DIXONS SIXTH FORM BRADFORD

SUMMER WORK

A LEVEL CHEMISTRY

2026

STUDENT NAME:

About the Summer Work

This booklet contains a number of tasks that you are expected to complete to a hgih standard in order to be able to be enrolled in this subject.

You can use GCSE revision guides, and the links to websites and suggested textbooks below, to help you complete the tasks and to carry out further reading.

In September you will sit an induction test, this will be based on this summer.

Remember - first impressions are important.

Key websites and textbooks:

Hodder AQA A-level Chemistry textbook

Oxforad AQA A-level Chemistry textbook

CGP A-level Chemistry Revision Guide

https://www.aqa.org.uk/subjects/chemistry/a-level/chemistry-7405/specification/specification-at-a-glance

GCSE Chemistry (Single Science) - AQA - BBC Bitesize

https://chemrevise.org/

https://www.physicsandmathstutor.com/chemistry-revision/a-level-aqa/

Welcome to Chemistry

Chemistry is the central science, with links to Biology, Physics, Mathematics, and Engineering. Chemists design and synthesise medicines, investigate climate change and energy, create our everyday products, and develop new materials.

Chemistry is split into three disciplines:

Physical Chemistry

This is the study of macroscopic, atomic, subatomic and particulate phenomena in chemical systems. This encompasses a diverse range of topics including; atomic structure, bonding, amount of substance, kinetics, equilibria and pH.

Inorganic Chemistry

This is the study of the structure, properties and reactions of inorganic compounds. This section focuses on trends visible in the periodic table and properties exhibited by elements in particular groups, e.g. the transition metals.

Organic Chemistry

This is the study of the structure, properties and reactions of organic compounds. This section is concerned with the vast amount of covalent compounds of the element carbon, and links to topics studied in Biology. The nomenclature, structure, properties and reactions of these compounds is studied, in addition to analytical techniques.

As an excellent student in Chemistry you will follow these routines in order to develop your understanding and master concepts.

1. Review work after each lesson and complete the relevant independent study tasks.
2. Seek understanding and not be content until you have mastered a concept.
3. Practice daily.
4. Ask your teacher if unsure.

Outline of the course

	Physical Chemistry	Inorganic Chemistry	Organic Chemistry
Y1 Topics	Atomic structure Amount of substance bonding Bonding Energetics Kinetics Chemical equilibrium & Le Chatelier's Principle and Kc Oxidation and reduction	Periodicity Group 2 the alkaline earth metals Group 7 the halogens	Introduction to organic chemistry Alkanes Halogenoalkanes Alkenes Alcohols Organic analysis
Links to GSCE (AQA unit)	Atomic structure (C1) The periodic table (C1) Bonding, structure and properties (C2) Quantitative chemistry (C3) Energy changes (C5) Rate of reaction (C6) Reversible reactions and equilibrium (C6) Oxidation and reduction (C4)	The periodic table (C1) Group 1 (C1) Group 7 (C1) Group 0 (C1) Metal oxides (C4) Oxidation and reduction (C4) Reactivity series (C4) Extraction of metals (C4) Gas tests (C8)	Crude oil (C8) Hydrocarbons (C8) Alkanes and alkenes (C8) Fractional distillation (C8) Cracking (C8) Earth's atmosphere (C9) Atmospheric pollutants (C9) Earth's resources and sustainable development (C10)

	Physical Chemistry	Inorganic Chemistry	Organic Chemistry
Y2 topics	Thermodynamics Rate equations Equilibrium constant Kp Electrode Potentials and electrochemical cells Acids and bases	Properties of Period 3 elements and their oxides Transition Metals Reactions of ions in aqueous solution	Optical isomerism Aldehydes and ketones Carboxylic acids Aromatic chemistry Amines Polymers Amines, amino acids, proteins and DNA Organic Synthesis NMR Spectroscopy Chromatography
Links to GSCE (AQA unit)	Reactions of acids (C4) Electrolysis (C4) Energy changes (C5) Rate of reaction (C6) Reversible reactions and equilibrium (C6)	The periodic table (C1) Bonding, structure and properties (C2) Oxidation and reduction (C4)	Hydrocarbons (C8) Alkanes and alkenes (C8) Chromatography (C7)

Assessment

Topic	Paper 1	Paper 2	Paper 3
Physical Chemistry			
3.1.1 Atomic structure	Y		Y
3.1.2 Amount of substance	Y	Y	Y
3.1.3 Bonding	Y	Y	Y
3.1.4 Energetics	Y	Y	Y
3.1.5 Kinetics		Y	Y
3.1.6 Chemical equilibria, Le Chatelier's principle and Kc	Y	Y	Y
3.1.7 Oxidation, reduction and redox equations	Y		Y
3.1.8 Thermodynamics	Y		Y
3.1.9 Rate equations		Y	Y
3.1.10 Equilibrium constant Kp for homogeneous systems	Y		Y
3.1.11 Electrode potentials and electrochemical cells	Y		Y
3.1.12 Acids and bases	Y		Y
Inorganic Chemistry			
3.2.1 Periodicity	Y		Y
3.2.2 Group 2, the alkaline earth metals	Y		Y
3.2.3 Group 7(17), the halogens	Y		Y
3.2.4 Properties of Period 3 elements and their oxides	Y		Y
3.2.5 Transition metals	Y		Y
3.2.6 Reactions of ions in aqueous solution	Y		Y
Organic Chemistry			
3.3.1 Introduction to organic chemistry		Y	Y
3.3.2 Alkanes		Y	Y
3.3.3 Halogenoalkanes		Y	Y
3.3.4 Alkenes		Y	Y
3.3.5 Alcohols		Y	Y
3.3.6 Organic analysis		Y	Y
3.3.7 Optical isomerism		Y	Y
3.3.8 Aldehydes and ketones		Y	Y
3.3.9 Carboxylic acids and derivatives		Y	Y
3.3.10 Aromatic chemistry		Y	Y
3.3.11 Amines		Y	Y
3.3.12 Polymers		Y	Y
3.3.13 Amino acids, proteins and DNA		Y	Y
3.3.14 Organic synthesis		Y	Y
3.3.15 Nuclear magnetic resonance spectroscopy		Y	Y
3.3.16 Chromatography		Y	Y

Paper 1	+	Paper 2	+	Paper 3
What's assessed • Relevant Physical chemistry topics (sections 3.1.1 to 3.1.4, 3.1.6 to 3.1.8 and 3.1.10 to 3.1.12) • Inorganic chemistry (Section 3.2) • Relevant practical skills		**What's assessed** • Relevant Physical chemistry topics (sections 3.1.2 to 3.1.6 and 3.1.9) • Organic chemistry (Section 3.3) • Relevant practical skills		**What's assessed** • Any content • Any practical skills
How it's assessed • written exam: 2 hours • 105 marks • 35% of A-level		**How it's assessed** • written exam: 2 hours • 105 marks • 35% of A-level		**How it's assessed** • written exam: 2 hours • 90 marks • 30% of A-level
Questions 105 marks of short and long answer questions		**Questions** 105 marks of short and long answer questions		**Questions** 40 marks of questions on practical techniques and data analysis 20 marks of questions testing across the specification 30 marks of multiple choice questions

- AO1: Demonstrate knowledge and understanding of scientific ideas, processes, techniques and procedures.
- AO2: Apply knowledge and understanding of scientific ideas, processes, techniques and procedures:
  - in a theoretical context
  - in a practical context
  - when handling qualitative data
  - when handling quantitative data.
- AO3: Analyse, interpret and evaluate scientific information, ideas and evidence, including in relation to issues, to:
  - make judgements and reach conclusions
  - develop and refine practical design and procedures.

Assessment objectives (AOs)	Component weightings (approx %)			Overall weighting (approx %)
	Paper 1	Paper 2	Paper 3	
AO1	30	30	32	30
AO2	48	48	34	45
AO3	22	22	34	25
Overall weighting of components	35	35	30	100

20% of the overall assessment of A-level Chemistry will contain mathematical skills equivalent to Level 2 or above.

At least 15% of the overall assessment of A-level Chemistry will assess knowledge, skills and understanding in relation to practical work.

Higher Education & Careers

There are a range of undergraduate courses you can access with an A-level in Chemistry, for example:

- Chemistry MChem, BSc
- Chemical Engineering MEng, BEng
- Chemistry and Mathematics BSc
- Chemistry and Mathematics MChem, BSc
- Medicinal Chemistry BSc
- Medicinal Chemistry MChem, BSc
- Medicine
- Dentistry
- Pharmacy
- Radiology
- Optometry

With a Chemistry degree you'll be able to work in a diverse range of industries and roles, for example:

- Research Chemist
- Chemical Engineer
- Cosmetic Chemist
- Resin Technology Scientist
- Analytical Development Technician
- Equity Analyst
- Analytical Chemist
- Process Technologist
- Systems Manager
- Graduate Chemist
- Product Testing Analyst
- Health Care Assistant, NHS
- Drug Discovery Chemist
- Chemical Analyst
- Graduate Trainee

****SOME OF THE TASKS WILL REQUIRE YOU TO USE THIS PERIODIC TABLE****

The Periodic Table of the Elements

Key

relative atomic mass
symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
(1)	(2)	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	(13)	(14)	(15)	(16)	(17)	(18)
							1.0 **H** hydrogen 1										4.0 **He** helium 2
6.9 **Li** lithium 3	9.0 **Be** beryllium 4											10.8 **B** boron 5	12.0 **C** carbon 6	14.0 **N** nitrogen 7	16.0 **O** oxygen 8	19.0 **F** fluorine 9	20.2 **Ne** neon 10
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12											27.0 **Al** aluminium 13	28.1 **Si** silicon 14	31.0 **P** phosphorus 15	32.1 **S** sulfur 16	35.5 **Cl** chlorine 17	39.9 **Ar** argon 18
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30	69.7 **Ga** gallium 31	72.6 **Ge** germanium 32	74.9 **As** arsenic 33	79.0 **Se** selenium 34	79.9 **Br** bromine 35	83.8 **Kr** krypton 36
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	96.0 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48	114.8 **In** indium 49	118.7 **Sn** tin 50	121.8 **Sb** antimony 51	127.6 **Te** tellurium 52	126.9 **I** iodine 53	131.3 **Xe** xenon 54
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La *** lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium 77	195.1 **Pt** platinum 78	197.0 **Au** gold 79	200.6 **Hg** mercury 80	204.4 **Tl** thallium 81	207.2 **Pb** lead 82	209.0 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac †** actinium 89	[267] **Rf** rutherfordium 104	[268] **Db** dubnium 105	[271] **Sg** seaborgium 106	[272] **Bh** bohrium 107	[270] **Hs** hassium 108	[276] **Mt** meitnerium 109	[281] **Ds** darmstadtium 110	[280] **Rg** roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated						

* 58 – 71 Lanthanides	140.1 **Ce** cerium 58	140.9 **Pr** praseodymium 59	144.2 **Nd** neodymium 60	[145] **Pm** promethium 61	150.4 **Sm** samarium 62	152.0 **Eu** europium 63	157.3 **Gd** gadolinium 64	158.9 **Tb** terbium 65	162.5 **Dy** dysprosium 66	164.9 **Ho** holmium 67	167.3 **Er** erbium 68	168.9 **Tm** thulium 69	173.1 **Yb** ytterbium 70	175.0 **Lu** lutetium 71
† 90 – 103 Actinides	232.0 **Th** thorium 90	231.0 **Pa** protactinium 91	238.0 **U** uranium 92	[237] **Np** neptunium 93	[244] **Pu** plutonium 94	[243] **Am** americium 95	[247] **Cm** curium 96	[247] **Bk** berkelium 97	[251] **Cf** californium 98	[252] **Es** einsteinium 99	[257] **Fm** fermium 100	[258] **Md** mendelevium 101	[259] **No** nobelium 102	[262] **Lr** lawrencium 103

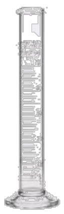
Practical Equipment

LOs

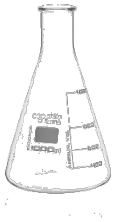
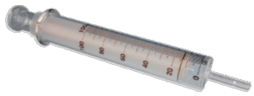
1. Recall the names and uses of common apparatus

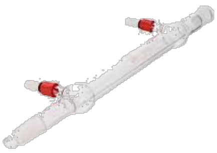
Below are images of most of the practical equipment that you will use during your Chemistry lessons. Name the pieces of apparatus using the options below and state what each is used for.

Beaker	Round bottom flask	Bunsen burner	Mass balance
Stopwatch	Test tube	Boiling tube	Condenser
Conical flask	Measuring cylinder	Filter funnel	Separating funnel
Burette	Volumetric pipette	Dropping pipette	Gas syringe

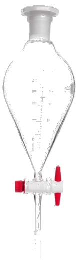


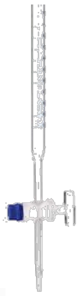











<u>**Variables**</u>

LOs

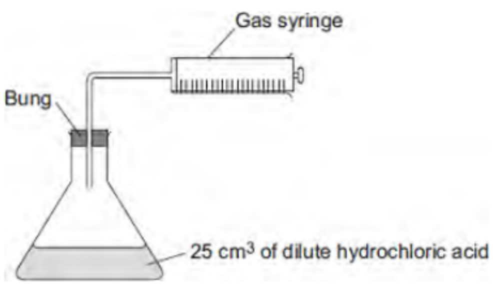
1. Recall the different types of variables
2. Identify the variables in a given experiment
3. Explain how control variables are kept the same

- <u>**Independent variable**</u> is the factor you change/are investigating the effects of (there is only 1 of these)
- <u>**Dependent variable**</u> is the factor you are measuring in your experiment (there is only one of these)
- <u>**Control variables**</u> are the factors you need to keep the same in order to make sure only one factor (the independent variable) is affecting the dependent variable (not the same equipment e.g. beaker). You should usually describe how you will keep these variables the same e.g. to keep the temperature the same you would use a water bath set at 30 °C

Identify the variables in the investigation below. Explain how you would keep the control variables the same.

Q2. A student investigated the reaction between magnesium metal and dilute hydrochloric acid.

The student placed 25 cm^3 of dilute hydrochloric acid in a conical flask and set up the apparatus as shown in the diagram.



The student:

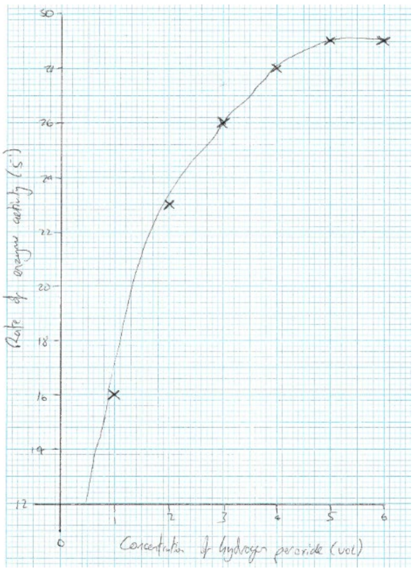
- took the bung out of the flask and added a single piece of magnesium ribbon 8 cm long
- put the bung back in the flask and started a stopwatch
- recorded the volume of gas collected after 1 minute
- repeated the experiment using different temperatures of acid.

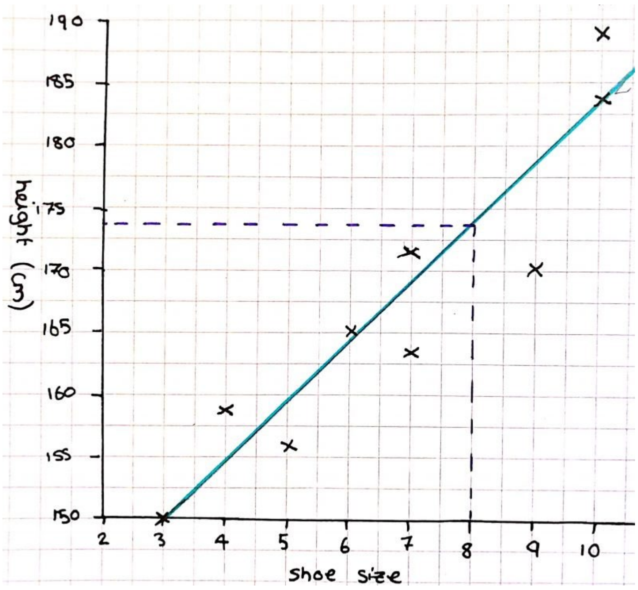
<u>Graphs</u>

LOs

1. Plot appropriate graphs
2. Use a tangent to determine the gradient of a curve at a particular point.

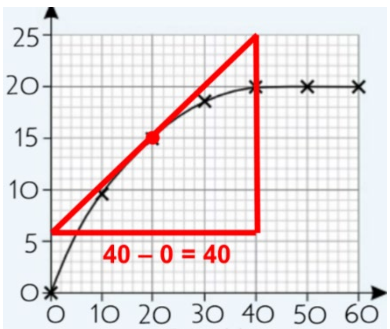
Most graphs that you draw in Chemistry will be scatter graphs e.g.





Points to remember when drawing a graph:

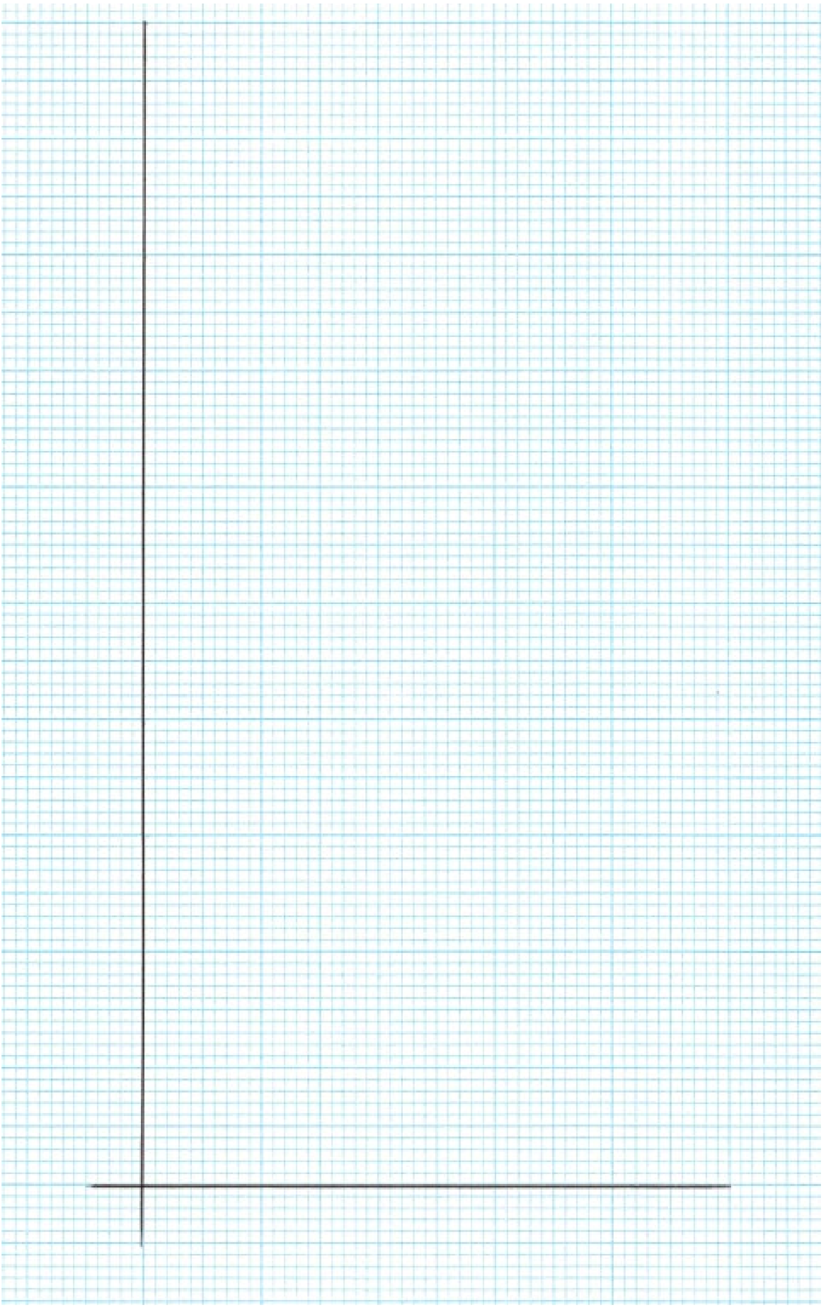
- Appropriate scales need to be chosen; going up by the same amount, the plotted points need cover at least half the scale, you don't need to start the axes at 0
- Label the axes, include units
- Plot points as clear crosses rather than small dots
- Circle any anomalous points
- Lines of best fit will either be a straight line using a ruler or a smooth curve which goes through or very close to each point. They will never be dot-to-dot
- The gradient can be calculated by doing Δy ÷ Δx (for a curve you will need to draw a tangent):



Gradient = 19 ÷ 40 = 0.475

Use the data in the table below to plot a graph of rate of reaction (mol dm^{-3} s^{-1}) against concentration (mol dm^{-3}). Plot both reactants on the same graph using a key.

Concentration (mol dm^{-3})	**Rate of reaction (mol dm^{-3} s^{-1})**	
	Reactant A	Reactant B
0.0	0.00	0.00
0.2	0.11	0.01
0.4	0.19	0.03
0.6	0.21	0.08
0.8	0.40	0.16
1.0	0.49	0.29



Draw a tangent on the line of best fit for reactant B at 0.7 mol dm^{-3} and calculate its gradient.

Converting Units

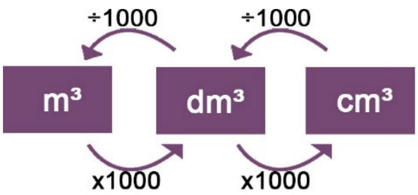
LOs

- Recall and convert units of mass, volume and temperature

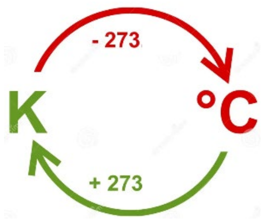
Each equation used in chemistry require values to be measured in specific units. What these units are depends on the equation, for example, the ideal gas equation requires volumes in m^3 whereas when calculating concentration volumes must be in dm^3.

You need to remember the units required and will be expected to convert values given to the required units.

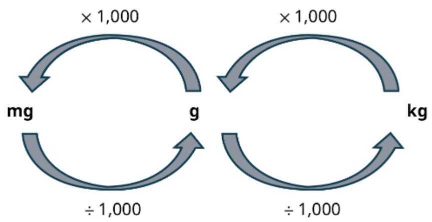
Converting volume:



Converting temperature:



Converting mass:



Convert the following quantities:

1. $25cm^3$ to dm^3
2. $5.1x10^{-3}$ m^3 to dm^3
3. $0.5x10^{-2}$ dm^3 to cm^3
4. 2.5 dm^3 to m^3
5. 25°C to K
6. 0K to °C
7. 100°C to K
8. 500°C to K
9. 250mg to g
10. $4.9x10^{-3}$Kg to g
11. 58g to Kg
12. 102Kg to mg

Rearranging Equations

LOs

- Rearrange equations involving common functions

You will be expected to rearrange equations to make a different term the subject.

To move a term/function from one side to the other you must perform the inverse of that on both sides.

$$y + 4 = x^2$$

$$\boxed{\sqrt{\ }}$$

$$\sqrt{y+4} = x$$

$$v = u + at$$
$$\boxed{-u} \quad \boxed{-u}$$
$$v - u = at$$
$$\boxed{\div a} \quad \boxed{\div a}$$
$$\frac{v-u}{a} = t$$

Rearrange the following equations to give the letter shown as the subject.

a) $T = \frac{2ab}{c}$ **b**

b) $PV = nRT$ **T**

c) $B = A(2r - p)$ **p**

(3)

a) $q = mcT$ **T**

b) $2a = 3(b - 2c)$ **c**

c) $F = mc + \frac{Rk}{T}$ **T**

d) $E = \frac{3pq}{r^2}$ **q**

e) $a^2 = bc^3$ **a**

f) $G = H - TS$ **T**

(6)

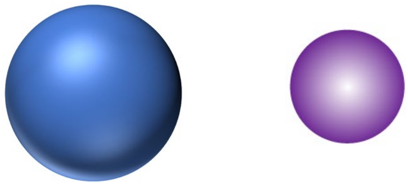
Atomic Structure

LOs

1. Outline the development of the model of the atom
2. Describe the current model of the atom
3. Recall the properties of subatomic particles

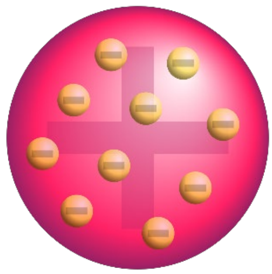
The model of the atom has been developed over time. Fill in the blanks to describe each model of the atom.

John Dalton 1803



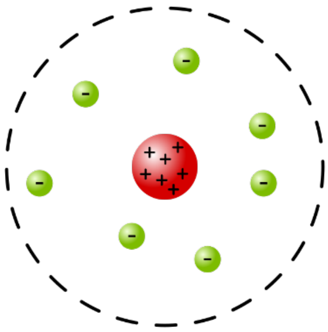
John Dalton described atoms as being s____________ s______________.

J J Thompson 1897



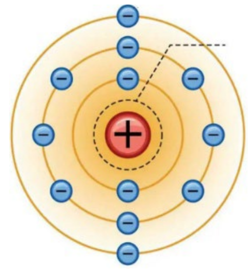
J J Thompson's model was called the p_________ p____________ model. It was a ball of p____________ charge that was evenly spread out, with n____________ charged e___________ embedded in it.

Ernest Rutherford 1909



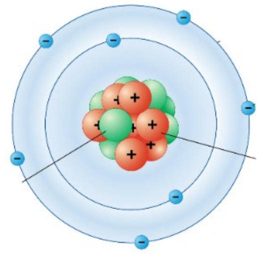
Rutherford fired a positively charged a________ particle at a sheet of g_______. Some of the particles went straight through showing that an atom is mostly e__________ s__________, and some of the particles were deflected straight which is evidence of a positively charged n__________.

Niels Bohr 1913



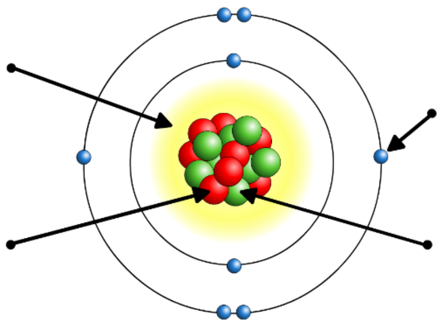
Bohr proposed that in order to prevent the atoms from collapsing, the electrons must be in fixed s__________ set at a fixed distance from the nucleus.

Rutherford and Chadwick 1932



Rutherford and Chadwick proposed that the nucleus contains positively charged p____________ and neutral n____________, making the model that we still use today.

Label this model of the atom below



Complete the table to show the relative masses and relative charges of the subatomic particles found in an atom.

Subatomic particle	Relative mass	Relative charge
Proton		
Electron		
Neutron		

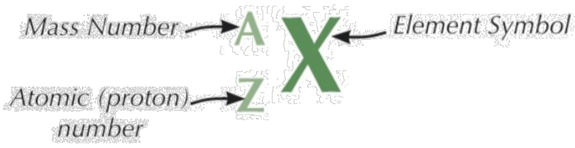
Atomic Number, Mass Number & Isotopes

LOs

1. Determine the number of protons, neutrons and electrons in an atom or ion using atomic number and mass number
2. Explain what an isotope is

The mass number of an atom/ion tells you the number of protons and neutrons in the nucleus.

The atomic number tells you the number of protons in the nucleus.



You can use these values to work out the number of protons, electrons, and neutrons that an atom/ion has. Atoms are neutral because they have the same number of protons and electrons so their charges cancel out. Positive ions have lost electrons and negative ions have gained electrons.

For example: an atom of ^{19}F has 9 protons, 9 electrons, and 10 neutrons.

$$^{19}_{9}F$$

Calculate the number of protons, electrons, and neutrons for the following atoms or ions.

Atom/ion	Protons	Neutrons	Electrons
$^{12}_{6}C$			
$^{31}_{15}P$			
$^{40}_{18}Ar$			
$^{23}_{11}Na^{+}$			
$^{35}_{17}Cl^{-}$			
$^{24}_{12}Mg^{2+}$			
$^{16}_{8}O^{2-}$			
$^{51}_{23}V^{3+}$			

Give the definition of an isotope.

Explain how these are all isotopes of lithium, with reference to the number of protons and neutrons in each.

$^{6}_{3}\mathrm{Li}$ $^{7}_{3}\mathrm{Li}$ $^{8}_{3}\mathrm{Li}$

Why do isotopes have the same chemical properties (react in the same way)?

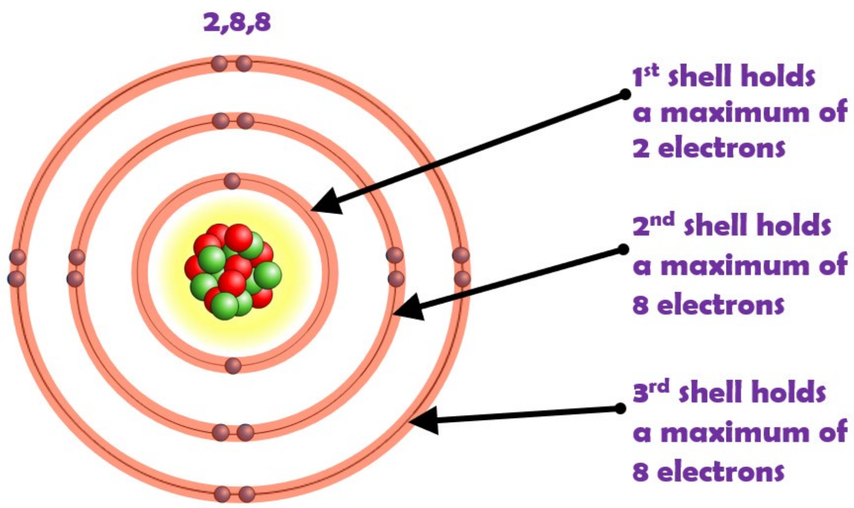
Electron configuration

LOs

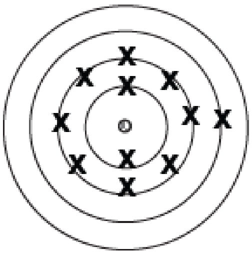
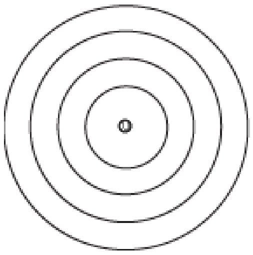
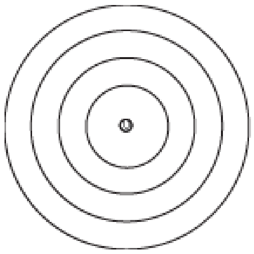
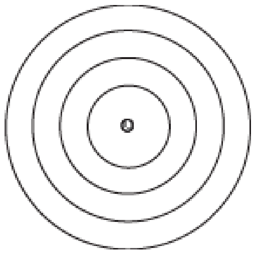
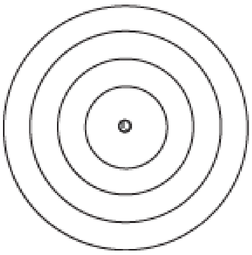
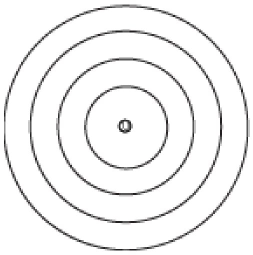
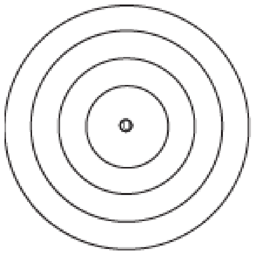
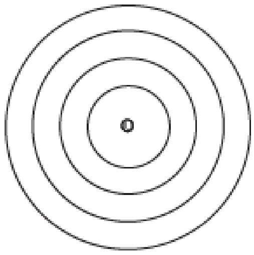
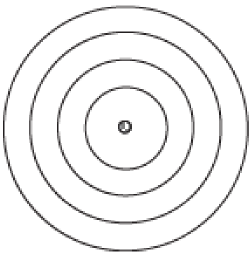
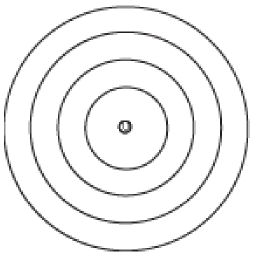
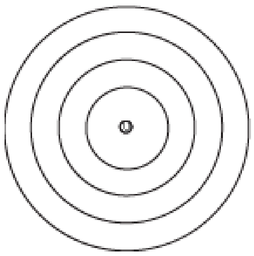
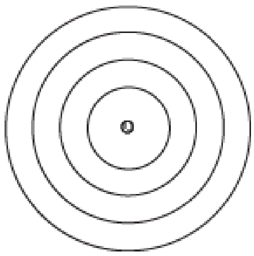
1. Draw the electron configuration of atoms and ions

Electrons orbit the nucleus in shells (energy levels). When drawing the electronic structure of an atom or ion, the inner shell is filled first, then each subsequent shell until all the electrons have been added.

When drawing the electronic structure of an ion, one electron is removed for each positive charge the ion has, and one electron is added for each negative charge it has. E.g. Cl^- will have one more electron than a Cl atom and Ca^{2+} will have two less electrons than a Ca atom.



Complete the electronic structure of the atoms and ions below.

1 Na structure = **2,8,1** 	2 Na^{+} structure = 	3 Al structure = 	4 Al^{3+} structure = 
5 Cl structure = 	6 Cl^{-} structure = 	7 O structure = 	8 O^{2-} structure = 
9 Ne structure = 	10 Ar structure = 	11 Mg structure = 	12 Mg^{2+} structure = 

<u>**The Periodic Table**</u>

LOs

- Explain how the arrangement of atoms in the periodic table is linked to their atomic structure.
- Explain reactivity and trends in reactivity of group 0, 1 and 7.

You will be given the following periodic table in your exams:

Key

relative atomic mass
symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
							1.0 **H** hydrogen 1										(18) 4.0 **He** helium 2
(1)	(2)											(13)	(14)	(15)	(16)	(17)	
6.9 **Li** lithium 3	9.0 **Be** beryllium 4											10.8 **B** boron 5	12.0 **C** carbon 6	14.0 **N** nitrogen 7	16.0 **O** oxygen 8	19.0 **F** fluorine 9	20.2 **Ne** neon 10
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	27.0 **Al** aluminium 13	28.1 **Si** silicon 14	31.0 **P** phosphorus 15	32.1 **S** sulfur 16	35.5 **Cl** chlorine 17	39.9 **Ar** argon 18
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30	69.7 **Ga** gallium 31	72.6 **Ge** germanium 32	74.9 **As** arsenic 33	79.0 **Se** selenium 34	79.9 **Br** bromine 35	83.8 **Kr** krypton 36
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	96.0 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48	114.8 **In** indium 49	118.7 **Sn** tin 50	121.8 **Sb** antimony 51	127.6 **Te** tellurium 52	126.9 **I** iodine 53	131.3 **Xe** xenon 54
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La** * lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium 77	195.1 **Pt** platinum 78	197.0 **Au** gold 79	200.6 **Hg** mercury 80	204.4 **Tl** thallium 81	207.2 **Pb** lead 82	209.0 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac** † actinium 89	[267] **Rf** rutherfordium 104	[268] **Db** dubnium 105	[271] **Sg** seaborgium 106	[272] **Bh** bohrium 107	[270] **Hs** hassium 108	[276] **Mt** meitnerium 109	[281] **Ds** darmstadtium 110	[280] **Rg** roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated						

* 58 – 71 Lanthanides	140.1 **Ce** cerium 58	140.9 **Pr** praseodymium 59	144.2 **Nd** neodymium 60	[145] **Pm** promethium 61	150.4 **Sm** samarium 62	152.0 **Eu** europium 63	157.3 **Gd** gadolinium 64	158.9 **Tb** terbium 65	162.5 **Dy** dysprosium 66	164.9 **Ho** holmium 67	167.3 **Er** erbium 68	168.9 **Tm** thulium 69	173.1 **Yb** ytterbium 70	175.0 **Lu** lutetium 71
† 90 – 103 Actinides	232.0 **Th** thorium 90	231.0 **Pa** protactinium 91	238.0 **U** uranium 92	[237] **Np** neptunium 93	[244] **Pu** plutonium 94	[243] **Am** americium 95	[247] **Cm** curium 96	[247] **Bk** berkelium 97	[251] **Cf** californium 98	[252] **Es** einsteinium 99	[257] **Fm** fermium 100	[258] **Md** mendelevium 101	[259] **No** nobelium 102	[262] **Lr** lawrencium 103

What is used to order the elements in the periodic table? ____________________________________

How is an element's atomic structure linked to the group it's in?

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Explain the reactivity of group 0 elements with reference to their atomic structure.

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Explain the trends in reactivity down group 1 and group 7, use the following terms in your answer:

Electron shell attraction lose gain nucleus

Group 1:

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Group 7:

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Ionic Bonding

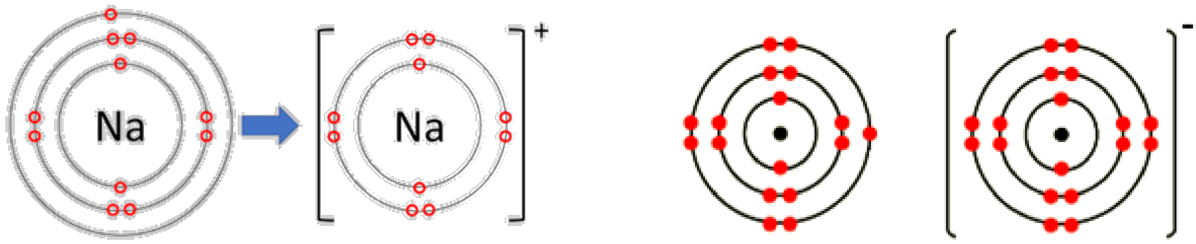
LOs

- Recall the definition of an ionic bond
- Describe how ionic bonds are formed
- Recall the formula of common compound ions
- Determine the formula of ionic compounds
- Describe the properties of ionic compounds and explain how they are linked to their structure

An ionic bond is the electrostatic attraction between oppositely charged ions.

Ionic bonding occurs between a metal and a non-metal. The metal atom loses electrons to become a positively charged ion with a full outer shell of electrons; the non-metal atom gains electrons to form a negatively charged ion with a full outer shell of electrons.

For example, when sodium reacts with chlorine to form sodium chloride (NaCl):



- Na atom loses 1 electron to form an Na^{1+} ion
- Cl atom gains 1 electron to form a Cl^{1-} ion

Describe what happens when magnesium reacts with oxygen. Name and give the formula of the product.

Describe what happens when sodium reacts with oxygen. Name and give the formula of the product.

Describe what happens when magnesium reacts with chlorine. Name and give the formula of the product.

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Compound ions are groups of atoms that have an overall charge. Common examples include:

Phosphate= PO_4^{3-}

Sulphate = SO_4^{2-}

Carbonate = CO_3^{2-}

Nitrate = NO_3^-

Hydroxide = OH^-

Ammonium= NH_4^+

Ionic compounds are neutral overall, so the charges of the positive and negative ions have to balance each other. For example, magnesium bromide is $MgBr_2$ because you need 2 Br- ions to balance the charge of the Mg^{2+} ion, sodium carbonate is Na_2CO_3 because you need 2 Na^+ ions to balance the charge of the CO_3^{2-} ion, and, calcium hydroxide is $Ca(OH)_2$ because you need 2 OH^- ions to balance the charge of the Ca^{2+} ion.

Name or give the formula of the following ionic compounds:

1) Lithium fluoride ____________
2) Calcium chloride ____________
3) Aluminium oxide ____________
4) Magnesium hydroxide ____________
5) Sodium nitrate ____________
6) Ammonium sulphate ____________

1) Na_2O ________________________

2) KNO_3 ________________________

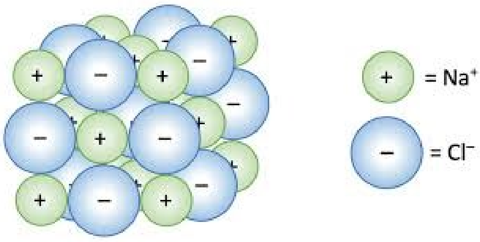
3) $NaOH$ ________________________

4) $Mg_3(PO_4)_2$ ________________________

5) Li_2CO_3 ________________________

6) $(NH_4)_2O$ ________________________

Ionic compounds have strong electrostatic forces of attraction between oppositely charged ions, in a giant lattice structure:



Ionic compounds have high melting and boiling points due to the strong electrostatic forces of attraction between oppositely charged ions in the giant lattice which requires lots of energy to overcome, so they are solids at room temperature. The greater the charge of the ion (more positive or more negative), the stronger the electrostatic forces of attraction between the oppositely charged ions.

Explain why $MgCl_2$ has a higher melting point than $NaCl$.

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Explain why MgO has a higher melting point than $MgCl_2$.

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Which ionic compound has a higher melting point: $CaBr_2$ or KCl? Explain your answer.

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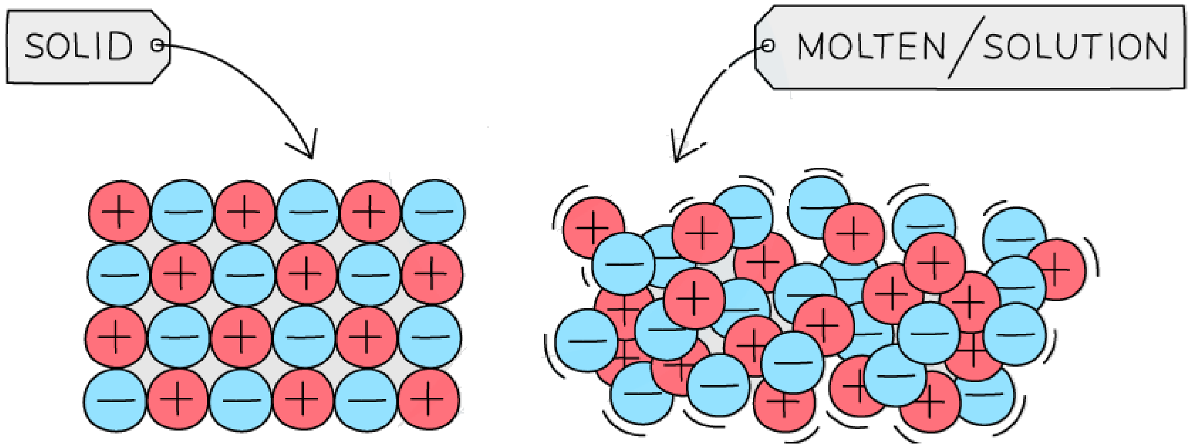

In which state(s) can ionic compounds conduct electricity. Explain your answer.

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Covalent Bonding

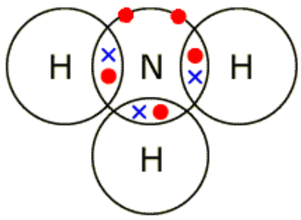
LOs

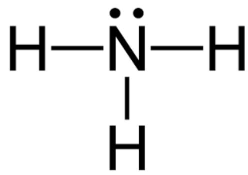
1. Recall the definition of a covalent bond
2. Draw dot and cross diagrams to represent covalent bonds
3. Describe the properties of covalent substances and explain how they are linked to their structure

A covalent bond is a shared pair of electrons, it occurs between 2 or more non-metal atoms, where they share electrons to have a full outer shell of electrons.

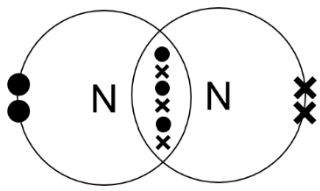
Dot and cross diagrams can be used to show how atoms share electrons in a covalent bond and the displayed formula shows the how atoms are bonded, with a line representing a covalent bond (a shared pair of electrons). Any electrons in the outer shell not involved in a covalent bond are called lone pairs and are represented as 2 dots in the displayed structure.

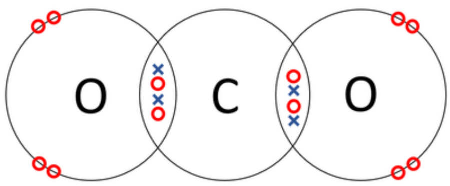
For example, ammonia has 3 H atoms bonded to 1 N atom. N is in group 5, so has 5 electrons in its outer shell. It needs 3 more electrons to have a full outer shell so needs to share a total of 3 pairs of electrons. H has 1 electron (in the first shell, which can hold a maximum of 2), so to have a full outer shell each atom needs to share 1 pair of electrons.

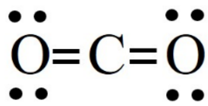




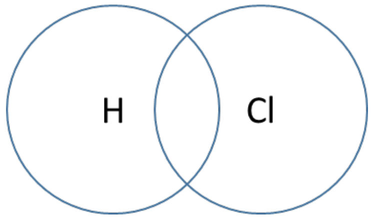
When atoms share more than 1 pair of electrons, this results in a double bond (2 shared pairs) or a triple bond (3 shared pairs). For example, in N_2 and CO_2.

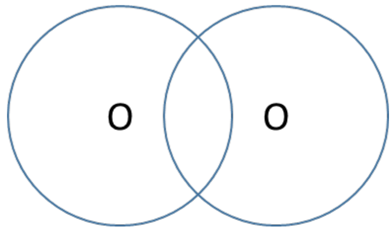


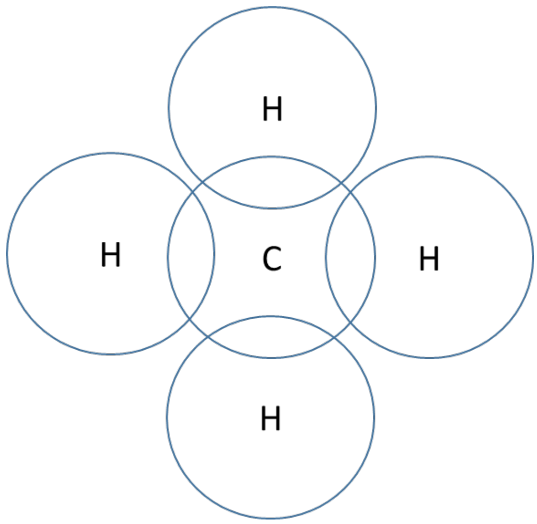


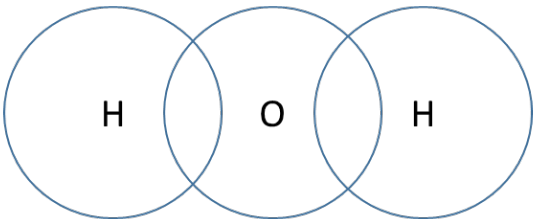


Complete the dot and cross diagram and draw the displayed formula for the molecules below.

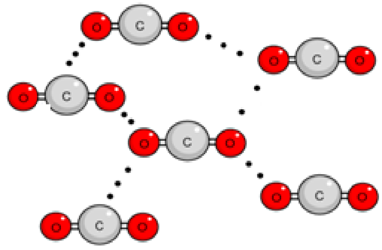




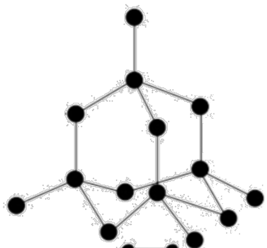


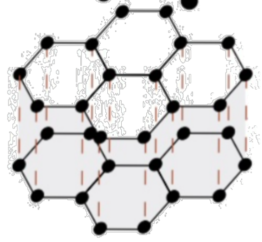


Simple covalent molecules (e.g. H_2O, CO_2, NH_3, etc.) are small molecules made up of only a few atoms. The small molecules are held together by weak intermolecular forces. When they are heated, the covalent bonds between the atoms (e.g. the double bond between the C=O) are not being broken, but the weak intermolecular forces holding the CO_2 molecules together require little energy to overcome, so they have low melting and boiling points and are normally gases or liquids at room temperature.



Giant covalent structures (macromolecules) such as diamond, graphite, and graphene are made up of lots of atoms. They have high melting and boiling points because rather than overcoming intermolecular forces, you have to break the many strong covalent bonds holding all of the atoms together, so they are solids at room temperature.

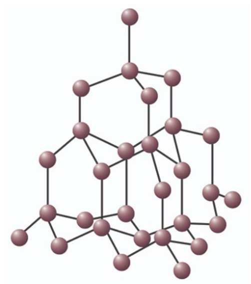




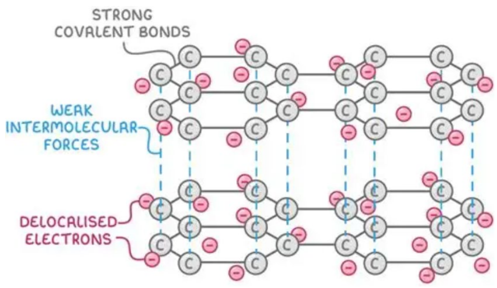
Explain why F_2 has a low melting point.

Silicon dioxide has a giant covalent structure. Explain why it has a high melting point.

Which has got a higher melting point: NO_2 or diamond? Explain your answer



Diamond



Graphite

Explain why graphite can conduct electricity but diamond can't.

Explain why graphite is soft but diamond is hard.

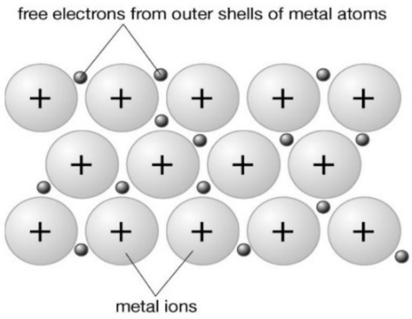
<u>Metallic Bonding</u>

LOs

1. Recall the definition of a metallic bond
2. Describe the properties of metallic substances and explain how they are linked to their structure

A metallic bond electrostatic attraction between positive ions and delocalised electrons.

In metals the outer electron(s) of the metal atom become delocalised to form a sea of delocalised electrons and a positive metal ion. These are held in a giant lattice structure by the electrostatic attraction between the positive metal ions and the sea of delocalised electrons. This attraction is strong and requires a lot of energy to overcome. In an electrical circuit, the delocalised electrons are able to carry charge and flow through the structure.



Explain why iron is a solid at room temperature.

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Explain why copper is used in electrical wires.

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Identifying Structure Types

LOs

1. Identify the bonding and structure type of substances from their name and formula

Structure type	Particles	Which substances
Monatomic	atoms	Group 0 elements
Simple molecular	molecules	Most non-metal elements (except Group 0), most compounds made from non-metals combined
Giant covalent	atoms	Diamond, graphite, silicon, silicon dioxide
Ionic	Ions	Most compounds made from metals and non-metals combined
Metallic	Ions & delocalised electrons	Metals

Substance	Formula	Structure type					Particles
		Monatomic	Simple molecular	Giant covalent	Ionic	Metallic	
sulfur	S_8						
magnesium	Mg						
hydrogen sulphide	H_2S						
neon	Ne						
ammonia	NH_3						
ammonium chloride	NH_4Cl						
calcium bromide	$CaBr_2$						
silicon dioxide	SiO_2						
silver nitrate	$AgNO_3$						
sodium hydroxide	NaOH						
magnesium oxide	MgO						
nitrogen dioxide	NO_2						
xenon	Xe						
bromine	Br_2						
copper carbonate	$CuCO_3$						
aluminium	Al						
ethane	C_2H_6						
glucose	$C_6H_{12}O_6$						
sodium sulphide	Na_2S						
methylamine	CH_3NH_2						

Balancing equations

LOs

1. Balance equations for given reactions

Balanced equations need to have the same number of all atoms on both sides of the equation. If the number of atoms is not the same on both sides, it must be balanced by placing 'big number' in front of the substance rather than by changing the 'small number' e.g. if you need 4 oxygen atoms you would do 2 CO_2 rather than changing it to CO_4.

It might be useful to use a table and count the number of each atoms on both sides of the equation:

Atom	Left	Right

Balance the equations below:

1) CH_4 + ___ O_2 → CO_2 + ___ H_2O

2) $Mg(OH)_2$ + ___ HCl → $MgCl_2$ + ___ H_2O

3) ___ Na + ___ H_2O → ___ NaOH + H_2

4) N_2 + ___ H_2 → ___ NH_3

5) ___ NaCl + F_2 → ___ NaF + Cl_2

6) ___ V + ___ O_2 → ___ V_2O_5

7) S_8 + ___ O_2 → ___ SO_3

8) ___ $AlBr_3$ + ___ K_2SO_4 → ___ KBr + $Al_2(SO_4)_3$

Relative atomic mass (A_r) & relative molecular mass (M_r)

LOs

1. Recall the definitions of relative atomic mass and relative molecular mass
2. Calculate relative formula mass

Relative atomic mass it the average mass of an atom of an element divided by 1/12 the mass of an atom of carbon 12.

Relative molecular mass is the average mass of a molecule divided by 1/12 the mass of an atom of carbon 12. It is calculated by adding up the relative atomic masses of all the atoms in a molecule using the top number from the periodic table (there is one at the front of this booklet.). In A-level Chemistry, M_r and A_r should be recorded to 1 decimal place.

Examples:

- CO_2 – there is 1 C atom which has a mass of 12.0 and 2 O atoms which each have a mass of 16.0. So, $M_r = (1 \times 12.0) + (2 \times 16.0) = 44.0$
- $Mg(OH)_2$ – there is 1 Mg atom which has a mass of 24.3, 2 O atoms which each have a mass of 16.0, and 2 H atoms which each have a mass of 1.0. So, $M_r = (1 \times 24.3) + (2 \times 16.0) + (1 \times 1.0) = 58.3$
- $Al_2(SO_4)_3$ – there are 2 Al atoms which each have a mass of 27.0, 3 S atoms which each have a mass of 32.1, and 12 O atoms which each have a mass of 16.0. So, $M_r = (2 \times 27.0) + (3 \times 32.1) + (12 \times 16.0) = 342.3$

Use the periodic table at the front of the booklet to calculate the M_r of the molecules below. Make sure all your answers are given to 1 decimal place.

1) HNO_3 ____________________
2) Fe_2O_3 ____________________
3) $CaCO_3$ ____________________
4) $Zn(OH)_2$ ____________________
5) $(NH_4)_2SO_4$ ____________________
6) $Ca(ClO_3)_2$ ____________________
7) $PtCl_2(NH_3)$ ____________________

Moles

LOs

1. Use the equation linking mass, M_r and moles to calculate unknown values

To make the number of atoms easier to use, in Chemistry we have a unit called the mole: 1 mole of particles contains 6.022×10^{23} particles, 0.5 moles of particles contain 3.011×0^{23} particles, and 2 moles of particles contains 1.2044×10^{24} particles.

The mass and moles of a substance are linked by the equation: Mass = M_r x moles

1) Calculate the mass of 2 moles of CO_2

2) Calculate the mass of 0.1 moles of NaOH

3) Calculate the mass of 0.4 moles of $MgBr_2$

4) Calculate the number of moles of Al_2O_3 in 3.62 g of Al_2O_3

5) Calculate the number of moles of $CaCl_2$ in 2.18 g of $CaCl_2$

6) Calculate the number of moles of $BaSO_4$ in 1.89 g of $BaSO_4$

7) Calculate the number of moles of KNO_3 in 3.25 g of KNO_3

Reacting Mass calculations

LOs

1. Calculate the mass of product made or reactant used when given the mass of another reactant or product.

If that mass of reactant used or product formed is known, the mass of other reactants used, or product made can be calculated.

Examples:

What mass of MgO is formed when 12.2 g of Mg reacts with an excess of O_2?

$$2\ Mg + O_2 \rightarrow 2\ MgO$$

1) Use mass and M_r to calculate the moles of the reactant
2) Use the ratio of reactant:product to calculate the moles of product that can be formed
3) Use the moles and M_r of product to calculate the mass produced.

1) Moles(Mg) = 12.2÷24.3=0.50

2) Mg:MgO = 2:2 = 1:1

Moles(MgO) = 0.5x1=0.50

3) Mass = 40.3x0.50=20.15g

What mass of Fe is formed when 3.6 g of Fe_2O_3 reacts with an excess of Al.

$$Fe_2O_3 + 2\ Al \rightarrow 2\ Fe + Al_2O_3$$

1) Use mass and M_r to calculate the moles of the reactant
2) Use the ratio of reactant:product to calculate the moles of product that can be formed
3) Use the moles and M_r of product to calculate the mass produced.

1) Moles(Fe_2O_3) = 3.6÷159.6=0.023

2) Fe_2O_3: Fe = 1:2

Moles(Fe) = 0.023x2=0.046

3) Mass = 56.0x0.046=2.58g

1) What mass of SO_3 is formed when 2.5 g of SO_2 reacts with an excess of O_2.

$$2\ SO_2 + O_2 \rightarrow SO_3$$

2) What mass of CO_2 is formed when 4.6 g of Fe_2O_3 reacts with an excess of CO.

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

3) What mass of SO_2 is formed 0.8 g of H_2S reacts with an excess of O_2.

$$2\ H_2S + 3\ O_2 \rightarrow 2\ SO_2 + 2\ H_2O$$

4) What mass of NH_3 is formed when an excess of N_2 reacts with 0.6 g of H_2.

$$N_2 + 3\ H_2 \rightarrow 2\ NH_3$$

5) What mass of CO_2 is formed when 1.7 g of C_4H_8 reacts with an excess of O_2.

$$C_4H_8 + 6\ O_2 \rightarrow 4\ CO_2 + 4\ H_2O$$