Introductory Day

A-Level Chemistry

Student Name……………………………

Summer Homework Submission Date: First Chemistry lesson in September

A-level Chemistry goes into much more detail than GCSE. It attempts to answer the big question 'what is the world made of' and it's the search for this answer that makes this subject so fascinating. From investigating how one substance can be changed drastically into another, to researching a new wonder drug to save millions of lives, the opportunities that chemistry provides are endless.

Possible degree options

According to bestcourse4me.com, the top five degree courses taken by students who have an A-level in Chemistry are:

• Chemistry
• Biology
• Pre-clinical medicine
• Mathematics
• Pharmacology.

Possible career options

Studying an A-level Chemistry related degree at university gives you all sorts of exciting career options, including:

• Analytical chemist
• Chemical engineer
• Clinical biochemist
• Pharmacologist
• Doctor
• Research scientist (physical sciences)
• Toxicologist
• Chartered certified accountant
• Environmental consultant
• Higher education lecturer
• Patent attorney
• Science writer
• Secondary school teacher.

Topics covered

A-level Chemistry lasts two years, with exams at the end of the second year.

First year of A-level

- **Physical chemistry** - Including atomic structure, amount of substance, bonding, energetics, kinetics, chemical equilibrium and Le Chatelier's principle

- **Inorganic chemistry** - Including periodicity, Group 2 the alkaline earth metals, Group 7(17) the halogens

- **Organic chemistry** - Including introduction to organic chemistry, alkanes, halogen alkanes, alkenes, alcohols, organic analysis

Second year of A-level

- **Physical chemistry** – Including thermodynamics, rate equations, the equilibrium constant, electrode potentials and electrochemical cells

- **Inorganic chemistry** - Including properties of Period 3 elements and their oxides, transition metals, reactions of ions in aqueous solution

- **Organic chemistry** - Including optical isomerism, aldehydes and ketones, carboxylic acids and derivatives, aromatic chemistry, amines, polymers, amino acids, proteins and DNA, organic synthesis, NMR spectroscopy, chromatography

Practicals

Chemistry, like all sciences, is a practical subject. Throughout the course you will carry out practical activities including:

- measuring energy changes in chemical reactions
- tests for identifying different types of compound
- different methods for measuring rates of reaction
- studying electrochemical cells
- preparation of organic solids and liquids
- an advanced form of chromatography for more accurate results

Exams

There is no coursework on this course. However, your performance during practicals will be assessed.

LEVEL	UNIT TITLE	ASSESSMENT DATE (PROVISIONAL)	MODE OF ASSESSMENT	Duration	A-level %
Year 1 progression	Progression exam	June, End of First year	Written Exam	1hr 30mins	0%
A-level	PAPER 1: Physical, Inorganic and Practical Skills (105 raw marks)	June, End of Second year	Written Exam	2hrs	35%
A-level	PAPER 2: Physical, Organic and Practical Skills (105 raw marks)	June, End of Second year	Written Exam	2hrs	35%
A-level	PAPER 3: Any content and any practical skills (90 raw marks)	June, End of Second year	Written Exam	2hrs	30%

Atomic Number, Mass Number and isotopes

- The atomic number of an element is given the symbol Z.
- It is sometimes called the proton number as it represents the number of protons in the nucleus of the element.
- For atoms the number of protons equals the number of electrons, but you need to take care when considering ions as the number of electrons changes when an ion forms from an atom.
- The mass number of an atom is given the symbol A. It represents the total number of neutrons and protons in the nucleus. Subtracting Z from A allows you to calculate the number of neutrons in the nucleus.
- Complete the following table (refer to the periodic table):

Element	Symbol	Z	A	No. protons	No. Neutrons	No. Electrons
Sodium			23			
		6	12			
		12			12	
		84	210			
Chlorine		17	35			
Chlorine		17	37			

Isotopes

The last two examples in the table above show two chlorine atoms with different numbers of neutrons. These are called isotopes of chlorine. Both are chlorine atoms because they have the same number of protons — but they have different numbers of neutrons. In other words they have the same atomic number but different mass numbers. Isotopes are very common: some occur naturally and some are man-made. Some elements may have a large number of isotopes.

Questions:

1. In terms of the numbers of subatomic particles, state one difference and two similarities between two isotopes of the same element.

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2. Give the chemical symbol, mass number and atomic number of an atom which has 3 electrons and 4 neutrons.

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3. Three isotopes of carbon are: carbon-12, carbon-13 and carbon-14. State the numbers of protons, neutrons and electrons in each.

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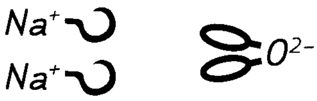
Deducing the Formulae of Ionic Compounds

The formula of a compound tells you the ratio of the elements that it contains. This ratio is fixed, and for ionic compounds that means it's easy to work out the formula from the charges on the ions.

Metal ions (and hydrogen ions) always carry a positive charge whilst non-metal ions carry a negative charge If you imagine that a positive charge is a 'hook' and a negative charge is an 'eye' then the formula can be deduced by exactly matching up the hooks and eyes. (This is to make the compound electrically neutral).

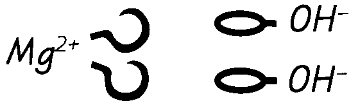
<table>
<tr>
<td>Example 1: What is the formula of sodium oxide?

Na^+ (sodium ion) has +1 charge so 1 hook
O^{2-} (oxide ion) has -2 charge so 2 eyes



We need an extra Na^+ to give us a second hook to match the second of the eyes on the O^{2-} ion.

We have 2 Na^+ ions to every O^{2-} ion, so the formula is Na_2O</td>
<td>Example 2: What is the formula of magnesium hydroxide?
Mg^{2+} (magnesium ion) has +2 charge so 2 hooks
OH^- (hydroxide ion) has -1 charge so 1 eye



There are 2 OH^- ions to every Mg^{2+} ion so the formula is $Mg(OH)_2$

Note the use of a bracket to show 2 lots of OH which is not the same as OH_2. Brackets are most often used when the non-metallic ion contains more than one element.</td>
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Questions:

Deduce the formulae for the following ionic compounds:

1. sodium chloride
2. calcium bromide
3. sodium carbonate
4. aluminium oxide
5. iron(II) chloride
6. potassium oxide
7. aluminium chloride
8. potassium nitrate
9. aluminium sulfate
10. iron (III) nitrate

Charges on ions	
aluminium	Al^{3+}
chloride	Cl^-
oxide	O^{2-}
bromide	Br^-
iron(II)	Fe^{2+}
potassium	K^+
calcium	Ca^{2+}
iron(III)	Fe^{3+}
sodium	Na^+
carbonate	CO_3^{2-}
nitrate	NO_3^-
sulfate	SO_4^{2-}

Moles

A Mole Is a Number of Particles

If you wanted to count the number of atoms that you had in a sample of a substance, you'd have to use some very big numbers, and spend a very long time counting. So you need a unit to describe the amount of a substance that you have — that unit is the mole.

One mole of substance contains 6.02×10^{23} particles
This number is known as Avogadro's number

The particles can be anything — e.g. atoms or molecules. So 6.02×10^{23} atoms of carbon is 1 mole of carbon, and 6.02×10^{23}3 molecules of CO_2 is 1 mole of CO_2

Molar Mass Is the Mass of One Mole

One mole of atoms or molecules has a mass in grams equal to the relative formula mass (A_r or M_r) of that substance.

Carbon has an A_r of 12 : 1 mole of carbon weighs 12 g : The molar mass of carbon is 12g/mole

CO_2, has an M_r of 44 : 1 mole of CO_2 weighs 44 g : The molar mass of CO_2 44g/mole

So you know that 12 g of carbon and 44 g of CO must contain the same number of particles.

You can use molar mass in calculations to work out how many moles of a substance you have.

$$\text{Number of moles} = \frac{\text{Mass of substance (g)}}{\text{Molar mass (g/mol)}}$$

Example:

- How many moles of sodium oxide are present in 24.8g of Na_2O?
- Molar mass of Na_2O = (2 x 23) ÷ (1 x 16) = 62g/mol
- Number of moles of Na_2O = 24.8 ÷ 62 = 0.4 moles

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Questions (refer to the periodic table)

1. Find the molar mass of zinc.

2. Find the molar mass of sulfuric acid, H_2SO_4

3. How many moles of sodium chloride are present in 117g of NaCl?

4. I have 54 g of water (H_2O) and 84 g of iron (Fe). Do I have more moles of water or of iron?

Equations and calculations

Calculate the relative molecular mass (M_r) of;

a) H_2 b) Ne

c) $Ca(OH)_2$ d) NH_4NO_3

Calculate the percentage by mass of the elements shown in the following compounds;

a) Ca in $Ca(OH)_2$ b) O in $Ca(OH)_2$

c) N in NH_4NO_3 d)Fe in $Fe(NO_3)_3$

Balance the following equations;

a)	F_2	+	KBr	$\rightarrow$	KF	+	Br_2
b)	Al	+	O_2	$\rightarrow$	Al_2O_3		
c)	$CuCO_3$			$\rightarrow$	CuO	+	CO_2
d)	C_4H_8	+	O_2	$\rightarrow$	CO_2	+	H_2O
e)	$Ba(OH)_2$	+	H_2SO_4	$\rightarrow$	$BaSO_4$	+	H_2O
f)	$FeCl_3$	+	NaOH	$\rightarrow$	$Fe(OH)_3$	+	NaCl

Calculating reacting masses

a) What mass of carbon dioxide is formed when 20g of calcium carbonate reacts with hydrochloric acid?

$$CaCO_3 + 2\,HCl \rightarrow CaCl_2 + H_2O + CO_2$$

b) What mass of carbon monoxide is needed to react with 1 kg of iron oxide?

$$Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$$

c) What mass of oxygen is needed to react with 184 g of sodium?

$$4\,Na + O_2 \rightarrow 2\,Na_2O$$

d) What mass of sodium carbonate is formed when 8.0g of sodium hydrogencarbonate ($NaHCO_3$) is decomposed by heat?

$$2\,NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

Research Exercise

Developing theories of atomic structure.

- You are to research how the accepted model of atomic structure has changed through history. Briefly outline each theory, highlighting their differences, and then explain why each theory was acceptance at the time and then finally rejected.

Useful Websites

- www.a-levelchemistry.co.uk/
- www.chemguide.co.uk/
- www.thestudentroom.co.uk
- www.alevelchem.com/
- www.creative-chemistry.org.uk
- www.revisionworld.co.uk/a2-level-level-revision/chemistry
- www.aqa.org.uk/subjects/science/as-and-a-level/chemistry-7404-7405

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