**Chemistry**

**The Summer Bridging Work MUST be handed into one of your subject teachers no later than Friday 13 September 2019.**

**Your work will be assessed in September by your class teachers.**

**Anyone not completing the work or producing work of poor quality will be re-interviewed regarding their place on the course and in the Sixth Form.**

**The aims are for you to understand if you like the course and for you to be ready to start learning at post-16 level.**

**Please hand your work in no later than Friday 13 September 2019.**

**­­­­­­­­­**

**Things you will need to succeed every day in the Sixth Form:**

* Pens
* Highlighters
* A pencil case
* Your own lined paper
* A single-hole punch (available from the school shop for £1)
* ****A pair of scissors
* ****Glue

**Things you will need for this course:**

* A lever-arch folder for storing work at home
* A ring-binder for work for the current unit
* A pack of at least 20 file dividers
* A Scientific calculator
* A ruler
* Summer Bridging Work Chemistry

**The books you need to buy are:**

There are a number of textbooks available for this course. Students in the past have found *A-Level Chemistry: OCR A Year 1 & 2 Complete Revision & Practice with Online Edition* a good option.

**Your Summer Bridging Work Project: Please read the information below and complete the questions as specified.**

**COURSE OUTLINE**

**The syllabus we follow is OCR Chemistry A**

A copy of the specification can be obtained from the OCR web site

**Autumn Term 2019**

* **Atoms and Electron Structure**
* **Moles, Equations and Acids**
* **Structure and Bonding**
* **Redox, Group 2 and Group 7**

**Spring and Summer Term 2020**

* **Basic Concepts in Organic Chemistry**
* **Alkanes and Alkenes**
* **Haloalkanes and Alcohols**
* **Enthalpy Changes**
* **Rates and Equilibrium**
* **Modern Analytical Techniques**
* **Chemistry of the Air**
* **Resources and Green Chemistry**

**Other important aspects of the course**

**Practical work**

Practical work is an essential element of the A Level Chemistry course. It is assessed during the course by your teachers who will track your skills as you improve. It is important that during the course you develop practical skills that allow you to complete experiments accurately, efficiently and safely. Some of the skills you have already mastered, others will be new to you. Along with an A-Level grade, your teachers will award you a pass or fail for your practical skills, this is called the practical endorsement.

The purpose of the practical work is to:

* illustrate the theory work
* to develop the practical skills needed for further study of Chemistry
* develop investigative and planning skills.

**Independent Learning**

When you study A Level Chemistry at least one section one the course will involve independent learning. You will be provided with a textbook and guidance on the necessary criteria, and your teacher will be one of your resources. It is a challenging opportunity for you to develop your skills of organisation. You will of course be expected to read about the topics taught and to make your own notes and supplement those given by your teacher. You should expect to show a great deal of independence in terms of the way that you tackle the workload presented.

**Working out of the classroom**

You will need to do about 5 hours work out of the classroom per week. If homework is not specifically set there is always background reading or revision to do. It is much easier to revise and sort your problems out straight away than find out in an examination you cannot do it. 5 hours does mean 5 solid hours with no interruptions - it is amazing how much work you can get done in that time.

Homework will be regularly set. Hand in your work on time since it allows your teacher the time to mark your work and comment constructively about your efforts. Failure to hand work in at the right time may result in it not being marked. When you get your work back check and if necessary seek clarification about your errors.

Finally remember that you can always see your teacher for help before you hand in your work in.

**Revision**

Success in A Level Chemistry will depend on how much revision you do. Students who complete all the past papers set and use a revision plan are much more likely to get the grade of which they are capable than those students who do not complete the past papers and revise effectively.

**PART 1: MEASURING AMOUNT OF SUBSTANCE**

**MASS VOLUME MOLAR MASS AVOGADRO**



**CONCENTRATION ATOM ION MOLECULE**

**MEASUREMENTS IN CHEMISTRY**

**Mass**

Convert the following into grams:

1. 0.25 kg

Helpful conversions

1 tonne = 1000kg

1 dm3 = 1000 cm3

1. 15 kg
2. 100 tonnes
3. 2 tonnes

**Volume**

Convert the following into dm3:

1. 100 cm3
2. 25 cm3
3. 50 m3
4. 50000 cm3

Tip – always use standard form for very large and very small numbers!

**Moles**

What is a mole? (GCSE)

Atoms and molecules are very small – far too small to count individually!

It is important to know how much of something we have, but we count particles in MOLES because you get simpler numbers

1 mole = 6.02 x 1023 particles

(6.02 x 1023 is known as Avogadro’s number)

1. If you have 2.5 x 1021 atoms of magnesium, how many moles do you have?
2. If you have 0.25 moles of carbon dioxide, how many molecules do you have?

How can you work out how many moles you have?

1. From a measurement of **MASS**: (GCSE)

You can find the number of moles of a substance if you are given its **mass** and you know its **molar mass**:

 **number of moles = mass (g)/molar mass (gmol-1)**

 **n = m/mr**

**Mass must be measured in grams. Molar mass has units of gmol-1**

|  |  |  |
| --- | --- | --- |
| 1. Calculate the number of moles present in: | 2. Calculate the mass of: | 3. Calculate the molar mass of the following substances: |
| a) 2.3 g of Na | a) 0.05 moles of Cl2 | a) 0.015 moles, 0.42 g |
| b) 2.5 g of O2 | b) 0.125 moles of KBr | b) 0.0125 moles, 0.50 g |
| c) 240000 g of CO2  | c) 0.075 moles of Ca(OH)2 | c) 0.55 moles, 88 g |
| d) 12.5 g of Al(OH)3 | d) 250 moles of Fe2O3 | d) 2.25 moles, 63 g |
| e) 5.2 g of PbO2 | e) 0.02 moles of Al2(SO4)3 | e) 0.00125 moles, 0.312 g |

1. From a measurement of AQUEOUS VOLUME: (New at A Level)

You can find the number of moles of a substance dissolved in water (aqueous) if you are given the **volume** of solution and you know its **molar concentration**:

**number of moles = aqueous volume x molar concentration**

 **n = V x C**

**Aqueous volume must be measured in dm3! Concentration has units of moldm-3.**

**What is a dm? What is a dm2? What is a dm3?**

**If you know the molar mass of the substance, you can convert the molar concentration into a mass concentration:**

**Molar concentration (moldm-3) x mr = mass concentration (gdm-3)**

|  |  |  |
| --- | --- | --- |
| 1. Calculate the number of moles of substance present in each of the following solutions: | 2. Calculate the molar concentration and the mass concentration of the following solutions: | 3. Calculate the molar concentration and the mass concentration of the following solutions: |
| a) 25 cm3 of 0.1 moldm-3 HCl | a) 0.05 moles of HCl in 20 cm3 | a) 35 g of NaCl in 100 cm3 |
| b) 40 cm3 of 0.2 moldm-3 HNO3 | b) 0.01 moles of NaOH in 25 cm3 | b) 20 g of CuSO4 in 200 cm3 |
| c) 10 cm3 of 1.5 moldm-3 NaCl | c) 0.002 moles of H2SO4 in 16.5 cm3 | c) 5 g of HCl in 50 cm3 |
| d) 5 cm3 of 0.5 moldm-3 AgNO3 | d) 0.02 moles of CuSO4 in 200 cm3 | d) 8 g of NaOH in 250 cm3 |
| e) 50 cm3 of 0.1 moldm-3 H2SO4 | e) 0.1 moles of NH3 in 50 cm3 | e) 2.5 g of NH3 in 50 cm3 |

1. From a measurement of GASEOUS VOLUME: (New at A Level)

You can find the number of moles of a gas if you are given the **volume** of the gas:

**number of moles = volume / 24**

 **n = V / 24**

**24 dm3 is the volume occupied by 1 mole of any gas at room temperature and pressure**

**Volume MUST be measured in dm3!**

**Hint: Check that all the volumes are in dm3 BEFORE carrying out your calculation.**

|  |  |  |
| --- | --- | --- |
| 1. Calculate the number of moles present in: | 2. Calculate the volume of gas occupied by: | 3. Calculate the mass of the following gas samples: |
| a) 48 dm3 of O2 | a) 0.05 moles of Cl2 | a) 48 dm3 of O2 |
| b) 1.2 dm3 of CO2 | b) 0.25 moles of CO2 | b) 1.2 dm3 of CO2 |
| c) 200 cm3 of N2 | c) 28 g of N2 | c) 200 cm3 of N2 |
| d) 100 dm3 of Cl2 | d) 3.2 g of O2 | d) 100 dm3 of Cl2 |
| e) 60 cm3 of NO2 | e) 20 g of NO2 | e) 60 cm3 of NO2 |

**PART 2: USING CHEMICAL EQUATIONS**

**How many moles?** (GCSE)

1. John weighs a sample of CaCO3 and records a mass of 5.0 g. How many moles of calcium carbonate are present?
2. Fatima measures out 50 cm3 of 0.1 moldm-3 hydrochloric acid. How many moles of hydrochloric acid are present?
3. Hussain collects 48 cm3 of carbon dioxide in a gas syringe. How many moles of carbon dioxide are present?

Using Chemical Equations

Chemical Equations show the ratio in which different species react in a chemical equation.



This equation shows that **6** moles carbon dioxide of react with **6** mole of water to make **1** mole of glucose and **6** moles of oxygen.

6: 6: 1: 6

1. How many moles of water are needed to react with 0.03 moles of carbon dioxide?
2. How many moles of glucose can you make from 0.03 moles of carbon dioxide?
3. How many moles of oxygen can you make from 0.03 moles of carbon dioxide?

**Equation 1:** Mg + 2 HCl → MgCl2 + H2

1. How many moles of magnesium would be needed to react with 0.01 moles of hydrochloric acid?
2. How many moles of hydrogen could be produced from 0.01 moles of hydrochloric acid?

Equation 2: 2 H2S + 3 O2  2 SO2 + 2 H2O

1. How many moles of oxygen is needed to react with 0.5 moles of hydrogen sulphide?
2. How many moles of sulphur dioxide can be made from 0.5 moles of hydrogen sulphide?

Equation 3: 4 K + O2  2 K2O

1. How many moles of oxygen are needed to react with 0.05 moles of potassium?
2. How many moles of potassium oxide can be made from 0.05 moles of potassium?

Calculating Reacting Quantities from Chemical Equations

*Use your notes from the GCSE unit CH2.3 “How much” to help you with these.*

**You perform these calculations in three steps:**

**-**  calculate the number of moles of one of the substances (you will either be given the mass, or the aqueous volume and the concentration, or the gaseous volume)

- use the equation to work out the number of moles of the other substance

- use one of the mole relationships to work out the quantity you need

1) What mass of hydrogen is produced when 192 g of magnesium is reacted with hydrochloric acid?

 Mg + 2 HCl → MgCl2 + H2 (3)

2) What mass of oxygen is needed to react with 8.5 g of hydrogen sulphide (H2S)?

 2 H2S + 3 O2  2 SO2 + 2 H2O (3)

3) What mass of potassium oxide is formed when 7.8 g of potassium is burned in oxygen?

 4 K + O2  2 K2O (3)

4) What mass of oxygen is required to oxidise 10 g of ammonia to NO?

 4 NH3 + 5 O2 → 4 NO + 6 H2O (3)

5) What mass of aluminium oxide is produced when 135 g of aluminium is burned in oxygen?

 2 Al + 3 O2  Al2O3 (3)

6) What mass of iodine is produced when 7.1 g of chlorine reacts with excess potassium iodide?

 Cl2 + 2 KI  2 KCl + I2 (3)

7) What volume of hydrogen is needed to react with 32 g of copper oxide?

 CuO + H2 → Cu + H2O (3)

8) What volume of oxygen is formed when 735 g of potassium chlorate decomposes?

 2 KClO3 → 2 KCl + 3 O2 (3)

9) What volume of hydrogen is produced when 195 g of potassium is added to water?

 2 K + 2 H2O  2 KOH + H2 (3)

10) What mass of calcium carbonate is required to produce 1.2 dm3 of carbon dioxide?

 CaCO3  CaO + CO2 (3)

11) What mass of magnesium oxide is formed when magnesium reacts with 6 dm3 of oxygen?

 2 Mg + O2  2 MgO (3)

12) What volume of carbon dioxide is produced when 5.6 g of butene (C4H8) is burnt?

 C4H8 + 6 O2  4 CO2 + 4 H2O (3)

13) The pollutant sulphur dioxide can be removed from the air by reaction with calcium carbonate in the presence of oxygen. What mass of calcium carbonate is needed to remove 480 dm3 of sulphur dioxide?

 2 CaCO3 + 2 SO2 + O2  2 CaSO4 + 2 CO2 (3)

14) 25 cm3 of a solution of sodium hydroxide reacts with 15 cm3 of 0.1 mol/dm3 HCl. What is the molar concentration of the sodium hydroxide solution?

HCl + NaOH 🡪 NaCl + H2O (3)

**Using state symbols**

1. What are the 4 state symbols?
2. In your exams you will be expected to include state symbols in all of your equations. There are marks available for them. Select 3 equations from the calculations in the previous section and write them out, in full, including the correct state symbol for each substance in the equation.

**PART 3: STRUCTURE AND BONDING**

**The following notes should be familiar, with some extension to your GCSE knowledge. This content is a good introduction to the “structure and bonding” part of unit 1. On your return to school you will be asked to complete a practice exam question using the content outlined in part 3 of this booklet.**

**TYPES OF BOND**

Atoms bond to each other in one of four ways:

1. Ionic bonding

**An ionic bond is a electrostatic attraction between oppositely charged ions, which are formed by the transfer of electrons from one atom to another**.

E.g. In sodium chloride, each sodium atom transfers an electron to a chlorine atom. The result is a sodium ion and a chloride anion. These two ions attract each other to form a stable compound.



1. Covalent bonding

**A covalent bond is a pair of electrons shared between two atoms**.

In a normal covalent bond, each atom provides one of the electrons in the bond. A covalent bond is represented by a short straight line between the two atoms.

E.g. water



Covalent bonds should not be regarded as shared electron pairs in a fixed position; the electrons are in a state of constant motion and are best regarded more as **charge clouds**.

Task: Practice your dot and cross diagrams.

On the following page, draw dot and cross diagrams for the following compounds:

1. NaCl

2. MgCl2

Remember- ionic bonds are between a metal and a non-metal, covalent bonds between non-metals only.

3. Na2O

4. MgO

5. Cl2

6. NH3

7. CH4

8. AlCl3

9. CO2

10. H2O

iii) Metallic bonding

**A metallic bond is an attraction between positive ions and a sea of electrons**.

Metallic bonds are formed when atoms lose electrons and the resulting electrons are attracted to all the resulting positive ions.

E.g. Magnesium atoms lose two electrons each, and the resulting electrons are attracted to all the positive ions.



Metallic bonding happens because the electrons are attracted to more than one nucleus and hence more stable. The electrons are said to be delocalized – they are not attached to any particular atom but are free to move between the atoms.

**STRUCTURES AND PROPERTIES**

# Bonding in ionic compounds

An ionic bond is electrostatic attraction between oppositely charged ions. After the ions are formed they all come together to form a **lattice**. A lattice is an infinite and repeating arrangement of particles. All the anions are surrounded by cations and all the cations are surrounded by anions.

**Example – sodium chloride**

The diagram below shows the structure of sodium chloride. The pattern repeats in this way and the structure extends (repeats itself) in all directions over countless ions. You must remember that this diagram represents only a tiny part of the whole sodium chloride crystal.

|  |  |
| --- | --- |
| naclexpl | naclspfilEach sodium ion attracts several chloride ions and vice versa so the ionic bonding is not just between one sodium and one chloride ion. There is a 3-D lattice. |

1. Melting and boiling point

The attraction between opposite ions is very strong. A lot of kinetic energy is thus required to overcome them and the melting point and boiling point of ionic compounds is very high.

In the liquid or aqueous state, the ions still retain their charge and the attraction between the ions is still strong. Much more energy is required to separate the ions completely and the difference between the melting and boiling point is thus large.

|  |  |  |
| --- | --- | --- |
| Compound | NaCl | MgO |
| Melting point/oC | 801 | 2852 |
| Boiling point/oC | 1459 | 3600 |

The higher the charge on the ions, and the smaller they are, the stronger the electrostatic attraction between them will be and the higher the melting and boiling points. In MgO, the ions have a 2+ and 2- charge and thus the attraction between them is stronger than in NaCl, so the melting and boiling points are higher.

1. Electrical Conductivity

Since ionic solids contain ions, they are attracted by electric fields and will, if possible, move towards the electrodes and thus conduct electricity. In the solid state, however, the ions are not free to move since they are tightly held in place by each other. Thus ionic compounds do not conduct electricity in the solid state. Ionic solids are thus good insulators.

In the liquid or aqueous state, the ions are free to move and so can move towards their respective electrodes, thus ionic compounds can conduct electricity.

1. Mechanical properties

Since ions are held strongly in place by the other ions, they cannot move or slip over each other easily and are hence hard and brittle.



opposite ions attract like ions repel – structure breaks

# Bonding in metals

Metallic bonding is the electrostatic attraction between cations and a sea of delocalised electrons. The cations are arranged to form a lattice, with the electrons free to move between them.

The structure of the lattice varies from metal to metal, and they do not need to be known in detail. It is possible to draw a simplified form of the lattice:

**Example - magnesium**



This is a simplified 2D form of the metal lattice

Properties of metals

1. Electrical conductivity

Since the electrons in a metal are delocalised, they are free to move throughout the crystal in a certain direction when a potential difference is applied and metals can thus conduct electricity in the solid state. The delocalised electrons are still present in the liquid state, so metals can also conduct electricity in the liquid state.

1. Melting and boiling point

 Although not generally as strong as in ionic compounds, the bonding in metals is relatively strong, and as a result the melting and boiling points of metals are relatively high.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Metal | Na | K | Be | Mg |
| Melting point/ oC | 98 | 64 | 1278 | 649 |
| Boiling point/ oC | 883 | 760 | 2970 | 1107 |

Smaller ions, and those with a high charge, attract the electrons more strongly and so have higher melting points than larger ions with a low charge. Na has smaller cations than K so has a higher melting and boiling point. Mg cations have a higher charge than Na so has a higher melting and boiling point.

1. Other physical properties

 Since the bonding in metals is non-directional, it does not really matter how the cations are oriented relative to each other. The metal cations can be moved around and there will still be delocalized electrons available to hold the cations together. The metal cations can thus slip over each other fairly easily. As a result, metals tend to be soft, malleable and ductile.

# Bonding in covalent structures

Covalent bonding can result in three very different types of substance:

1. **Molecular**

In many cases, the bonding capacity is reached after only a few atoms have combined with each other to form a molecule. If no more covalent bonds can be formed after this, the substance will be made up of a larger number of discreet units (molecules) with no strong bonding between them.

Such substances are called **molecular substances**, and there are many examples of them:

CH4, Cl2, He, S8, P4, O2, H2O, NH3

The molecules are held together by **intermolecular forces**, which are much weaker than covalent bonds but are often strong enough to keep the substance in the solid or liquid state.

**Example - Iodine**



There are attractive forces between these molecules, known as intermolecular forces, but they are weak. In the gaseous state, the intermolecular forces are broken but the bonds within the molecule remain intact - they are not broken. The gas phase consists of molecules, not atoms.

Molecular substances have certain characteristic properties:

1. Melting and boiling point

 These are generally low, since intermolecular forces are weak.

Intermolecular forces also decrease rapidly with increasing distance, so there is often little difference in the melting and boiling points.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Substance | CH4 | H2O | H2 | He |
| Melting point /oC | -184 | 0 | -259 | -272 |
| Boiling point /oC | -166 | 100 | -253 | -268 |

1. Electrical conductivity

There are no ions and no delocalised electrons, so there is little electrical conductivity in either solid or liquid state.

1. Other physical properties

 The intermolecular forces are weak and generally non-directional, so most molecular covalent substances are soft, crumbly and not very strong.

1. **Giant covalent**

In some cases, it is not possible to satisfy the bonding capacity of a substance in the form of a molecule; the bonds between atoms continue indefinitely, and a large lattice is formed. There are no discrete molecules and covalent bonding exists between all adjacent atoms.

Such substances are called giant covalent substances, and the most important examples are C, B, Si and SiO2.

**Example – diamond** (diamond is an allotrope of carbon)

|  |  |
| --- | --- |
| diamond | Don't forget that this is just a tiny part of a giant structure extending on all 3 dimensions.drawdiamond |

In giant covalent compounds, covalent bonds must be broken before a substance can melt or boil.

Giant covalent compounds have certain characteristic properties:

1. Melting and boiling point

These are generally very high, since strong covalent bonds must be broken before any atoms can be separated. The melting and boiling points depend on the number of bonds formed by each atom and the bond strength. The difference between melting and boiling points is not usually very large, since covalent bonds are very directional and once broken, are broken completely.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Substance | C | Si | B | SiO2 |
| Melting point /oC | 3550 | 1410 | 2300 | 1510 |
| Boiling point /oC | 4827 | 2355 | 2550 | 2230 |

1. Electrical conductivity

There are no ions or delocalised electrons, so there is little electrical conductivity in either solid or liquid state.

1. Other physical properties

Since the covalent bonds are strong and directional, giant covalent substances are hard, strong and brittle.

Diamond is in fact the hardest substance known to man. For this reason it is used in drills, glass-cutting and styluses for turntables.

1. **Giant covalent layered**

Some substances contain an infinite lattice of covalently bonded atoms in two dimensions only to form layers. The different layers are held together by intermolecular forces, and there are often delocalized electrons in between the layers. Examples of these structures are graphite and black phosphorus.

**Example - graphite**

 or 

In graphite, each carbon atom is bonded to three others. The spare electron is delocalized and occupies the space in between the layers. All atoms in the same layer are held together by strong covalent bonds, and the different layers are held together by intermolecular forces.

A number of characteristic properties of graphite result from this structure:

1. Electrical conductivity

 Due to the delocalised electrons in each plane, graphite is a very good conductor of electricity in the x and y directions, even in the solid state (unusually for a non-metal). However, since the delocalisation is only in two dimensions, there is little electrical conductivity in the z direction (i.e. perpendicular to the planes).

1. Density:

Graphite has a much lower density than diamond (2.25 gcm-3) due to the relatively large distances in between the planes.

1. Hardness

Graphite is much softer than diamond since the different planes can slip over each other fairly easily. This results in the widespread use of graphite in pencils and as an industrial lubricant.

**SUMMARY OF DIFFERENT TYPES OF COMPOUND AND THEIR PROPERTIES**

**Task:** Complete the table to summarize the bonding and properties of each type of substance. The first row has been done for you. The numbers indicate the quantity of responses required in each box.

|  |  |  |
| --- | --- | --- |
| **SUBSTANCE** | Nature of bonding | Physical properties |
| **IONIC****E.g. NaCl** | Electrostatic ttraction between oppositely charged ions. Infinite lattice of oppositely charged ions in three dimensions | High mpt, bptGood conductors in liquid statePoor conductors in solid stateHard, strong, brittle. |
| **METALLIC****E.g.**  | 1) 2) | 1)2)3)4) |
| **GIANT COVALENT****E.g.**  | 1)2) | 1)2)3)4) |
| **MOLECULAR****E.g.**  | 1)2) | 1)2)3)4) |
| **GIANT COVALENT LAYERED****E.g.**  | 1)2)3) | 1)2)3)4) |

Structure and bonding summary exercises

**Task 1:** Predict which of the following pairs of substances is likely to have the higher melting point, giving reasons for your choice:

1. Na and Mg
2. Na and K
3. NaCl and NaBr
4. NaCl and MgO
5. C and Si
6. Ne and Ar
7. F2 and Cl2
8. NH3 and PH3
9. NaCl and HCl
10. SiO2 and CO2

**Task 2:** Describe the bonding in the following compounds, and briefly describe their main physical properties (limited to melting point and electrical conductivity).

1. magnesium
2. diamond
3. silicon dioxide
4. magnesium oxide
5. carbon dioxide
6. graphite
7. sodium nitrate
8. water
9. sulphur dioxide
10. helium

**Well done for completing this transition course!**

If you have any questions or concerns about A Level Chemistry please contact Dr Balster at rbl@cheney.oxon.sch.uk, Miss Rayment era@cheney.oxon.sch.uk or Ms Elias at lel@cheney.oxon.sch.uk.