16. Bonds in which a hydrogen atom are bonded to a much more electronegative element are

very strongly polar. The elements which are much more electronegative are nitrogen, ........................,....................................... and .......................................... Hydrogen bonds are .................................... forces ofattraction between these highly polar molecules eg +H-F+H-F hydrogen bonds).

17. Hydrogen bonds are ............................ than other forms of permanent dipole attractions but they are weaker than covalent bonds.

18. A monatomic structure consists of individual atoms with weak .................... .....................

forces between the atoms.

London Dispersion force

He He

19. The first 20 elements in the Periodic Table can be categorised according to bonding and structure.

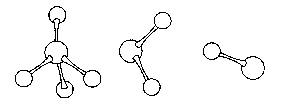
* Metallic (Li, Be, .........., ............, .............., ..............., ................, )
* Covalent molecular gases (H2, N2, ............, ................, .................)
* Covalent molecular solids (P4, ................ and fullerines C60)
* Covalent networks (B, ........... and .............)
* Monatomic (He, Ne, .........., ............., ................, .................)

**b. Compounds**

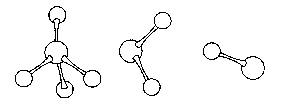
1. Compounds adopt one of three different structures.

* Covalent molecular – discrete molecules with ................................ bonds between the atoms inside the molecule and weak ................ ................. ..................... forces between the molecules.

.................................................... bond between atoms within the molecule



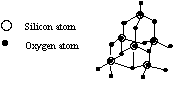
eg methane (CH4)



weak van der Waals forces

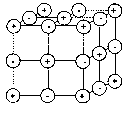
....................... the molecules

* Covalent network – a giant ............................ of covalently bonded atoms.



eg silicon dioxide

* Ionic - a giant ............................... of oppositely charged ions



eg sodium chloride

# c – Properties associated with bonding

1. The melting and boiling points of polar substances is ................................... than the melting and boiling points of non-polar substances with similar molecular mass.

2. Ionic and polar substances tend to be ........................................ in polar solvents such as water.

Non-polar substances tend to be .......................................... in polar solvents such as water

3. Water (H2O), ammonia (............) and hydrogen fluoride (...........) have ................................... than expected boiling points due to ............................................ bonding between the polar molecules.

4. Hydrogen bonding between water molecules results in ice having a very open structure.

This results in ice having a ....................................... density than liquid water and so ice .....................on water.

5. Diamond is a very .................... and strong substance due to its giant ....................................... structure.

In diamond every carbon atoms is bonded to .................. other atoms. This means that all 4 electrons in a carbon atoms ............................ shell are bonded and so carbon does not ............................................electricity.

6. In graphite the carbon atoms are arranged in ........................................ with weak van der Waals forces between the layers. This makes graphite ......................... and slippery. As only 3 of the 4 electrons in the outer energy level are shared there are ........................................... electrons which makes it an electrical ............................................

7. Silicon carbide (SiC) has a giant covalent ....................................... structure. This makes silicon carbide very ........................ and strong. It is used in cutting and grinding tools.

CfE Higher Chemistry Summary Notes

Unit 2

**Section 2A – Esters, Fats & Oils**

1. Alcohols contain the ......................................... (-OH) functional group and have names that end in –ol.

The general formula for straight chain alcohols (alkanols) is ...........................................

2. Carboxylic acids contain the ......................................... ( -COOH) functional group. Carboxylic acids have a name ending of -......................... acid.

eg

is called ..................................................... acid.

3. Esters are used as solvents, ....................................... and ..............................................

4. Esters contain the ester functional group ( .....................) and have a name ending of –oate.

The first part of an esters name comes from the parent alcohol eg ethanol gives an ....................... ester.

The second part of the ester name comes from the parent carboxylic acid eg propanoic acid gives a ......................................... ester.



eg

....................................... + ............................... acid --> .............................. + water

5. Esters are made in a ................................................... reaction between an alcohol and a

...................................................... acid.

 eg

+ H2O

....................................

.................................... acid ........................... methylethanoate + water

 The shows that the reaction is ....................................................

6. The breaking of an ester into an alcohol and a carboxylic acid is an example of a .................................reaction.



eg

+ H2O

........................ + ……………………….

…………….……………………………………….. + …….. ………………………..

7. Fats and oils are naturally occurring esters. Fats and oils can be classified according to their origin

* animal, ......................... or ......................................

8. Oils have ........................................... melting points than fats. The ....................................... melting points of oils is due to the oils being ............................................................. molecules. The fact that oils are..................................................... can be shown by the fact that oils quickly decolourise .............................solution.

9. Fat molecules are more linear in shape than oil molecules. This means that fat molecules can pack together .................... closely than the more angular oil molecules. Since the fat molecules can pack together more closely than oil molecules the van der Waals forces between fat molecules are ..................................... than between oil molecules.

10. The conversion of unsaturated oils into solid fats in margarine manufacture involves the addition of ......................................... (hydrogenation) across some of the double bonds in the oil molecules.

11. Fats and oils are a ....................... concentrated source of energy in the body than carbohydrates.

12. The hydrolysis of a fat (or oil) produces a mixture of ...................................... (propane-1,2,3-triol) and ............................. acids (long chain carboxylic acids) in the ratio of one molecule of

........................................... to three molecules of fatty acid.

13. Glycerol (propane-1,2,3-triol) is a trihydric alcohol. This means it contains three ..........................(-OH) functional groups per molecule.

14. Fats and oils are triglycrerides formed when one molecule of .......................................... combines with ............. fatty acid molecules. The fatty acid molecules may, or may not, be identical.

**Section 2B– Proteins**

1. Proteins are condensation polymers made by combining together many ......................... ...........................monomer molecules.

2. Proteins are required by the body for ........................ and tissue ...................................

3. Proteins and amino acids contain the elements carbon, hydrogen, oxygen and...........................

4. The body cannot make all the amino acids required to make body proteins. The supply of these .............................................. amino acids depends on the diet we eat.

5. Amino acids are molecules that contain a basic amino (...............) group and an acidic .............................(-COOH) group.

The simplest amino acid is 2-aminoethanoic acid. The extended structural formula is shown below

6. Shown below is a small part of a protein polymer chain



Circle the peptide (amide) links shown in this part of the protein molecule.

Draw the extended structural formula of the two amino acids produced by the hydrolysis of this part of the protein molecule.

7. Protein molecules can be classed as fibrous (..................... and thin) or globular (spiral chains folded into ........................). Fibrous proteins are form many of the bodies major structural units such as muscles and tendons. Globular proteins form many of the enzymes and hormones in the body. For example haemoglobin in .................. blood cells and insulin are both ............................. proteins.

**Section 2C – Chemistry of Cooking**

1. Some flavour molecules are soluble in water because they are ………………………….. If they are cooked in ………………………….. the flavour molecules will be lost.

2. Some flavour molecules are soluble in fats and oils because they are ………………………….. If they are cooked in ………………………….. the flavour molecules will be lost.

3. When protein molecules are heated the ………………………….. bonds holding them in shape are broken. The protein molecules can unwind and lose their ……………………….. This is called …………………………….. the protein.

4. Aldehydes and ……………………………….. are common flavour molecules.

5. Aldehydes and ketones contain the .......................................... (-C=O) functional group.

6. In aldehydes the ................................. (-C=O) group is on the ................ of the carbon chain.

 In ketones the .................................... (-C=O) group is **not** on the ..................... of the carbon chain.



eg

and

is the .............................. called ............................. and this is the ............................. called ...........................

7. Aldehydes have a name ending in ..................... while ketones have a name ending in ........................

8. Aldehydes but **not** ketones can be oxidised to ................................................. acids.

Suitable oxidising agents include Fehlings solution, acidified …………………………….. …………………… and …………………………………… reagent..

 eg

+ O

butanal .......................................................

**Section 2D – Oxidation of Food**

1. Alcohols burn in oxygen, or air, forming ..................................... ................................... and water.

 eg. .................. + ............................

2. Alcohols can be dehydrated (loss of .............................) by passing the alcohol vapour over heated aluminium oxide (catalyst) to form alkenes.



eg.

-H2O

propan-1-ol ..................................

3. Alcohols can be classified as primary, ........................................... or .................................................

 eg

propan-2-ol is a 2-methylpropan-2-ol is a propan-1-ol is a

........................... alcohol ............................... alcohol ............................ alcohol

4. Primary and secondary alcohols can be oxidised using oxidising agents such as

a) acidified potassium dichromate solution. In this reaction the orange .............................. ion is reduced to the ............................ chromium(III) ion.

b) hot copper(II) oxide. In this reaction the black copper(II) oxide is reduced to ....................... copper metal.

5. Oxidation of a primary alcohol first produces an ............................................... which then is oxidised further to a ............................................... acid.



eg

+ O

- 2H

ethanol .................................. ………………… ………….

6. Oxidation of a secondary alcohol produces a ...........................................



eg

- 2H

propan-2-ol ...................................

7. Some alcohols contain two or more hydroxyl groups. In the space below draw the named diol and triol.

ethane-1,2-diol propane-1,2,3-triol (known as ………………..)

8. Diols and triols contain multiple hydroxyl groups. This means they can undergo …………………….bonding. This causes an …………………………….. in viscosity and boiling point as the intermolecular attractions are ………………………………..

9. Carboxylic acids can react with alkalis in a …………………………………….. reaction to produce a ……………………… and …………………….

e.g. ethanoic acid + …………………. ……………………. → ………………… ………………. + water

………………………. + NaOH → ……………………. + H2O

10. Salts of carboxylic acids can be used as …………………………………. to stop food spoiling.

11. A dilute solution of ethanoic acid is used to ................. foods like onions. The acids prevents the growth of ………………… and ……………………….

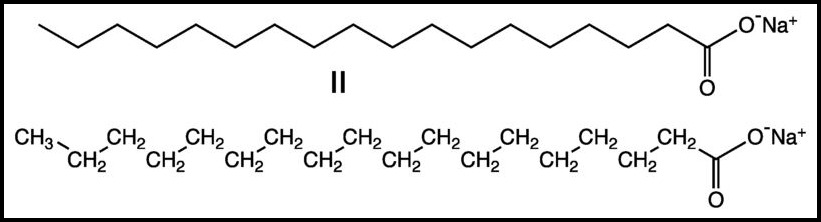
12. Vitamin C (ascorbic acid) is a common ………………………….. The vitamin C molecule will easily be oxidised – this saves the food from becoming oxidised.

**Section 2E - Soaps Detergents & Emulsions**

1. Soaps are produced by the alkaline hydrolysis of ................... and .....................

2. Soaps are the sodium or potassium salts of ………………….. …………………….

3. Label the hydrophobic and hydrophilic parts of this soap molecule.



4. The hydrophobic …………………….. dissolves in ……………………. And the hydrophilic ……………… dissolves in …………………….. The charged ……………… repel each other to prevent the fats from sticking together.

5. Hard water contains dissolved calcium ( ) and ………………………………. ( Mg2+) ions.

When hard water is mixed with soap it forms a **………………….** instead of a lather.

6. In hard water areas **………………………………** are used instead of soaps. A detergent has a structure similar to soap but it will not form a …………………..

7. An **……………………………** contains small droplets of one liquid dispersed in another liquid. They are common in food e.g. in salad dressing, mayonnaise and milk.

8. Oil and water normally separate into layers. To prevent this happening an **………………….** is added.

9. Emulsifiers are soap like molecules – they have a …………………………….. part and a ………………………….. part.

**Section 2F - Fragrances**

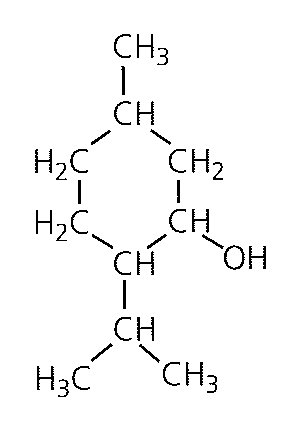
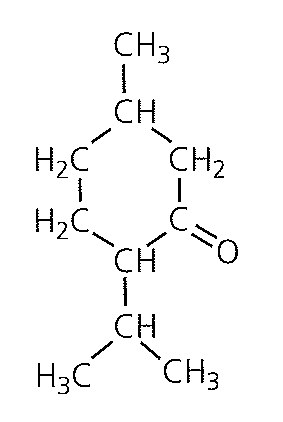
1. Essential oils are concentrated extracts from ………………………….., and are mixtures of organic compounds.

2. …………………………………….. are key components in most essential oils.

3. Terpenes are unsaturated compounds formed by joining together …………………………… (2-methylbuta-1,3-diene) units. Draw an isoprene molecule in the space below.

4. Terpenes can be oxidised to form new compounds which have different ……………………….. from the original terpene. This can alter the fragrance and other properties of a product. During oxidation the ratio of oxygen to hydrogen …………………………………..

e.g. the oxidation of menthol to menthone:

O:H ratio: …………. ……………..

**Section 2G – Skin Care**

1. …………………… radiation (UV) is a high-energy form of light, present in sunlight. UV light has enough ……………………………..to break chemical bonds. It is the UV light that is responsible for …………………… and ageing our skin.

2. The energy from ………… ……………… can break …………………………. bonds. This leaves an atom or group of atoms with an unpaired ………………………. These highly ……………………… particles are called ………………….. ………………………..

3. A free radical reaction (……………………. reaction) can be broken down into three stages:

Stage 1 is called ………………………………….. This is when the ………………………. ………………………… are formed and the reaction starts.

Stage 2 is called ………………………………….. This is when more ………………. ………………………. are formed and the reaction continues again and again.

Stage 3 is called ……………………………………… This is when the free radicals join together to form a ……………………… compound and the reaction …………………………..

4. Free radical reactions on our ……………………. can cause ageing and wrinkles. Free radical …………………………………… can be added to cosmetics to prevent this. Vitamins …….. and ……….. are examples of natural free radical …………………………………..

Unit 3

**3A – Getting the Most From Reactants**

1. The chemical industry is a major contributor to the ...................................... of the UK.

2. Stages in the production of a new product involve research, pilot production, scaling up to

............................... scale production and review.

3. A ...................................... is a chemical from which other chemicals can be synthesised or extracted.

4. The major ............. materials used by the chemical industry are fossil ..................., metal .............. and minerals, air and ...............................

5. You must be able to carry out calculations using a balanced equation.

e.g. What mass of carbon dioxide is formed when 4g of methane is burned?

CH4 + 2O2  ****  CO2 + 2H2O

GFM’s:

CH4

CO2

1 mole  1 mole

4 x 1 = 4 1x 12 = 12

GFM= 16

n=m/GFM

= 4/16

2 x 16 =32

1x 12 = 12

GFM= 44

= 0.25 moles 0.25 moles

m = n x GFM

= 0.25 x 44

= 11g

Examples for practice.

(a) Calculate the mass of oxygen that can be formed when 68g of hydrogen peroxide (H2O2) decomposes.

2H2O2  2H2O + O2

(b) Calculate the mass of calcium carbonate required to react completely with 300cm3 of 0.1 moll-1 hydrochloric acid.

CaCO3 + 2HCl → CaCl2 + CO2 + H2O

6. Excess calculations can be carried out to work out which ………………………….. is in ………………. The reactant which determines the quantity of product made is called the …………………………. reactant.

Example

30cm3 of 0.1 moll-1 copper(II) sulfate solution was added to 2.6g of zinc.

Zn + CuSO4 → ZnSO4  + Cu

a. Show by calculation that zinc was in excess.

b. Calculate the mass of copper produced.

7. The volume of one mole of gas at a specified temperature and pressure is known as the **……………. …………………..**.

volume = moles x molar volume

The molar volume of all gases is approximately the ……………………………..

Example:

Calculate the volume of 0.05 moles of carbon dioxide gas assuming the molar volume is 24lmol-1.

8. Molar volumes can also be used in balanced equation calculations.

Example:

Calculate the volume of hydrogen produced when 13.1g of zinc are added to excess dilute hydrochloric acid. Take the molar volume to be 24 litres mol-1.

Zn + 2HCl → ZnCl2 + H2

9. Percentage yield can be used to work out how ………………… a reaction is.

Percentage yield = actual yield x 100

theoretical yield

Example:

Excess ethene was reacted with 3.65g of hydrogen chloride and 4.1g of the product was made. Calculate the percentage yield.

C2H4  + HCl → C2H5Cl

11. **Atom economy** allows us to examine the proportion of ……………………… converted into desired ……………………..

Atom economy = mass of desired product(s) x 100

total mass of reactants

Example:

Calculate the atom economy for the production of carbon dioxide in the following reaction.

Fe2O3 + 3CO → 2Fe + 3CO2

**3B – Equilibria**

1. Reversible reactions attain a state of dynamic equilibrium when the ................... of the forward reaction equals the .......................... of the ................................ reaction.

2. At equilibrium the concentrations of the reactants and products remains..................................

It is important to realise that at equilibrium the concentrations of reactants and products will not usually be .....................................

3. Increasing the concentration of one of the reactants will speed up the .................................... reaction and shift the equilibrium from ....................... to right ie it will result in an increased yield of...............................

4. ............................................ the temperature will favour the endothermic step in a reversible reaction.

............................................. the temperature will favour the exothermic step in a reversible reaction.

5. Increasing the pressure will favour the reaction which forms fewer ........................ molecules.

Decreasing the pressure will favour the reaction which forms more ....................... molecules.

6. In the Haber process



The yield of ammonia will increase by ........................................ the temperature, .................................. the pressure and recycling unreacted nitrogen and hydrogen (as this has the same effect as increasing the ........................................... of these reactants).

7. Low temperatures favour the formation of .......................................... in the Haber Process. However, if the temperature is too low the ............................. of production of ammonia is too slow.

8. A catalyst speeds up the rate at which a reaction reaches .................................................

Using a catalyst does not, however, alter the position of the equilibrium as it speeds up the forward and reverse reactions by the ........................ amount.

**3C – Chemical Energy**

1. ………………………………… reactions release energy into the surroundings (ΔH is …………………) and ………………………………… reactions absorb energy from their surroundings (ΔH is …………………). There are costs involved in supplying energy to …………………………….. reactions and cooling down ……………………… reactions.

2. Chemical energy is also known as ……………………

3. The enthalpy of …………………………….. of a substance is the enthalpy change when one mole of a substance burns ………………………… in oxygen. This is always an …………………………….. reaction.

The enthalpy of combustion of ethane is shown below:

C2H6 + 3½O2 → …………….. + ………………. ΔH =

4. Enthalpy change can be calculated using the relationship H= - cmT where

c is the ..................................................................................... of the substance (usually water with a value of 4.18 kJ kg-1 oC-1)

m is the mass in ................................................ (kg) of the solution being heated.

T is the temperature change (in ............).

5. The enthalpy of ………………………… is the enthalpy change when one mole of a substance dissolves in water. This can be an ………………………………….. reaction or an …………………………… reaction.

6. The enthalpy of ………………………… is the enthalpy change when an acid is neutralised to produce one mole of water. This is an ………………………………….. reaction and is approximately …………………… for every acid.

7. The enthalpy of ………………..……. of a compound is the energy change involved in the formation of one …………….…. of the compound from its constituent ………………. in their standard states i.e. their states at …….°C and …… atmosphere pressure.

8. ……………………… law states that the enthalpy change of a chemical reaction will be …………. …………….. no matter which route is taken.

9. Enthalpy changes can be calculated using Hess’s Law.

Example: 2C + 3H2 → C2H6 ∆H ?

Enthalpies of combustion of carbon, hydrogen and ethane are

a. C + O2 → CO2 ∆H = …………….kJmol-1

b. H2 + ½ O2 → H2O ∆H = …………….kJmol-1

c. C2H6 + 2O2 → 2CO2 + 3H2O ∆H = …………….kJmol-1

ΔH = 2a + 3b + (-c)

=………… + ……………+ …………….

= …………………… kJmol-1

10. Bond enthalpy is the ………………….. required to ………………… one mole of bonds. Mean molar bond enthalpies are …………………… values which are quoted for bonds which occur in …………………. molecular environments.

11. Breaking bonds is an …………………………… process and making bonds is an ………………………. process.

**Enthalpy change = total bonds broken + total bonds made**

Example:

H2(g) + Br2(g) → 2HBr(g)

Bonds broken: Bonds made:

H-H 432kJmol-1

Br-Br …………kJmol-1

total = …………. kJmol-1 total = ……………….kJmol-1

Enthalpy change = …………… + ………………..

= ………………. kJmol-1

**3D – Oxidising & Reducing Agents**

1. Oxidation involves a ........................... of electrons. Reduction involves a ...................... of electrons.

2. An oxidising agent is itself ......................................... An oxidising agent is an electron ...............................

A reducing agent is itself ........................................ A reducing agents is an electron ....................................

For example in the displacement reaction shown below

 2+

In this reaction the zinc is the .......................................... agent. The ion-electron equation for the oxidation of the zinc atoms is:-

...........  ................... + .................

In this reaction the silver(I) ions are the .......................................... agent. The ion-electron equation for the reduction of the silver(I) ions is:-

............... + .............  ..................

In this reaction the nitrate ions are called the ........................................................... ions.

3. Elements and compounds at the top right of the electrochemical series are strong …………………………. agents.

Elements and compounds at the bottom left of the electrochemical series are strong ………………………… agents.

4. To combine together the oxidation and reduction ion-electron equations into the balanced

redox equation we must first balance the number of .................................... lost and gained in the oxidation and reduction reactions.

 eg oxidation

 reduction

To construct the redox equation we must first multiply the oxidation reaction by ...........

 This gives

Adding together the oxidation and reduction reactions gives the balanced redox equation

.............................................................................................................................................................................

From this equation it can be seen that ............ moles of Fe2+(aq) ions are oxidised by ............... mole of MnO4-(aq) ions.

5. Ion-electron equations not listed in the data booklet can be constructed.

eg Construct the ion-electron equation for the conversion of iodate (IO3-) ions to iodine.

Step 1:- balance the iodine atoms on each side:



Step 2:- Add H2O(l) molecules to balance the oxygen atoms on each side



Step 3 :- Add H+(aq) ions to balance the hydrogens on each side



Step 4:- Add electrons to balance the charge on each side



In this reaction the iodate is ................................. to iodine. The iodate is acting as an

............................................ agent.

6. You must be able to carry out calculations based on redox titrations using the relationship

C1V1 = C2V2

n1 n2

where c = ……………………………

V = ……………………………

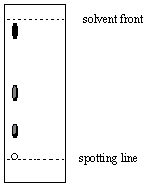
n = ………………………………………………………

**3E - Chemical Analysis**

1. In chromatography, differences in the polarity and/or …………….. of molecules are exploited to ……………………… the components present within a mixture.

2. In ………………….chromatography a small sample of the mixture is placed at the bottom of the paper and the paper is dipped in the …………………. By capillary action, the ……………….. moves up the paper. Different compounds move up the paper at different ……………………. Thus the components of the mixture are …………………………………..

3. It is common to measure the distance the sample has moved from its original starting point and compare it with the ………………………. the solvent has moved. This allows us to calculate ratios known as ………. values.



Rf = \_\_\_\_\_

4. In paper chromatography the ………………………. phase is the paper and the mobile phase is the …………………………..

5. Gas-liquid chromatography (GLC) involves a sample being vaporised and injected into a chromatographic column. The sample is transported through the column by the flow of inert gas – the …………………………….phase. The column is made of a liquid soaked into a solid – the…………………………………… phase.

6. A chromatogram obtained from GLC shows the quantity of substance on the y-axis and the **…………………………** time on the x-axis. The ……………………… time is how long it takes the sample to pass through the column. Different substances have …………………. ………………… times which can be used to identify the compounds present.



7. Volumetric analysis involves using a solution of accurately known concentration in aquantitative reaction to determine the …………………………. of another substance.

8. A solution of accurately known concentration is known as a **……………………..** solution.

9. ……………………………. can be used to measure the volume of reactant solution required to complete the reaction. The end-point is the point at which the reaction is just complete. ……………………….. can be used to detect the end-point of a reaction.

10. Sometimes you do not need an …………..……….. as the reaction is self-indicating, i.e. the product is a different colour from the ……………………………..

e.g. permanganate ions, ……………., are a ………..………. colour and manganese ions, …………, are ……………………………

11. The relationship C1V1 = C2V2 can be used to work out the concentrations of reactants in

n1 n2

acid - ………………… titrations and ……………………… titrations.

Example:

25 cm3 of a solution of sodium hydroxide was added to a flask and titrated with a 0.2 mol l-1 solution of hydrochloric acid.

***HCl + NaOH → NaCl + H2O***

The experiment was carried out three times and the volumes of HCl titrated in each

experiment are shown in the table.

|  |  |
| --- | --- |
| Titration | Volume of 0.2 mol l-1 solution of HCl (cm3) |
| 1 | 11.3 |
| 2 | 10.4 |
| 3 | 10.6 |

Calculate the concentration of the NaOH solution, in mol l –1.