CHSG Sixth Form
A-level Chemistry
Bridging Work Booklet

Name: ______________________________

SECTION A: INTRODUCTION

Welcome to Carshalton High School for Girls Sixth Form A-level Chemistry!

This bridging work is designed to help you bridge the gap between your GCSE Science studies and the AS/A Level Chemistry course. It includes a list of topics from GCSE that will be helpful for you to review and practice. You may find it easy, not-so-easy, tricky or really tricky. There may be some questions you can't do at all. The aim is for you to practice your Chemistry and identify your strengths and weaknesses in the subject.

Why do bridging work? Because we want you to be successful and what it takes to be successful at GCSE is different from being successful at A-level. Although you have fewer subjects there are different skills at post 16 and the volume of work is greater because the detail and depth is more demanding. Bridging work should help you gauge whether the subject is for you, so you can change your mind within the first 2 weeks – as long as there is space and you meet the entry criteria. We would rather you study courses that interest you and you are sufficiently qualified to study. If you would like further work, or an insight into the wonderful world of AS-Level Chemistry and beyond, there are some further reading suggestions towards the end. There is also a list of websites you will undoubtedly find useful throughout the course and may need to use to complete this task.

The booklet is subject based and will build on your chemistry knowledge. Is the bridging work assessed? Yes. In September, your subject teacher will ask you for your bridging work and it will be assessed **for effort and completion**. Teachers can diagnose your strengths and weaknesses and begin to support you in a more targeted way. Bridging work also assesses your work ethic and so the sixth form team will pick up on anyone with a low work ethic and support you accordingly. This leads into the fact should you decide to change, you would need to complete the bridging work for the new course you choose.

Chemistry A-level Studying Chemistry at A-level will require you to be highly organised and effective with your own independent work. Not only will you have to balance the workload of this subject and the other subjects you have chosen, we require you to commit and do the very best that you can. Anyone not completing the work or producing poor quality will be spoken to and asked to reconsider if this is the correct course for you. Please use resources such as the internet, library and your Chemistry GCSE notes to help you complete this booklet.

You are also expected to spend at least 3-5 hours a week on your Chemistry work outside of lessons. This will include homework tasks, pre-reading/flip learning, independent study tasks, making additional notes, reviewing lesson materials and reading around the subject. You should aim to purchase a textbook for your course – OCR A A-level Chemistry. Your teachers are also, of course, an excellent source of support both in and out of lessons. Other support includes, drop-in support classes outside of school hours, extended interventions and 1-1s. A full copy of the specification, past papers etc. can be accessed through the OCR website: http://www.ocr.org.uk/qualifications/as-a-level-gce-chemistry-a-h032-h432-from2015/

To complete this module of work, answer the questions provided in this workbook by using your GCSE knowledge. If you cannot complete these you will need to go over these topics before starting to ensure a solid foundational knowledge.

Key areas from your GCSE Science work that you will need for AS/A Level Chemistry

1) Atomic structure – protons, neutrons, electrons, mass number, isotopes etc.

2) Electron arrangement – how many electrons each shell can hold etc.

3) Ionic compounds – dot and cross diagrams, properties, examples.

4) Covalent compounds – dot and cross diagrams, properties, examples, diamond vs graphite.

5) Metallic bonding – diagram, properties of metals.

6) Quantitative Chemistry Calculations – relative atomic mass, relative molecular mass, atom economy, percentage yield, conversions.

7) Organic compounds – alkanes and alkenes, alcohols.

8) Rates of reaction – collision theory, how to speed up reactions, catalysts etc.

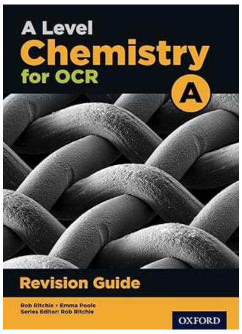
9) Endothermic and exothermic reactions.

10) Periodic table – overall arrangement in groups and periods.

11) Practical techniques, experiments you have studied and carried out.

Useful Resources:

1. OCR A-level Chemistry Text Book



2. Useful websites:
   a. Chemguide - www.chemguide.co.uk
   b. Knockhardy - www.knockhardy.org.uk/sci.htm
   c. BBC Bitesize - www.bbc.co.uk/schools/cgsebitesize/chemistry
   d. Royal Society of Chemistry - www.rsc.co.uk
   e. Khan Academy - www.khanacademy.org

SECTION B: WHY STUDY A-LEVEL CHEMISTRY?

A-level Chemistry provides a gateway into so many university degrees and career pathways. University degrees such as chemistry, chemical engineering, medicine, environmental science and environmental engineering, materials and mechanical engineering, health and health sciences, agriculture and biochemistry are some of the obvious ones. These degrees can lead to careers in a plethora of jobs including engineers, doctors, chemist, pharmacists, environmental scientists, geologists, forensic science and robotics to name a few.

Use the questions below to understand more about why you would like to study chemistry!

1. What is chemistry? ____

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2. Why is the study of chemistry been very useful in the past? What problems has it helped to solve?

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3. How may chemistry be useful going forward in the future? What problems that exist in our world may it help to solve and that you may want to be involved in as part of the solution?

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4. Name three scientists from before the 1900's that contributed great work in the field of chemistry and describe what their contribution was.

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SECTION C: GCSE REFRESHER (Questions should be completed on lined paper.)

PART 1 BALANCING EQUATIONS

It's a key skill in chemistry. You must be able to do it. Have a go and if you are struggling, get it sorted. Balance the following equations:

1) $Mg(s) + O_2(g) \rightarrow MgO(s)$

2)$H_2(g) + 0_2(g) \rightarrow H2O(l)$

3)$Fe(s) + HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

4)$CuO(s) + HNO3(aq) \rightarrow Cu(NO_3)_2\ (aq) + H_2O(l)$

5)$Ca(OH)_2(aq) + HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$

6)$KHCO_3(s) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + CO_2(g) + H_2O(l)$

7)$Al(s) + Cl_2(g) \rightarrow AlCl_3(s)$

Part 2 BONDING

A. IONIC BONDING

Table salt (sodium chloride, NaCl) is our most common ionic compound. It is also an excellent exemplar of how ionic substances behave. Under a microscope, or even on your kitchen table, you can see the beautiful crystalline lattice structure. Whilst it adds flavour to our food it doesn't melt when added to hot fish and chips. However, it dissolves readily in water, providing an ideal habitat for crocodiles and other marine organisms which rely on a salty aqueous environment. Brine conducts electricity and the products of its electrolysis provide us with vital chemical ingredients for our everyday life.

1) Complete the passage below using the following words: loses ions ionic protons negative electrons positive gains Atoms are neutral because they have the same number of ………………….. and ………………… . If atoms lose or gain electrons they become electrically charged and are called ………… (they are not atoms any more). If atoms gain electrons they become ………………... ions, and if they lose electrons they become ………………… ions. When a metal reacts with a non-metal, the metal atoms …………….. electrons and the non-metal atoms …………… electrons, forming an ……………….. compound.

2) Describe the structure of sodium chloride.

3) a) Explain why ionic substances have high melting and boiling points. b) Explain why ionic substances can conduct electricity when molten or dissolved. c) Explain why ionic substances cannot conduct electricity when solid.

4) Name the three products from the electrolysis of brine and give one example of how each is useful to us in everyday life. Product Use

5) Deduce the chemical formulae of the following ionic compounds:- a) calcium chloride d) aluminium hydroxide b) sodium oxide e) potassium carbonate c) magnesium sulfide f) calcium nitrate

B. COVALENT BONDING

Covalently bonded molecules are everywhere! In fact, you are breathing some in (and out) as you read this. Their simple molecular structure is crucial to your survival. When you use your pencil to answer these questions you are relying on the properties of one of the World's most useful giant covalent structures, graphite. At the Brit Awards, Adele and other starlets adorn themselves with the World's strongest naturally occurring covalent structure, diamond. Which, as it just so happens, was also instrumental in the Hatten Garden robberies as a consequence of this very property! Simple covalent molecules

1) Circle the correct answer.

Covalent bonding occurs between: Metal - Non-metal ; Metal – Metal ; Non-metal - Non-metal

2) How does a covalent bond form

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3) What are the properties of simple covalent substances such as chlorine or oxygen?

Melting point and boiling point: High/Low

Solubility in water: Soluble/Insoluble

Conducts electricity: Conductors/Insulators

Bonding between molecules (intermolecular bonding): Weak/strong

4) Draw dot-and-cross diagrams of the following simple molecules:

a. Methane

b. Water

5) Describe and explain the difference in the boiling point of water compared to chlorine and oxygen.

C. GIANT COVALENT STRUCTURES

1. What are the properties of giant covalent substances?

Types of covalent structures – name four:

Properties - High or low bp and mp:

Conductor or insulator:

Hard or soft:

Solubility in H_2O:

Uses:

2. Giant covalent structures tend to have low melting and boiling points. True/false

3. Most intermolecular forces are strong and make it difficult to separate the molecules. True/false

4. Most covalent substances do not conduct electricity. True/false

5. Graphite conducts electricity. True/false

6. Graphite is slippery because the intramolecular bonds are weak covalent bonds. True/false

Now explain your answer to each of the above statements.

PART 3 - ACIDS AND ALKALIS

Acids and alkalis play a crucial part in our everyday lives. Indigestion is caused by excess stomach acid. Gaviscon contains an alkali to neutralise the excess acid. Our breathing is controlled by the pH of our blood. Bee stings hurt thanks to formic acid. The effects can be neutralised by bicarbonate of soda. Chemists often carry out titrations to determine unknown concentrations of acids or alkali, particularly when quality checking products. A good example is checking the concentration of alkali in fertilisers before they go on shop shelves for us to buy; too much alkali can be just as bad (if not worse) than too much acid (caused by acid rain).

1) Acids have a pH of ……………… than 7. Alkalis have a pH of …………. than 7. Neutral substances have a pH of ………….. .

2) Acid + Metal → ………………… + …………………….

Acid + Metal Oxide → …………………….. + ……………………..

Acid + Metal Hydroxide → ………………………. + ……………………….

Acid + Metal Carbonate → …………………………. + ………………………. + ………………………….

3) Mr Withers needs to know how acidic the soil is in the school grounds. He decides to ask the Chemistry A-Level students to find out by doing a titration. They decide to use sodium hydroxide as their alkali of known concentration.

a) Balance the equation for this reaction. $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

b) The chemistry students use 24.2 cm3 of sulfuric acid, extracted from the soil, to neutralise 25.0 cm3 of 0.010 moldm-3 sodium hydroxide. Determine the concentration of sulfuric acid in the school soil.

PART 4 - REDOX

Without redox we wouldn't be able to get energy from our food. On a slightly less essential level, batteries and hydrogen fuel cells rely on redox to switch on torches and power modern cars. The key rule to remember in redox is that "the electrons have got to go somewhere!"…more on that in lesson time.

1. What is "redox"?

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2. Give two examples of useful redox reactions in everday life excluding those mentioned above (there are millions!).

3. What does oxidation mean – give two definitions?

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4. What does reduction mean – give two definitions?

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5. Which element is oxidised and which is reduced in the reaction below? Explain your answer.

$$2Na + Cl_2 \rightarrow 2NaCl$$

Oxidised .... Reduced ....

PART 5 – QUANTITATIVE CHEMISTRY

Calculations are a part of every chemist's world. They are sometimes something that A Level students find tricky but you can do it! The key is to sort out anything you don't understand and get plenty of practice to improve your confidence. These calculations build up in difficulty to those found on AS Level papers. Give them a shot; you may be surprised by how much you can do.

1. Magnesium sulfate is one of the chemicals in detergent powder. Ana makes some magnesium sulfate using this reaction.

   magnesium carbonate + sulfuric acid → magnesium sulfate + water + carbon dioxide

$$MgCO_3 + H_2SO_4 \rightarrow MgSO_4 + H_2O + CO_2$$

a) The theoretical yield for Ana's experiment is 12.0 g. Ana dries and weighs the magnesium sulfate she makes. This is her actual yield. Actual yield = 10.8 g. Work out the **percentage yield** for Ana's experiment.

b) The relative formula mass of magnesium carbonate is 84. The relative formula mass of magnesium sulfate is 120. Calculate the mass of magnesium carbonate that must react with sulfuric acid to produce 12.0 g of magnesium sulfate.

2. Many elements have variable oxidation states. What does this mean and how is it useful to us?

3. The ore haematite contains iron(III) oxide. Iron is extracted from this ore by reduction with carbon. The products of this reaction are iron and carbon dioxide.

(a) Finish this symbol equation for the reaction. ………… Fe2O3 + ………… C → ………… + …………

(b) A haematite ore contains 80% by mass of iron(III) oxide. Calculate the maximum mass of iron that can be extracted from each tonne of this ore. Show each step of your calculation as indicated below.

HINTS: 1 tonne = 1000 kg

4. Titrations

a. A solution of hydrochloric acid (HCl) has a concentration of 1.50 mol per dm3. Calculate its concentration in grams per dm3. (3 marks)

b. 0.0350 dm3 of sodium hydroxide (NaOH) solution was put in a flask. The concentration of the sodium hydroxide was 0.500 mol/dm3. 0.0250dm3 of hydrochloric acid (HCl) was needed to neutralise it. Calculate the concentration of the hydrochloric acid in mol/dm3 (6)

c. 25.0 cm3 of sodium hydroxide (NaOH) solution was neutralised by 37.5 cm3 of hydrochloric acid (HCl). The concentration of the sodium hydroxide was 0.750 mol/dm3. Calculate the concentration of the hydrochloric acid in mol/dm3. (6)

SECTION D: QUESTIONS (Questions should be completed on lined paper if needed.)

1. (a) Define the term atomic number of an element. (1)

(b) Give the symbol, including mass number and atomic number, for an atom of an element which contains 12 neutrons and 11 electrons. (2)

(c) How many neutrons are there in one 27Al atom? (1)

(d) Define the term relative atomic mass of an element. (2)

2. At room temperature, both sodium metal and sodium chloride are crystalline solids which contain ions.

(a) Create diagrams for sodium metal lattice and a sodium chloride lattice below and mark the charge for each ion. (2)

(b) (i) Explain how the ions are held together in solid sodium metal.

(ii) Explain how the ions are held together in solid sodium chloride.

(iii)The melting point of sodium chloride is much higher than that of sodium metal. What can be deduced from this information? (3)

(a) Explain why sodium metal is malleable (can be hammered into shape). (1)

3. Sodium chlorate, NaClO3, contains 21.6% by mass of sodium, 33.3% by mass of chlorine and 45.1% by mass of oxygen.

(i) Use the above data to show that the empirical formula of sodium chlorate is NaClO3

(2) (ii) Sodium chlorate may be prepared by passing chlorine into hot aqueous sodium hydroxide.

(iv) Balance the equation for this reaction below.

$$\ldots\ldots\ Cl_2 + \ldots\ldots\ NaOH \rightarrow \ldots\ldots\ NaCl + NaClO_3 + 3H_2O\ (1)$$

4. (a) Give the relative mass and relative charge of a neutron.

Relative mass ....
Relative charge.. (2)

(b) In terms of the number of their fundamental particles, what do two isotopes of an element have in common and how do they differ?

In common ....
Difference .... (2)

(c) Give the complete atomic symbol, including mass number and atomic number, for an atom of the isotope with 22 neutrons and 19 electrons. (2)

5. (a) Describe the bonding in metals. (2)

(b) Explain why the melting point of magnesium is higher than that of sodium. (3)

(c) Explain how metals conduct electricity. (2)

6. (i) Showing the outer electrons only, draw a dot-and-cross diagram to indicate the bonding in calcium oxide (CaO). (2)

(ii) Describe the type and strength of the bonding in solid calcium oxide. (3)

(iii) Use ideas about solids, liquids and gases to describe the changes that take place as calcium oxide is heated from 25°C (room temperature) to a temperature above its melting point. (3)

(iv) State two properties of calcium oxide that depend on its bonding. (2)

7. (a) Give the relative charge and relative mass of an electron.

Relative charge...
Relative mass .... (2)

(b) Isotopes of chromium include 54Cr and 52Cr

(i) Give the number of protons present in an atom of 54Cr

(ii) Deduce the number of neutrons present in an atom of 52Cr (2)

(c) (i) State what is meant by the term empirical formula.

(ii) A chromium compound contains 28.4% of sodium and 32.1% of chromium by mass, the remainder being oxygen. Calculate the empirical formula of this compound. (4)

8. (a) Complete the following table: (b) An atom of element Q contains the same number of neutrons as are found in an atom of 27Al. An atom of Q also contains 14 protons.

(i) Give the number of protons in an atom of 27Al.

(ii)Deduce the symbol, including mass number and atomic number, for this atom of element Q. (3)

(c) Define the term relative atomic mass of an element. (2)

9. (a) (i) Describe the bonding in a metal (you may draw a diagram if it helps). (3)

ii) Explain why magnesium has a higher melting point than sodium. (4)

(b) Why do diamond and graphite both have high melting points? (3)

(c) Why is graphite a good conductor of (1)

(d) Why is graphite soft? (2)

10. Organic Chemistry

(a) Butane, C_4H_{10}, is a hydrocarbon which is used as a fuel. (i) Explain what is meant by the term hydrocarbon.

(ii) Write an equation for the complete combustion of butane.

(iii) Under what conditions would you expect incomplete combustion to occur? (3)

(b) Ethane (C_2H_6) can be cracked in the presence of a catalyst to produce ethene (C_2H_4) and hydrogen. (i) Write an equation for this reaction.

(ii) Give a suitable catalyst for this reaction.

(iii) State one reason why cracking is important. (3)

Alcohols

d. What is the common functional group to all alcohols? What do all their names end in? (2 marks)

e. Draw the structural formula for methanol and ethanol. (1 mark)

f. List the key properties of alcohols (3 marks) • • •

g. What are the main uses of alcohols? (1 mark)

h. What is the name of the alcohol found in alcoholic drinks? (1 mark)

Carboxylic Acids

i. What is the common functional group to all carboxylic acids? What do their names end in? (2 marks)

j. Draw the structural formula for ethanoic acid and propanoic acid. (1 mark)

k. Why do carboxylic acids make weakly acidic solutions with water? (1 mark)

l. What common compound do carboxylic acids react with are what are the products? (1 mark)

m. Which carboxylic acid is the main ingredient to vinegar, why is it an aqueous solution? (1 mark)

n. How can ethanol be made into ethanoic acid? (1 mark)

Esters

o. What two ingredients must be combined to make an ester? (1 mark)

p. Explain with the aid of a diagram, why ester formation is a condensation reaction (2 marks)

q. What is the functional group for an ester? (1 mark)

r. Name a common ester. (1 mark)

s. Describe some uses of esters, why are their properties suited to each? (3 marks)

11. Bond enthalpies

1) How can endothermic and exothermic reactions be explained in terms of breaking and making of chemical bonds? (1 mark)

2) Draw the energy level diagram for an endothermic reaction. Label the diagram with products, reactants, activation energy and energy absorbed. (3 marks)

3) The bond enthalpies for some common bonds are shown below. C-H: +413 kJ mol-1 , C-C: +347 kJ mol-1 , C-O: +358 kJ mol-1 , O=O: +497 kJ mol-1 , C=O: +805 kJ mol-1 , O-H: +463 kJ mol-1 , C=C: +612 kJ mol-1 , H-H: +436 kJ mol-1 , NΞN: +945 kJ mol-1 , N-H: +391 kJ mol-1

Calculate the enthalpy changes of reaction for each of the following reactions. (1 mark each)
a. $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$

b. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

12. Chemical analysis

A. Answer the following questions by identifying the anion and cation.

i. An unknown powered was dissolved in nitric acid and added to a solution of silver nitrate and produced a pale cream precipitate. The same powder, when burnt, gave a pale lilac flame. (1 mark)

ii. An unknown powder was dissolved in sodium hydroxide solution and then a little dilute ammonia solution was added, this formed a pale green precipitate. The powder was also dissolved in nitric acid and added to silver nitrate solution, this produced a white precipitate which dissolved in ammonia. (1 mark)

iii. A powder was dissolved in water and then a solution of barium chloride was added to this solution. This produced a white precipitate which did not dissolve in dilute hydrochloric acid. When the powder was burnt in a flame it produced a green/ yellow flame. (1 mark)

iv. A powder was dissolved in nitric acid and added to a solution of silver nitrate and produced a pale yellow precipitate formed. The same powder was burnt in a flame and produced a green colour. (1 mark)

Answer the following questions by stating which technique you would use to identify the compounds.

v. A chemist has made a new carbon and hydrogen based drug, which technique should he use to identify the structure of this drug? (1 mark)

vi. A water company wants to test a sample of their water for trace metals, which technique should they use? (1 mark)

vii. A chemist wants to test his proteins for their purity, which technique should she use? (1 mark)

viii. 4. A blood analyst wants to check a athletes blood for drugs, which technique should be used? (1 mark)

PART 6 - HANDLING NUMBERS

The ability to work with numbers is essential for Chemistry and the level of accuracy is very important. The numbers we use in Chemistry range from being extremely small to very large, and you must be able to deal with these.

Tasks

A. Converting to SI

a. Convert the following into SI units:

1. 67 cm 2. 30 minutes 3. 100 °c 4. - 27 °c 5. 0.1 g 6. 2.7 tonnes 7. 12 g carbon into moles

B. Decimal Places – dp

Your calculator can produce lots of digits after the decimal place, and you will need to record the answer accurately and appropriately to score marks in an exam. The answer will also need to be rounded up or down. Make sure you give the answer to the number of decimal places the exam question has asked for. If in doubt, 2 dp is the norm.

C. Significant Figures

SF Significant figures are useful when quoting numbers when decimal places are not appropriate. These numbers tell you about the magnitude of a figure. You will need to count the significant figure as soon as you come across a non-zero number reading from left to right. Examples to 3 SF: 3.81 0.0000381 3.81 3.00

D. Standard Form Some numbers are far too large to write out in full so a shorthand called 'standard form' or 'scientific notion' is used. Examples: 1.0 x 106 = 1, 000 000 1.0 x 106 = 0.00 0001