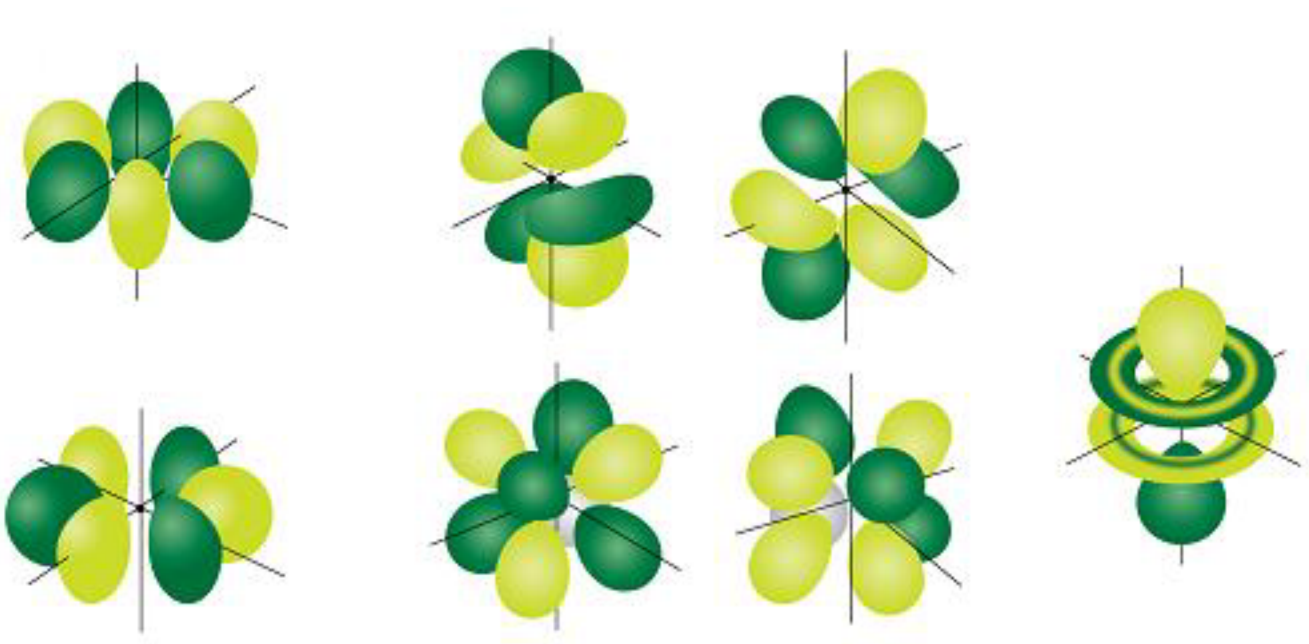
THE STRUCTURE OF ATOMS



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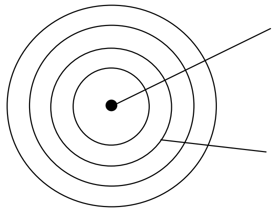
SECTION 1 – Basics from GCSE

Atoms consist of a central ……………………….. containing protons and ……………………… . The nucleus is ……………………. compared to the size of the whole atom. The nucleus is surrounded by ………………………… in energy levels (also called ………………). Atoms have no electric charge because they contain the same number of protons and ……………………….. .

sub-atomic particle	relative mass	relative charge
proton		
neutron		
electron		



Atomic number = number of …………………..

Mass number = number of ………………….. + number of …………………….

The number of protons, neutrons and electrons in an atom can be worked out using the atomic number and mass number.

Number of protons = ……………………………………………………..……………………………………………

Number of neutrons = ………………………………………………………………………………………………

Number of electrons = ………………………………………………………………………………………………

Atoms can be represented as follows:

$^{\text{mass number}}_{\text{atomic number}}\text{Symbol}$ e.g. $^{19}_{9}F$ protons = …..... neutrons = …..... electrons = …....

Atoms of the same element have the same number of ………….......... In fact, it is the number of ………………… that determines what type of atom it is (e.g. all atoms with 6 protons are carbon atoms). Atoms of different elements have different numbers of …………………. .

Isotopes are atoms with the same number of ………………. but a different number of ………………… . This means they are atoms of the same …………………….. with the same ……………… number but a different ……………… number.

	$^{35}_{17}Cl$	$^{37}_{17}Cl$
protons		
neutrons		
electrons		

Ions are charged particles with an unequal number of ……………….………. and ………………….………. .

Most ions have stable electron structures with the same electron structure as the elements in Group …………. .

Negative ions have ………….. electrons than protons.

Positive ions have ………….. electrons than protons.

TASK 1 – Atoms and ions

Species	Atom / ion	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
$^{14}_{7}N$	atom					
$^{31}_{15}P$	atom					
	atom	3	7			
	atom	10			10	
	atom		40	20		
	atom		40		22	
	atom			4	5	
	atom	82			126	
	atom	35			44	
	atom	35			46	
$^{23}_{11}Na^{+}$						
$^{16}_{8}O$						
$^{16}_{8}O^{2-}$						
		17	35			18
		19			20	19
		19			20	18
				20	20	18
		1			0	0
		53			74	54
			14		7	10

TASK 2 – Identify the particle

In each case identify the particle. The first one has been done for you.

1	An atom with 6 protons and the same number of neutrons as a ^{14}N atom	$^{13}_{6}C$
2	An atom with one more proton and the same number of neutrons than an atom of ^{39}K	
3	An atom with 10 protons and the same number of neutrons as an atom of ^{24}Mg	
4	An atom with one fewer proton and the same number of neutrons as an atom of ^{66}Zn	
5	An atom with the same number of protons and two more neutrons as an atom of ^{79}Br	
6	An atom with two fewer protons and the same number of neutrons as an atom of ^{50}Cr	
7	An ion with one more proton and two more neutrons as an atom of ^{20}Ne but the same number of electrons	
8	An ion with two fewer protons and two fewer neutrons as an atom of ^{40}Ar but the same number of electrons	
9	An ion with two more protons and two more neutrons as an atom of ^{60}Ni but the same number of electrons	
10	An ion with two more protons and three more neutrons as an atom of ^{20}Ne but the same number of electrons	
11	An ion with one fewer proton, one fewer neutron and the same number of electrons as an atom of ^{129}Xe.	
12	An ion with one more proton, two more neutrons, but the same number of electrons as an ion of $^{85}Rb^{+}$	
13	A particle with two fewer protons, two fewer neutrons and the same number of electrons as an atom of ^{20}Ne	
14	A particle with one fewer proton, two fewer neutrons and one more electron as a $^{48}Ti^{2+}$ ion	
15	A particle with one fewer proton, two more neutrons and the same number of electrons as a $^{127}I^{-}$ ion	

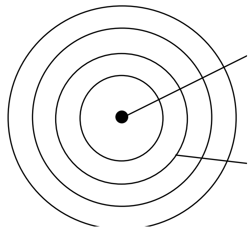
SECTION 2 – Development of atomic structure models

	Atoms	Plum-pudding model	Nuclear model		
lead scientist	John Dalton	JJ Thompson	Ernest Rutherford	Neils Bohr	James Chadwick
when	early 1800s	1897	1911	1913	1932
description of model					
what they discovered					
what they did					

SECTION 1 – Basics from GCSE

Atoms consist of a central **NUCLEUS** containing protons and **NEUTRONS**. The nucleus is **TINY** compared to the size of the whole atom. The nucleus is surrounded by **ELECTRONS** in energy levels (also called **SHELLS**). Atoms have no electric charge because they contain the same number of protons and **ELECTRONS**.

sub-atomic particle	relative mass	relative charge
proton	**1**	**+1**
neutron	**1**	**0**
electron	**0.0005 or $\frac{1}{1836}$**	**-1**



Atomic number = number of **PROTONS**

Mass number = number of **PROTONS** + number of **NEUTRONS**

The number of protons, neutrons and electrons in an atom can be worked out using the atomic number and mass number.

Number of protons = **ATOMIC NUMBER**

Number of neutrons = **MASS NUMBER - ATOMIC NUMBER**

Number of electrons = **ATOMIC NUMBER**

Atoms can be represented as follows:

$^{\text{mass number}}_{\text{atomic number}}\text{Symbol}$ e.g. $^{19}_{9}F$ protons = **9**   neutrons = **10**   electrons = **9**

Atoms of the same element have the same number of **PROTONS.** In fact, it is the number of **PROTONS** that determines what type of atom it is (e.g. all atoms with 6 protons are carbon atoms). Atoms of different elements have different numbers of **PROTONS**.

Isotopes are atoms with the same number of **PROTONS** but a different number of **NEUTRONS** . This means they are atoms of the same **ELEMENT** with the same **ATOMIC** number but a different **MASS** number.

	$^{35}_{17}Cl$	$^{37}_{17}Cl$
protons	**17**	**17**
neutrons	**18**	**20**
electrons	**17**	**17**

Ions are charged particles with an unequal number of **PROTONS** and **ELECTRONS**.

Most ions have stable electron structures with the same electron structure as the elements in Group **0/8/18**.

Negative ions have **MORE** electrons than protons.

Positive ions have **FEWER** electrons than protons.

TASK 1 – Atoms and ions

Species	Atom / ion	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
$^{14}_{7}N$	atom	**7**	**14**	**7**	**7**	**7**
$^{31}_{15}P$	atom	**15**	**31**	**15**	**16**	**15**
$^{7}_{3}Li$	atom	3	7	**3**	**4**	**3**
$^{20}_{10}Ne$	atom	10	**20**	**10**	10	**10**
$^{40}_{20}Ca$	atom	**20**	40	20	**20**	**20**
$^{40}_{18}Ar$	atom	**18**	40	**18**	22	**18**
$^{9}_{4}Be$	atom	**4**	**9**	4	5	**4**
$^{208}_{82}Pb$	atom	82	**208**	**82**	126	**82**
$^{79}_{35}Br$	atom	35	**79**	**35**	44	**35**
$^{81}_{35}Br$	atom	35	**81**	**35**	46	**35**
$^{23}_{11}Na^{+}$	**ion**	**11**	**23**	**11**	**12**	**10**
$^{16}_{8}O$	**atom**	**8**	**16**	**8**	**8**	**8**
$^{16}_{8}O^{2-}$	**ion**	**8**	**16**	**8**	**8**	**10**
$^{35}_{17}Cl^{-}$	**ion**	17	35	**17**	**18**	18
$^{39}_{19}K$	**atom**	19	**39**	**19**	20	19
$^{39}_{19}K^{+}$	**ion**	19	**39**	**19**	20	18
$^{40}_{20}Ca^{2+}$	**ion**	**20**	**40**	20	20	18
$^{1}_{1}H^{+}$	**ion**	1	**1**	**1**	0	0
$^{127}_{53}I^{-}$	**ion**	53	**127**	**53**	74	54
$^{14}_{7}N^{3-}$	**ion**	**7**	14	**7**	7	10

TASK 2 – Identify the particle

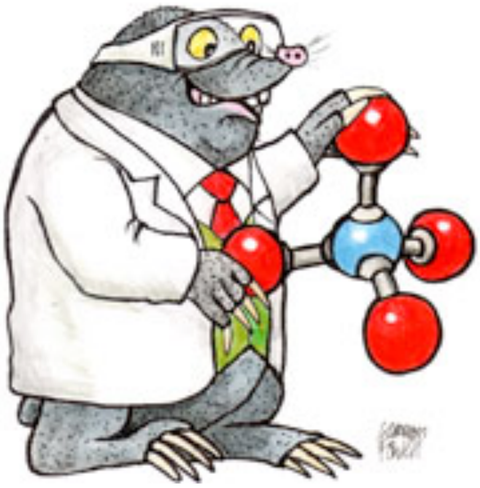
In each case identify the particle. The first one has been done for you.

1	An atom with 6 protons and the same number of neutrons as a ^{14}N atom	$^{13}_{6}C$
2	An atom with one more proton and the same number of neutrons than an atom of ^{39}K	$^{40}_{20}Ca$
3	An atom with 10 protons and the same number of neutrons as an atom of ^{24}Mg	$^{22}_{10}Ne$
4	An atom with one fewer proton and the same number of neutrons as an atom of ^{66}Zn	$^{65}_{29}Cu$
5	An atom with the same number of protons and two more neutrons as an atom of ^{79}Br	$^{81}_{35}Br$
6	An atom with two fewer protons and the same number of neutrons as an atom of ^{50}Cr	$^{48}_{22}Ti$
7	An ion with one more proton and two more neutrons as an atom of ^{20}Ne but the same number of electrons	$^{23}_{11}Na^{+}$
8	An ion with two fewer protons and two fewer neutrons as an atom of ^{40}Ar but the same number of electrons	$^{36}_{16}S^{2-}$
9	An ion with two more protons and two more neutrons as an atom of ^{60}Ni but the same number of electrons	$^{64}_{30}Zn^{2+}$
10	An ion with two more protons and three more neutrons as an atom of ^{20}Ne but the same number of electrons	$^{25}_{12}Mg^{2+}$
11	An ion with one fewer proton, one fewer neutron and the same number of electrons as an atom of ^{129}Xe.	$^{127}_{53}I^{-}$
12	An ion with one more proton, two more neutrons, but the same number of electrons as an ion of $^{85}Rb^{+}$	$^{88}_{38}Sr^{2+}$
13	A particle with two fewer protons, two fewer neutrons and the same number of electrons as an atom of ^{20}Ne	$^{16}_{8}O^{2-}$
14	A particle with one fewer proton, two fewer neutrons and one more electron as a $^{48}Ti^{2+}$ ion	$^{45}_{21}Sc$
15	A particle with one fewer proton, two more neutrons and the same number of electrons as a $^{127}I^{-}$ ion	$^{128}_{52}Te^{2-}$

SECTION 2 – Development of atomic structure models

	Atoms	Plum-pudding model	Nuclear model		
lead scientist	John Dalton	JJ Thompson	Ernest Rutherford	Neils Bohr	James Chadwick
when	early 1800s	1897	1911	1913	1932
description of model	**atom was the smallest particle**	**a ball of solid positive charge with negative electrons spread throughout**	**atom has a tiny positive nucleus surrounded by mainly empty space in which electrons are moving**	**atom has a tiny positive nucleus surrounded by mainly empty space in which electrons are moving in energy levels (shells)**	**atom has a tiny positive nucleus containing protons and neutrons, surrounded by mainly empty space in which electrons are moving in energy levels (shells)**
what they discovered		**electrons**	**the nucleus**	**electrons move in shells**	**neutrons in the nucleus**
what they did			**fired electrons at the thin piece of gold foil; a small proportion of alpha particles were deflected / bounced back**		

AMOUNT OF SUBSTANCE



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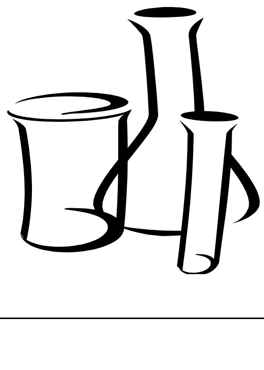


1 - FORMULAE

If you are serious about doing A level Chemistry, you **MUST** be able to write a formula without a second thought. It is the single most essential skill for an A level chemist.

You have to know and be able to use the information on this page – you should not be looking it up. There is no data sheet with ion charges at A level.

If you can't write a formula in an instant, **DROP CHEMISTRY NOW** and choose something else.

Elements

Monatomic	Simple molecular	Ionic	Metallic	Giant covalent
He helium Ne neon Ar argon Kr krypton Xe xenon Rn radon	H_2 hydrogen N_2 nitrogen O_2 oxygen F_2 fluorine Cl_2 chlorine Br_2 bromine I_2 iodine P_4 phosphorus S_8 sulfur	There are no ionic elements!!	The formula is just the symbol, e.g. Mg magnesium Fe iron Na sodium Ni nickel	The formula is just the symbol C diamond C graphite C graphine Si silicon

Compounds

Monatomic	Simple molecular	Ionic	Metallic	Giant covalent
There are no monatomic compounds!!	Some common molecular compounds: CO_2 carbon dioxide CO carbon monoxide NO nitrogen monoxide NO_2 nitrogen dioxide SO_2 sulfur dioxide SO_3 sulfur trioxide NH_3 ammonia CH_4 methane H_2S hydrogen sulfide	These have to be worked out using ion charges – you have to know these at AS/A level! LEARN them ASAP. Note these acids: HCl hydrochloric acid H_2SO_4 sulfuric acid HNO_3 nitric acid H_3PO_4 phosphoric acid	There are no metallic compounds!!	SiO_2 silicon dioxide

Positive ions		Negative ions	
Group 1 ions: Li^+ lithium Na^+ sodium K^+ potassium Group 2 ions: Mg^{2+} magnesium Ca^{2+} calcium Ba^{2+} barium	Group 3 ions: Al^{3+} aluminium Other common ions Mg^{2+} silver(I) Zn^{2+} zinc(II) NH_4^+ ammonium H^+ hydrogen	Group 7 ions: F^- fluoride Cl^- chloride Br^- bromide I^- iodide Group 6 ions: O^{2-} oxide S^{2-} sulfide	Other common ions NO_3^- nitrate SO_4^{2-} sulfate CO_3^{2-} carbonate HCO_3^- hydrogencarbonate OH^- hydroxide H^- hydride PO_4^{3-} phosphate

TASK 1 – WRITING FORMULAS OF IONIC COMPOUNDS

1) silver(I) bromide …………………………..
2) sodium carbonate …………………………..
3) potassium oxide …………………………..
4) iron(III) oxide …………………………..
5) chromium(III) chloride …………………………..
6) calcium hydroxide …………………………..
7) aluminium nitrate …………………………..
8) sodium sulfate …………………………..
9) lead(II) oxide …………………………..
10) sodium phosphate …………………………
11) zinc hydrogencarbonate …………………………..
12) ammonium sulfate …………………………..
13) gallium hydroxide …………………………..
14) strontium selenide …………………………..
15) radium sulfate …………………………
16) sodium nitride …………………………

TASK 2 – WRITING FORMULAS 1

1) lead(IV) oxide …………………………..
2) copper …………………………..
3) sodium …………………………..
4) ammonium chloride …………………………..
5) ammonia …………………………..
6) sulfur …………………………..
7) sulfuric acid …………………………..
8) neon …………………………..
9) silica …………………………..
10) silicon …………………………..
11) barium hydroxide …………………………..
12) tin(IV) chloride …………………………..
13) silver(I) nitrate …………………………..
14) iodine …………………………..
15) nickel …………………………..
16) hydrogen sulfide …………………………..
17) titanium(IV) oxide …………………………..
18) lead …………………………..
19) strontium sulfate …………………………..
20) lithium …………………………..

TASK 3 – WRITING FORMULAS 2

1) silver(I) carbonate …………………………..
2) gold …………………………..
3) platinum(II) fluoride …………………………..
4) nitric acid …………………………..
5) ammonia …………………………..
6) silicon(IV) hydride …………………………..
7) phosphorus …………………………..
8) diamond …………………………..
9) vanadium(V) oxide …………………………..
10) cobalt(II) hydroxide …………………………..
11) barium hydroxide …………………………..
12) ammonia …………………………..
13) hydrochloric acid …………………………..
14) fluorine …………………………..
15) silicon …………………………..
16) calcium phosphate …………………………..
17) rubidium …………………………..
18) germanium(IV) oxide …………………………..
19) magnesium astatide …………………………..
20) nitrogen monoxide …………………………..

2 - EQUATIONS

From an early age you should have been able to balance chemical equations. However, at A level, you will often need to:

- work out the formulas yourselves
- work out what is made (so you need to know some basic general equations)
- for reactions involving ions in solution, write ionic equations

Some general reactions you should know:

General Reaction	Examples
substance + oxygen → oxides	$2\ Mg + O_2 \rightarrow 2\ MgO$ $2\ H_2S + 3\ O_2 \rightarrow 2\ H_2O + 2\ SO_2$ $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$
metal + water → metal hydroxide + hydrogen	$2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$
metal + acid → salt + hydrogen	$Mg + 2\ HCl \rightarrow MgCl_2 + H_2$
oxide + acid → salt + water	$MgO + 2\ HNO_3 \rightarrow Mg(NO_3)_2 + H_2O$
hydroxide + acid → salt + water	$2\ NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
carbonate + acid → salt + water + carbon dioxide	$CuCO_3 + 2\ HCl \rightarrow CuCl_2 + H_2O + CO_2$
hydrogencarbonate + acid → salt + water + carbon dioxide	$KHCO_3 + HCl \rightarrow KCl + H_2O + CO_2$
ammonia + acid → ammonium salt	$NH_3 + HCl \rightarrow NH_4Cl$
metal carbonate → metal oxide + carbon dioxide (on heating)	$CaCO_3 \rightarrow CaO + CO_2$

TASK 4 – WRITING BALANCED EQUATIONS

1) Balance the following equations.

a) $Mg + HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

b) $CuCl_2 + NaOH \rightarrow Cu(OH)_2 + NaCl$

c) $SO_2 + O_2 \rightarrow SO_3$

d) $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$

2) Give balanced equations for the following reactions.

a) sodium + oxygen → sodium oxide

b) aluminium + chlorine → aluminium chloride

c) calcium + hydrochloric acid → calcium chloride + hydrogen

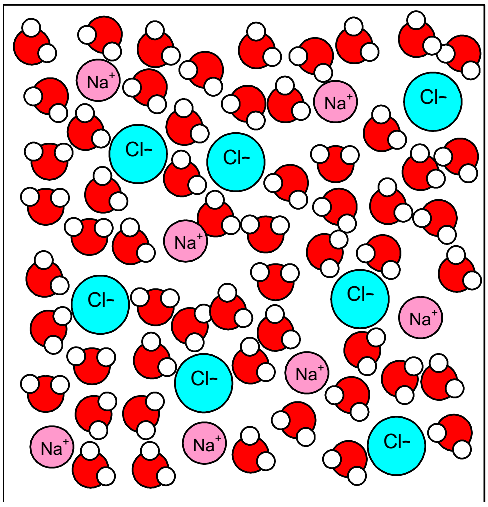
d) ammonia + sulfuric acid → ammonium sulfate

TASK 5 – WRITING BALANCED EQUATIONS 2

Write balance equations for the following reactions:

1) burning aluminium
2) burning hexane (C_6H_{14})
3) burning ethanethiol (CH_3CH_2SH)
4) reaction of lithium with water
5) reaction of calcium carbonate with nitric acid
6) thermal decomposition of lithium carbonate
7) reaction of ammonia with nitric acid
8) reaction of potassium oxide with sulfuric acid
9) reaction of calcium hydroxide with hydrochloric acid
10) reaction of zinc with phosphoric acid
11) reaction of sodium hydrogencarbonate with sulfuric acid
12) reaction of potassium hydroxide with sulfuric acid

Ionic equations



When an ionic substance dissolves in water, the positive and negative ions separate and become hydrated (they interact with water molecules rather than each other). For example, a solution of sodium chloride could also be described as a mixture of hydrated sodium ions and hydrated chloride ions in water.

In reactions involving ionic compounds dissolved in water, some of the ions may not be involved in the reaction. These are called ***spectator ions***. For such reactions, we can write an ***ionic equation*** that only shows the species that are involved in the reaction.

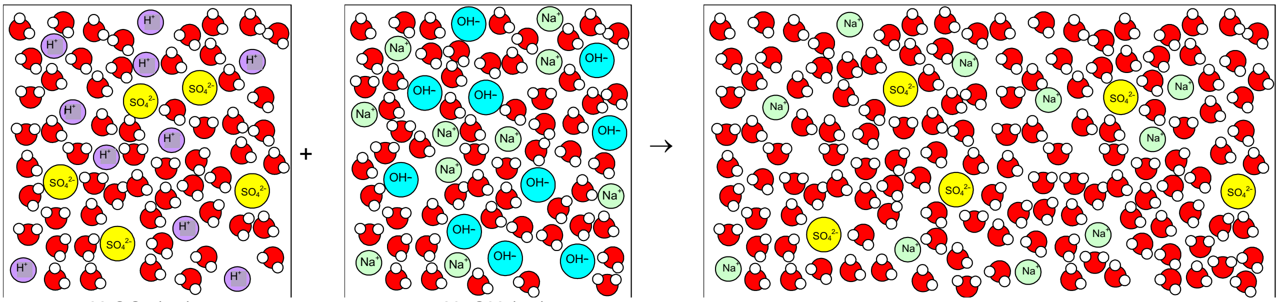
Simple examples are equations for which ionic equations can be written include:

Reactions of acids:

Common ionic equations are:	acid + hydroxide	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
	acid + carbonate	$2\ H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$
	acid + hydrogencarbonate	$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$
	acid + ammonia	$H^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq)$

We can even use these ionic equations to work out the ratio in which acids react without writing any equation.

For example, in the reaction of $H_2SO_4(aq)$ with $NaOH(aq)$ we know that one lot of H_2SO_4 contains two lots of H^+ ions. As H^+ ions react with OH^- ions in the ratio 1:1 [$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$] we know that we need two lots of NaOH to provide two lots of OH^- ions to react with the two lots of H^+ ions. Therefore, one lot of H_2SO_4 reacts with two lots of NaOH, i.e. the reacting ratio of H_2SO_4 : NaOH = 1:2

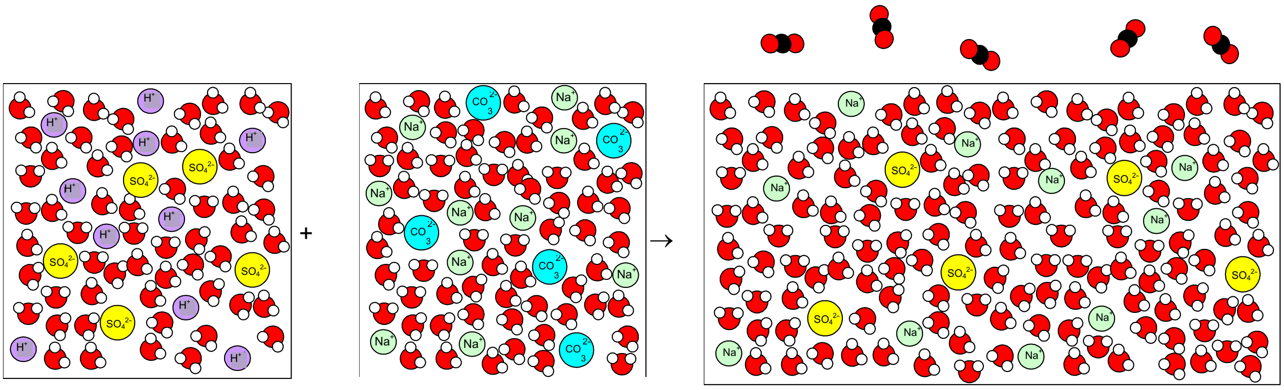


H_2SO_4 (aq) + NaOH (aq) →

contains H^+(aq) | contains OH^-(aq)

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

the Na^+(aq) and SO_4^{2-}(aq) ions are not involved



H_2SO_4 (aq) + Na_2CO_3 (aq) →

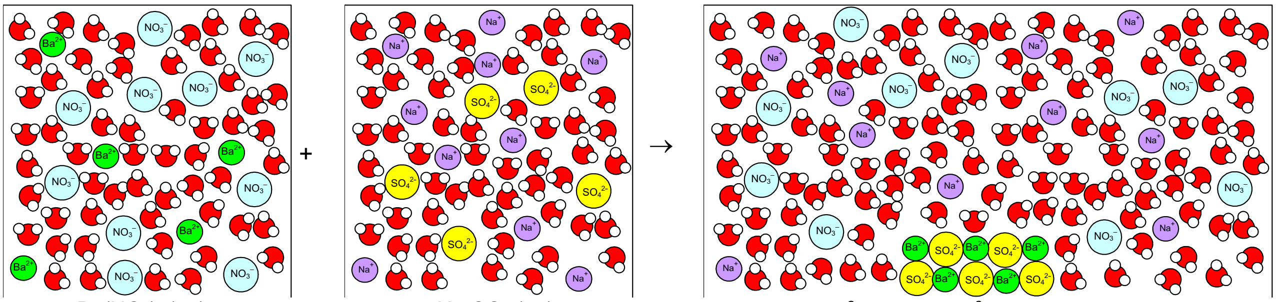
contains H^+(aq) | contains CO_3^{2-}(aq)

$2\,H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

the Na^+(aq) and SO_4^{2-}(aq) ions are not involved

Precipitation reactions

Some salts are insoluble in water. If solutions containing those ions are mixed, the insoluble salt forms as a solid as the solutions are mixed. This solid is known as a precipitate, and the reaction as precipitation.

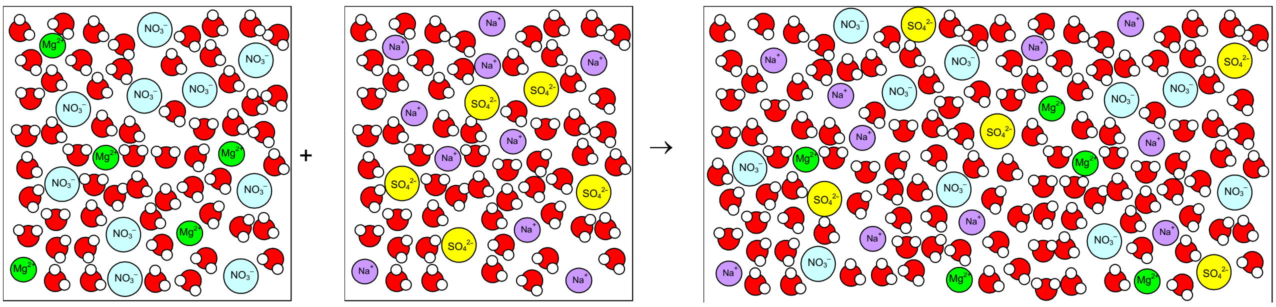


$Ba(NO_3)_2$ (aq) + Na_2SO_4 (aq) →

contains Ba^{2+}(aq) | contains SO_4^{2-}(aq)

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Most salts are soluble in water. Often when solutions of two salts are mixed, no such precipitation reaction will take place and the ions will remain dissolved in water.



$Mg(NO_3)_2$ (aq) + Na_2SO_4 (aq) →

Nothing happens – solutions just mix together (none of the combinations of ions give an insoluble compound)

TASK 6 – IONIC EQUATIONS

1) Use your knowledge of ionic equations to give the molar ratio in which the following acids react with bases. Complete the table to show your answers.

Acid	Formula of acid	Base	Formula of base	Molar ratio of acid:base
hydrochloric acid		lithium hydroxide		
sulfuric acid		sodium hydrogencarbonate		
nitric acid		ammonia		
sulfuric acid		potassium carbonate		
nitric acid		strontium hydroxide		

2) Write ionic equations for each of the following reactions.

a) reaction of hydrochloric acid (aq) with potassium hydroxide (aq)

b) precipitation of silver(I) iodide from reaction between silver(I) nitrate (aq) and potassium iodide (aq)

c) reaction of potassium carbonate (aq) with nitric acid (aq)

d) precipitation of calcium hydroxide from reaction between sodium hydroxide (aq) and calcium chloride (aq)

e) reaction of ammonia (aq) with hydrochloric acid (aq)

f) reaction of sodium hydrogencarbonate (aq) with sulfuric acid (aq)

g) precipitation of calcium sulfate from reaction between calcium chloride (aq) and sulfuric acid (aq)

h) precipitation of lead(II) chloride from reaction between lead(II) nitrate (aq) and sodium chloride (aq)

i) reaction of barium hydroxide (aq) with nitric acid (aq)

3 – SIGNIFICANT FIGURES & STANDARD FORM

Standard Form

- Standard form is very useful for writing very large or small numbers.
- They are written in the form A x 10^n where A is a number between 1 and 10.
- n represents the number of places the decimal point is moved (for +n values the decimal point has been moved to the left, for –n values the decimal point has been moved to the right).

Number	3435	1029000	0.025	23.2	0.0000278
Standard form	3.435 x 10^3	1.029 x 10^6	2.5 x 10^{-2}	2.32 x 10^1	2.78 x 10^{-5}

- To find the value of n:
  - for numbers greater than 1, n = number of places between first number and decimal place
  - for numbers less than 1, n = number of places from the decimal place to the first number (including that number)

Significant figures

Full number	1 sig fig	2 sig fig	3 sig fig	4 sig fig	5 sig fig
9.378652	9	9.4	9.38	9.379	9.3787
4204274	4000000	4200000	4200000	4204000	4204300
0.903521	0.9	0.90	0.904	0.9035	0.90352
0.00239482	0.002	0.0024	0.00239	0.002395	0.0023948

Significant figures for calculations involving multiplication / division

- Your final answer should be given to the same number of significant figures as the least number of significant figures in the data used.

e.g. Calculate the average speed of a car that travels 1557 m in 95 seconds.

average speed = $\frac{1557}{95}$ = 16 m s^{-1} (answer given to 2 sig fig as lowest sig figs in data is 2 sig fig for time)

e.g. Calculate the average speed of a car that travels 1557 m in 95.0 seconds.

average speed = $\frac{1557}{95}$ = 16.4 m s^{-1} (answer given to 3 sig fig as lowest sig figs in data is 3 sig fig for time)

Significant figures for calculations involving addition/subtraction ONLY

- Here the number of significant figures is irrelevant – it is about the place value of the data. For example

e.g. Calculate the total energy released when 263 kJ and 1282 kJ of energy are released.

Energy released = 263 + 1282 = 1545 kJ (answer is to nearest unit as both values are to nearest unit)

e.g. Calculate the total mass of calcium carbonate when 0.154 g and 0.01234 g are mixed.

Mass = 0.154 + 0.01234 = 0.166 g (answer is to nearest 0.001 g as least precise number is to nearest 0.001 g)

TASK 7 – SIGNIFICANT FIGURES & STANDARD FORM

1) Write the following numbers to the quoted number of significant figures.

a) 345789 4 sig figs …………………
b) 297300 3 sig figs …………………
c) 0.07896 3 sig figs …………………
d) 6.0961 3 sig figs ………………….
e) 0.001563 3 sig figs ………………….
f) 0.010398 4 sig figs ………………….

2) Complete the following sums and give the answers to the appropriate number of significant figures.

a) 6125 x 384 …………………
b) 25.00 x 0.010 …………………
c) 13.5 + 0.18 …………………
d) 7550 / 25 ………………….
e) 0.000152 x 13.00 ………………….
f) 0.0125 x 0.025 ………………….

3) Write the following numbers in non standard form.

a) 1.5×10^{-3} …………………..
b) 4.6×10^{-4} …………………......
c) 3.575×10^{5} …………………......
d) 5.34×10^{2} ……………………
e) 1.03×10^{6} ……………………
f) 8.35×10^{-3} ……………………

4) Write the following numbers in standard form.

a) 0.000167 …………………..
b) 0.0524 …………………......
c) 0.000000015 …………………......
d) 34500 ……………………
e) 0.62 ……………………
f) 87000000 ……………………

5) Complete the following calculations and give the answers to the appropriate number of significant figures.

a) $6.125 \times 10^{-3} \times 3.5$ ……………………………………………………………..
b) $4.3 \times 10^{-4} / 7.00$ ……………………………………………………………..
c) $4.0 \times 10^{8} + 35000$ ……………………………………………………………..
d) $0.00156 + 2.4 \times 10^{3}$ …………………………………………………………….
e) $6.10 \times 10^{-2} - 3.4 \times 10^{-5}$ …………………………………………………………….
f) $8.00 \times 10^{-3} \times 0.100 \times 10^{-3}$ …………………………………………………………….

4 – RELATIVE MASS

- Most elements are made of up atoms of different isotopes (e.g. chlorine contains both ^{35}Cl and ^{37}Cl atoms)
- The relative atomic mass (A_r) of an element is an average of the mass of the isotopes taking into account the relative abundance of each isotope.

Relative atomic mass, A_r	Average mass of an atom of an element relative to $\frac{1}{12}$th the mass of ^{12}C atom	
Relative formula mass, M_r	*If referring specifically to a molecule (relative molecular mass)* Average mass of a molecule of a substance relative to $\frac{1}{12}$th the mass of ^{12}C atom	
	More generally for any substance Sum of the relative atomic masses of all the atoms in the formula of a substance	

TASK 8 – RELATIVE FORMULA MASS

Calculate the M_r of each of these substances.

1 F_2 …………………………………………………………………………………………….

2 Fe …………………………………………………………………………………………….

3 H_2SO_4 …………………………………………………………………………………………….

4 Al_2O_3 …………………………………………………………………………………………….

5 $Mg(OH)_2$ …………………………………………………………………………………………….

6 $Al(NO_3)_3$ …………………………………………………………………………………………….

7 $(NH_4)_2SO_4$ …………………………………………………………………………………………….

8 $CuCO_3$ …………………………………………………………………………………………….

9 $AgNO_3$ …………………………………………………………………………………………….

10 NH_4NO_3 …………………………………………………………………………………………….

11 $CuSO_4.5H_2O$ …………………………………………………………………………………………….

12 magnesium …………………………………………………………………………………………….

13 oxygen …………………………………………………………………………………………….

14 sodium bromide …………………………………………………………………………………………….

15 calcium fluoride …………………………………………………………………………………………….

16 potassium sulfate …………………………………………………………………………………………….

17 chlorine …………………………………………………………………………………………….

18 iron(III) sulfate …………………………………………………………………………………………….

5 – THE MOLE & AVOGADRO CONSTANT

- One mole of anything contains 6.022 x 10^{23} of those things. One mole of bananas is 6.022 x 10^{23} bananas. One mole of water molecules is 6.022 x 10^{23} water molecules
- This number is known as the Avogadro constant (= 6.022 x 10^{23} mol^{-1}).
- The Avogadro number was chosen so that the mass of one mole of particles of a substance equals the M_r in grams. For example, the M_r of water is 18.0, and the mass of one mole of water molecules in 18.0 grams.

$$\text{Moles} = \frac{\text{mass (g)}}{M_r}$$

1 ton	=	1,000,000 g	(10^6)
1 kg	=	1,000 g	(10^3)
1 mg	=	0.001 g	(10^{-3})

Remember *Mr Moles!*

TASK 9A – MOLES

1) How many moles are there in each of the following?
   - a) 72.0 g of Mg
   - b) 4.00 kg of CuO
   - c) 39.0 g of $Al(OH)_3$
   - d) 1.00 tonne of NaCl
   - e) 20.0 mg of $Cu(NO_3)_2$

2) What is the mass of each of the following?
   - a) 5.00 moles of Cl_2
   - b) 0.200 moles of Al_2O_3
   - c) 0.0100 moles of Ag
   - d) 0.00200 moles of $(NH_4)_2SO_4$
   - e) 0.300 moles of $Na_2CO_3.10H_2O$

3) a) Calculate the number of moles of CO_2 molecules in 11.0 g of carbon dioxide.
   b) Calculate the number of moles of C atoms in 11.0 g of carbon dioxide.
   a) Calculate the number of moles of O atoms in 11.0 g of carbon dioxide.

4) a) Calculate the number of moles of Al_2O_3 in 5.10 g of Al_2O_3
   b) Calculate the number of moles of Al^{3+} ions in 5.10 g of Al_2O_3
   a) Calculate the number of moles of O^{2-} ions in 5.10 g of Al_2O_3

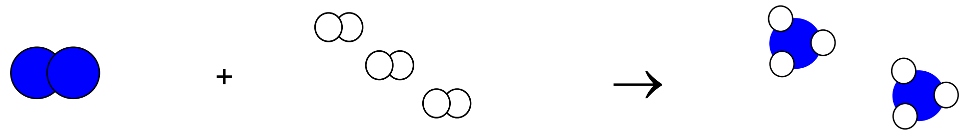
5) An experiment was carried out to find the M_r of vitamin C (ascorbic acid). It was found that 1.00 g contains 0.00568 moles of Vitamin C molecules. Calculate the M_r of vitamin C.

TASK 9B – AVOGADRO CONSTANT

1) Calculate the mass of one atom of ^{12}C
2) Calculate the number of H_2O molecules in 9.0 g of water.
3) Calculate the number of atoms in 42.0 g of nitrogen gas.

6 – REACTING MASS CALCULATIONS

What a chemical equation means



N_2	+	$3\ H_2$	→	$2\ NH_3$
1 molecule N_2		3 molecules H_2		2 molecules NH_3
12 molecules N_2 1 dozen molecules N_2		36 molecules H_2 3 dozen molecules H_2		24 molecules NH_3 2 dozen molecules NH_3
12 molecule N_2 1 dozen molecules N_2		36 molecules H_2 3 dozen molecules H_2		24 molecules NH_3 2 dozen molecules NH_3
6×10^{23} molecule N_2 1 moles N_2		18×10^{23} molecules H_2 3 moles H_2		12×10^{23} molecules NH_3 2 moles NH_3
10 moles N_2		30 moles H_2		20 moles NH_3
0.5 moles N_2		1.5 moles H_2		1 mole NH_3

TASK 10 – WHAT EQUATIONS MEAN

4 Na	**+**	**O_2**	→	**2 Na_2O**
12 mol				
0.1 mol				

2 Al	**+**	**3 Cl_2**	→	**2 $AlCl_3$**
5 mol				
0.1 mol				

C_4H_{10}	**+**	**6½ O_2**	→	**4 CO_2**	**+**	**5 H_2O**
0.5 mol						
20 mol						

4 NH_3	**+**	**3 O_2**	→	**2 N_2**	**+**	**6 H_2O**
0.5 mol						
10 mol						

Reacting mass calculations

- You can use balanced chemical equations to find out what mass of chemicals (or volume of gases) react or are produced in a chemical reaction. To do this, calculate:

  (a) moles of ✓ (b) moles of ? (c) mass of ?

e.g. What mass of iron is produced when 32.0 kg of iron(III) oxide is heated with CO?

✓ ?

$Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$

moles of $Fe_2O_3 = \frac{\text{mass (g)}}{M_r} = \frac{32000}{159.6} = 200.5$ mol

1 mole of Fe_2O_3 forms 2 moles of Fe

∴ moles of Fe = 2 x 200.5 = 401.0 mol

∴ mass of Fe = moles x M_r = 401.0 x 55.8 = **22,400 g (3 sig fig)**

e.g. What mass of oxygen is needed to convert 102 g of ammonia into nitrogen?

✓ ?

$4\ NH_3(g) + 3\ O_2(g) \rightarrow 2\ N_2(g) + 6\ H_2O(g)$

moles of $NH_3 = \frac{\text{mass (g)}}{M_r} = \frac{102}{17.0} = 6.00$ mol

4 moles of NH_3 reacts with 3 moles of O_2 ∴ 1 mole of NH_3 reacts with ¾ mole of O_2

∴ moles of O_2 = 6.00 x ¾ = 4.50 mol

∴ mass of O_2 = moles x M_r = 4.50 x 32.0 = **144 g (3 sig fig)**

e.g. When 5.00 g of crystals of hydrated tin (II) chloride, $SnCl_2.xH_2O$, are heated, 4.20 g of anhydrous tin(II) chloride are formed. Calculate the number of molecules of water of crystallisation are in $SnCl_2.xH_2O$ (i.e. the value of x).

$SnCl_2.xH_2O \rightarrow SnCl_2 + x\ H_2O$

moles of $SnCl_2 = \frac{\text{mass (g)}}{M_r} = \frac{4.20}{189.7} = 0.02214$ moles

∴ moles of $SnCl_2.xH_2O$ = 0.02214 mol

∴ M_r of $SnCl_2.xH_2O = \frac{\text{mass (g)}}{\text{moles}} = \frac{5.00}{0.02214} = 225.8$

∴ M_r of xH_2O = 225.8 – 189.7 = 36.1

∴ $x = \frac{36.1}{18.0} =$ **2** (x is a whole number and so the final answer is given as an integer)

TASK 11 – REACTING MASS CALCULATIONS 1

1) What mass of hydrogen is needed to react with 40.0 g of copper(II) oxide?

   $CuO + H_2 \rightarrow Cu + H_2O$

2) What mass of oxygen reacts with 192 g of magnesium?

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

3) What mass of sulfur trioxide is formed from 96.0 g of sulfur dioxide?

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

4) What mass of carbon monoxide is needed to react with 480 kg of iron(III) oxide?

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

5) What mass of carbon dioxide is produced when 5.60 g of butene is burnt.

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

6) What mass of oxygen is needed to react with 8.50 g of hydrogen sulfide (H_2S)?

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

7) 4.92 g of hydrated magnesium sulfate crystals ($MgSO_4.nH_2O$) gave 2.40 g of anhydrous magnesium sulfate on heating to constant mass. Work out the formula mass of the hydrated magnesium sulfate and so the value of *n*.

   $MgSO_4.nH_2O \rightarrow MgSO_4 + n\ H_2O$

8) In an experiment to find the value of *x* in the compound $MgBr_2.xH_2O$, 7.30 g of the compound on heating to constant mass gave 4.60 g of the anhydrous salt $MgBr_2$. Find the value of *x*.

   $MgBr_2.xH_2O \rightarrow MgBr_2 + x\ H_2O$

9) What mass of glucose must be fermented to give 5.00 kg of ethanol?

   $C_6H_{12}O_6 \rightarrow 2\ C_2H_5OH + 2\ CO_2$

10) The pollutant sulfur dioxide can removed from the air by reaction with calcium carbonate in the presence of oxygen. What mass of calcium carbonate is needed to remove 1.000 tonnes of sulfur dioxide?

    $2\ CaCO_3 + 2\ SO_2 + O_2 \rightarrow 2\ CaSO_4 + 2\ CO_2$

11) What mass of potassium oxide is formed when 7.80 mg of potassium is burned in oxygen?

    $4\ K + O_2 \rightarrow 2\ K_2O$

12) What mass of hydrogen is produced when 10.0 g of aluminium reacts with excess hydrochloric acid?

    $2\ Al + 6\ HCl \rightarrow 2\ AlCl_3 + 3\ H_2$

13) What mass of sodium just reacts with 40.0 g of oxygen?

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

14) What mass of nitrogen is produced when 2.00 tonnes of ammonia gas decomposes?

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

15) What mass of oxygen is produced when 136 g of hydrogen peroxide molecules decompose?

    $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$

16) What mass of lead(II) oxide is produced when 0.400 moles of lead(II) nitrate decomposes?

    $2\ Pb(NO_3)_2 \rightarrow 2\ PbO + 4\ NO_2 + O_2$

Limiting reagents

- In the real world of chemistry, it is rare that we react the exact right amount of chemicals together. Usually, we have more than we need of one of the reactants and so it doesn't all react – it is in excess.
- Sometimes in calculations, we need to work out if one of the reactants is in excess. The reactant that is not in excess is sometimes called the limiting reagent.

e.g. Propane reacts with oxygen as shown: $C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$

How many moles of products are formed when 1 mole of C_3H_8 is mixed with 8 moles of O_2?

	C_3H_8	+	$5\,O_2$	$\rightarrow$	$3\,CO_2$	+	$4\,H_2O$
moles at the start	1 mol		8 mol				
change in moles	1 mol react		5 mol react		3 mol made		4 mol made
moles at the end	1 – 1 = 0 mol		8 – 5 = 3 mol		0 + 3 = 3 mol		0 + 4 = 4 mol
	C_3H_8 limiting reagent		O_2 in excess				

e.g. Sulfur dioxide reacts with oxygen as shown: $2\,SO_2 + O_2 \rightarrow 2\,SO_3$

How many moles of SO_3 are formed when 5 mole of SO_2 is mixed with 2 moles of O_2?

	$2\,SO_2$	+	O_2	$\rightarrow$	$2\,SO_3$
moles at the start	5 mol		2 mol		
change in moles	4 mol react		2 mol react		4 mol made
moles at the end	5 – 4 = 1 mol		2 – 2 = 0 mol		0 + 4 = 4 mol
	SO_2 in excess		O_2 limiting reagent		

- In calculations you will be asked to work with masses, but you will need to convert to moles to find out which is the limiting reagent in order to work out the required answer.

e.g. In the manufacture of titanium, what mass of titanium can theoretically be formed when 1.00 kg of titanium chloride reacts with 0.100 kg of magnesium?

$$TiCl_4 + 2\,Mg \rightarrow Ti + 2\,MgCl_2$$

	$TiCl_4$	+	$2\,Mg$	$\rightarrow$	**Ti**	+	$2\,MgCl_2$
moles at the start	$\frac{1000}{189.9}$ = 5.266 mol		$\frac{100}{24.3}$ = 4.115 mol				

5.266 moles of $TiCl_4$ needs 10.53 moles of Mg to react

∴ $TiCl_4$ is in excess and does not all react, so Mg is the limiting reagent

∴ 2.058 moles of $TiCl_4$ reacts with 4.115 moles of Mg

	$TiCl_4$	+	$2\,Mg$	$\rightarrow$	**Ti**	+	$2\,MgCl_2$
change in moles	– 2.058 mol		– 4.115 mol		+ 2.058		+ 4.115 mol
moles at the end					0 + 2.058 = 2.058 mol		

∴ Mass of Ti = 2.058 x M_r = 2.058 x 47.9 = 98.6 g

TASK 12A – LIMITING REAGENTS 1

1		**CaO**	**+**	**H_2O**	**→**	**$Ca(OH)_2$**		
	a)	2 mol		3 mol				
	b)	10 mol		8 mol				
	c)	0.40 mol		0.50 mol				
2		**$2Ca$**	**+**	**O_2**	**→**	**$2CaO$**		
	a)	2 mol		2 mol				
	b)	10 mol		2 mol				
	c)	0.50 mol		0.20 mol				
3		**$2Fe$**	**+**	**$3Cl_2$**	**→**	**$2FeCl_3$**		
	a)	3 mol		3 mol				
	b)	12 mol		15 mol				
	c)	20 mol		40 mol				
4		**$TiCl_4$**	**+**	**$4Na$**	**→**	**Ti**	**+**	**$4NaCl$**
	a)	4 mol		4 mol				
	b)	2 mol		10 mol				
	c)	0.5 mol		1 mol				
5		**C_2H_5OH**	**+**	**$3O_2$**	**→**	**$2CO_2$**	**+**	**$3H_2O$**
	a)	15 mol		30 mol				
	b)	0.25 mol		1 mol				
	c)	3 mol		6 mol				
6		**N_2**	**+**	**$3H_2$**	**→**	**$2NH_3$**		
	a)	3 mol		6 mol				
	b)	0.5 mol		0.9 mol				
	c)	6 mol		20 mol				
7		**$4K$**	**+**	**O_2**	**→**	**$2K_2O$**		
	a)	10 mol		2 mol				
	b)	6 mol		4 mol				
	c)	0.50 mol		0.20 mol				

TASK 12B – LIMITING REAGENTS 2

1 What mass of calcium hydroxide is formed when 10.0 g of calcium oxide reacts with 10.0 g of water?

$CaO + H_2O \rightarrow Ca(OH)_2$

2 What mass of magnesium bromide is formed when 1.00 g of magnesium reacts with 5.00 g of bromine?

$Mg + Br_2 \rightarrow MgBr_2$

3 What mass of copper is formed when 2.00 g of copper(II) oxide reacts with 1.00 g of hydrogen?

$CuO + H_2 \rightarrow Cu + H_2O$

4 What mass of sodium fluoride is formed when 2.30 g of sodium reacts with 2.85 g of fluorine?

$2Na + F_2 \rightarrow 2NaF$

5 What mass of iron is formed when 8.00 g of iron(III) oxide reacts with 2.16 g of aluminium?

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

6 What mass of aluminium chloride is formed when 13.5 g of aluminium reacts with 42.6 g of chlorine?

$2Al + 3Cl_2 \rightarrow 2AlCl_3$

TASK 12C – REACTING MASS CALCULATIONS 2

1) 5.00 g of iron and 5.00 g of sulfur are heated together to form iron(II) sulfide. Which reactant is in excess and what is the maximum mass of iron(II) sulfide that can be formed?

   $Fe + S \rightarrow FeS$

2) In the manufacture of the fertiliser ammonium sulfate, what is the maximum mass of ammonium sulfate that can be obtained from 2.00 kg of sulfuric acid and 1.00 kg of ammonia?

   $H_2SO_4 + 2\ NH_3 \rightarrow (NH_4)_2SO_4$

3) In the Solvay process, ammonia is recovered by the reaction shown. What is the maximum mass of ammonia that can be recovered from 2.00 tonnes of ammonium chloride and 0.500 tonnes of calcium oxide?

   $2\ NH_4Cl + CaO \rightarrow CaCl_2 + H_2O + 2\ NH_3$

4) In the manufacture of titanium, what mass of titanium can theoretically be formed when 0.500 kg of titanium chloride reacts with 0.100 kg of magnesium?

   $TiCl_4 + 2\ Mg \rightarrow Ti + 2\ MgCl_2$

5) In the manufacture of ammonia, what mass of ammonia can theoretically be formed when 1.00 kg of nitrogen reacts with 0.500 kg of hydrogen?

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

6) In the manufacture of sulfur trioxide, what mass of sulfur trioxide can theoretically be formed when 1.00 kg of sulfur dioxide reacts with 0.500 kg of oxygen?

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

7) Hydrazine (N_2H_4) was used as the rocket fuel for the Apollo missions to the moon. It is by reaction of ammonia with sodium chlorate(I). What mass of hydrazine is made by reaction of 100 g of ammonia with 100 g of sodium chlorate(I)?

   $2\ NH_3 + NaOCl \rightarrow N_2H_4 + NaCl + H_2O$

1) A mixture of anhydrous sodium carbonate and sodium hydrogencarbonate of mass 10.000 g was heated until it reached a constant mass of 8.708 g. Calculate the composition of the mixture in grams of each component. Sodium hydrogencarbonate thermally decomposes to form sodium carbonate.

2) A mixture of calcium carbonate and magnesium carbonate with a mass of 10.000 g was heated to constant mass, with the final mass being 5.096 g. Calculate the percentage composition of the mixture, by mass.

3) 1 mole of a hydrocarbon of formula C_nH_{2n} was burned completely in oxygen producing carbon dioxide and water vapour only. It required 192 g of oxygen. Work out the formula of the hydrocarbon.

4) A mixture of $MgSO_4.7H_2O$ and $CuSO_4.5H_2O$ is heated at 120°C until a mixture of the anhydrous compounds is produced. If 5.00 g of the mixture gave 3.00 g of the anhydrous compounds, calculate the percentage by mass of $MgSO_4.7H_2O$ in the mixture.

Percentage yields

- When you make a new substance by a chemical reaction, you may not get all the expected amount of product. For example, if you reacted 4 g of hydrogen with 32 g of oxygen, you may get less than 36 g of water. Reasons include:
  - the reaction may be reversible (both the forwards and backwards reaction can take place)
  - some of the product may be lost when it is separated from the reaction mixture
  - some of the reactants may react in other reactions.

$$\% \text{ yield} = \frac{\text{mass of product obtained}}{\text{maximum theoretical mass of product}} \; x\ 100$$

e.g. Tungsten is made from tungsten(VI) oxide: $WO_3 + 3\ H_2 \rightarrow W + 3\ H_2O$

a) Calculate the maximum theoretical mass of tungsten that can be made from 0.500 tonne of tungsten(VI) oxide.

Moles of WO_3 = $\frac{\text{mass (g)}}{M_r} = \frac{500000}{231.8}$ = 2157 mol

∴ moles of W = 2157 mol

∴ mass of W = moles x M_r = 2157 x 183.8 = **396000 g** (3 sig fig)

b) In the reaction, only 350000 g of tungsten was made. Calculate the percentage yield.

% Yield = $\frac{\text{mass of product obtained}}{\text{maximum theoretical mass of product}} \; x\ 100 = \frac{350000}{396000} \; x\ 100$ = **88.3%** (3 sig fig)

TASK 13 – PERCENTAGE YIELD

1) Sulfur dioxide reacts with oxygen to make sulfur trioxide. $2\ SO_2 + O_2 \rightarrow 2\ SO_3$

   a) Calculate the maximum theoretical mass of sulfur trioxide that can be made by reacting 96.0 g of sulfur dioxide with an excess of oxygen.

   b) In the reaction, only 90.0 g of sulfur trioxide was made. Calculate the percentage yield.

   c) Give three reasons why the amount of sulfur trioxide made is less than the maximum theoretical maximum.

2) Iron is extracted from iron(III) oxide in the Blast Furnace as shown. $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$

   a) Calculate the maximum theoretical mass of iron that can be made from 1.00 tonne of iron(III) oxide.

   b) In the reaction, only 650000 g of iron (to 3 significant figures) was made. Calculate the percentage yield.

3) Nitrogen reacts with hydrogen to make ammonia. $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

   a) Calculate the maximum theoretical mass of ammonia that can be made by reacting 90.0 g of hydrogen with an excess of nitrogen.

   b) In the reaction, only 153 g of ammonia was produced. Calculate the percentage yield.

4) Titanium can be extracted from titanium chloride by the following reaction. $TiCl_4 + 2\ Mg \rightarrow Ti + 2\ MgCl_2$

   a) Calculate the maximum theoretical mass of titanium that can be extracted from 100 g of titanium chloride .

   b) In the reaction, only 20.0 g of titanium was made. Calculate the percentage yield.

   c) Give three reasons why the amount of titanium made is less than the maximum theoretical maximum.

5) Aluminium is extracted from aluminium oxide in the following reaction. $2\ Al_2O_3 \rightarrow 4\ Al + 3\ O_2$

   a) Calculate the maximum theoretical mass of aluminium that can be made from 1.00 kg of aluminium oxide.

   b) In the reaction, only 500 g of aluminium was made. Calculate the percentage yield.

6) The fertiliser ammonium sulfate is made as follows. $2\ NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$

   a) Calculate the maximum theoretical mass of ammonium sulfate that can be made by reacting 85.0 g of ammonia with an excess of sulfuric acid.

   b) In the reaction, only 300 g of ammonium sulfate was produced. Calculate the percentage yield.

7) 0.8500 g of hexanone, $C_6H_{12}O$, is converted into its 2,4-dinitrophenylhyrazone during its analysis. After isolation and purification, 2.1180 g of product $C_{12}H_{18}N_4O_4$ are obtained. Calculate the percentage yield.

Atom Economy

- Atom economy is a measure of what proportion of the products of a reaction are the desired product and how much is waste. The higher the atom economy, the less waste that is produced.

$$\% \text{ atom economy} = \frac{\text{mass of desired product as shown in equation}}{\text{total mass of products as shown in equation}} \times 100$$

e.g. making ethanol by fermentation

glucose		ethanol		carbon dioxide
$C_6H_{12}O_6(aq)$	→	$2\ CH_3CH_2OH(aq)$	+	$2\ CO_2$
180 g		92 g		88 g
		180 g products		

Atom economy = $\frac{92}{180} \times 100$ = 51%

Only 92 g of the 180 g of products is ethanol. This means that 51% of the mass of the products is ethanol, while the other 49% is waste.

TASK 14 – ATOM ECONOMY

1) Calculate the percentage atom economy to make sodium from sodium chloride. $2\ NaCl \rightarrow 2\ Na + Cl_2$

2) Calculate the percentage atom economy to make hydrogen from the reaction of zinc with hydrochloric acid. $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$

3) Calculate the percentage atom economy to make iron from iron(III) oxide in the Blast Furnace. $Fe_2O_3 + 3\ CO \rightarrow 2\ Fe + 3\ CO_2$

4) Calculate the percentage atom economy to make calcium oxide from calcium carbonate. $CaCO_3 \rightarrow CaO + CO_2$

5) Calculate the percentage atom economy to make sulfur trioxide from sulfur dioxide. $2\ SO_2 + O_2 \rightarrow 2\ SO_3$

6) Calculate the percentage atom economy to make oxygen from hydrogen peroxide. $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$

7) Hydrazine (N_2H_4) was used as the rocket fuel for the Apollo missions to the moon. It is by reaction of ammonia (NH_3) with sodium chlorate(I) (NaOCl).

ammonia + sodium chlorate → hydrazine + sodium chloride + water

$$2\ NH_3 + NaOCl \rightarrow N_2H_4 + NaCl + H_2O$$

a) Calculate the maximum theoretical mass of hydrazine that can be made by reacting 340 g of ammonia with an excess of sodium chlorate.

b) In the reaction, only 280 g of hydrazine was produced. Calculate the percentage yield.

c) Calculate the percentage atom economy for this way of making hydrazine.

d) Explain clearly the difference between atom economy and percentage yield.

7 – EMPIRICAL & MOLECULAR FORMULAS

- Every substance has an empirical formula. It shows the simplest ratio of atoms of each element in a substance.

  e.g. SiO_2 (giant covalent) – the ratio of Si:O atoms in the lattice is 1:2
  Al_2O_3 (ionic) – the ratio of Al^{3+}:O^{2-} ions in the lattice is 2:3
  H_2O (molecular) – the ratio of H:O atoms in the substance is 1:2

- Substances made of molecules also have a molecular formula. This indicates the number of atoms of each element in **one molecule**.

a) Finding the molecular formula from the formula mass and empirical formula

e.g. Empirical formula = CH_2, M_r = 42.0

Formula mass of empirical formula = 14.0 $\therefore \frac{M_r}{M_r \text{ of empirical formula}} = \frac{42.0}{14.0} = 3$

Molecular formula = 3 x empirical formula = C_3H_6

b) Finding the empirical formula of a compound from its composition by percentage or mass

i) Write out the mass or percentage of each element,

ii) Divide each mass or percentage by the A_r of the element **(not the M_r)**

iii) Find the simplest whole number ratio of these numbers by dividing by the smallest number. If the values come out as near $^1/_2$'s then times them by 2, if they are near $^1/_3$'s then times by 3.

e.g. i) A compound is found to contain, by mass, iron 72.4% and oxygen 27.6%.

Fe $\frac{72.4}{55.8} = 1.30$ O $\frac{27.6}{16.0} = 1.73$

Simplest ratio Fe:O = 1.30 : 1.73 (divide by smallest, i.e. 1.29)

$\frac{1.30}{1.30} : \frac{1.73}{1.30}$

1 : 1.33 (involves $^1/_3$'s so x3)

3 : 4

$\therefore$ empirical formula = $\mathbf{Fe_3O_4}$

e.g. ii) 0.25 g of hydrogen reacts with oxygen to produce 4.25 g of hydrogen peroxide (M_r = 34.0).

Mass of oxygen reacting with hydrogen = 4.25 – 0.25 = 4.00 g

H $\frac{0.25}{1.0} = 0.25$ O $\frac{4.00}{16.0} = 0.25$

Simplest ratio H : O = 0.25 : 0.25 (divide by smallest, i.e. 0.25)

1 : 1

$\therefore$ empirical formula = **HO**

Formula mass of empirical formula = 17.0

$\therefore \frac{M_r}{M_r \text{ of empirical formula}} = \frac{34.0}{17.0} = 2$

Molecular formula = 2 x empirical formula = $\mathbf{H_2O_2}$

TASK 15 – EMPRICIAL & MOLECULAR FORMULAS

1) The molecular formula of some substances are shown. Write the empirical formula of each substance.

a) C_2H_6	b) P_2O_3	c) SO_2	d) C_6H_{12}
e) $C_2H_4O_2$	f) C_2H_7N	g) B_6H_{10}	h) $C_{12}H_{22}O_{11}$

2) The empirical formula and relative molecular mass of some simple molecular compounds are shown below. Work out the molecular formula of each one.

a)	NH_2	$M_r = 32.0$	d)	PH_3	$M_r = 34.0$
b)	C_2H_5	$M_r = 58.0$	e)	CH	$M_r = 78.0$
c)	CH_2	$M_r = 70.0$	f)	CH_2	$M_r = 42.0$

3) Find the simplest whole number ratio for each of the following. The numbers come from experiments so there will be some small random errors which mean that you can round the numbers a little bit.

a) 1.5 : 1	b) 1 : 1.98	c) 4.97 : 1	d) 1 : 2.52
e) 1 : 1.33	f) 1.66 : 1	g) 1 : 1.26	h) 1 : 1.74

4) Find the empirical formulae of the following compounds using the data given.

a)	Ca	20 %	Br	80 %		
b)	Na	29.1 %	S	40.5 %	O	30.4 %
c)	C	53.3 %	H	15.5 %	N	31.1 %
d)	C	2.73 g	O	7.27 g		
e)	N	15.2 g	O	34.8 g		

5) 3.53 g of iron reacts with chlorine to form 10.24 g of iron chloride. Find the empirical formula of the iron chloride.

6) 50.0 g of a compound contains 22.4 g of potassium, 9.2 g of sulfur, and the rest oxygen. Calculate the empirical formula of the compound.

7) An oxide of phosphorus contains 56.4 % phosphorus and 43.6 % oxygen. Its relative molecular mass is 220. Find both the empirical and the molecular formula of the oxide.

8) A compound contains 40.0 g of carbon, 6.7 g of hydrogen and 53.5 g of oxygen. It has a relative molecular formula of 60. Find both the empirical and the molecular formula of the compound.

9) An organic compound X, which contains carbon, hydrogen and oxygen only, has an M_r of 85. When 0.43 g of X are burned in excess oxygen, 1.10 g of carbon dioxide and 0.45 g of water are formed. Find the empirical and molecular formulae of compound X.

10) When ammonium dichromate(VI) is added gradually to molten ammonium thiocyanate, Reinecke's salt is formed. It has the formula $NH_4[Cr(SCN)_x(NH_3)_y]$ and the following composition by mass: Cr = 15.5%, S = 38.15%, N = 29.2%. Calculate the values of *x* and *y* in the above formula.

9 – SOLUTION CALCULATIONS

$$\text{concentration} = \frac{\text{moles}}{\text{volume (dm}^3)}$$

a) Use the volume and concentration of one reactant to calculate the moles.

b) Use the balanced (or ionic) equation to find the moles of the other reactant.

c) Calculate the volume or concentration as required of that reactant.

Note

- Volume in dm^3 = $\frac{\text{volume in cm}^3}{1000}$ (e.g. 25 cm^3 → $\frac{25}{1000}$ = 0.025 dm^3)
- Concentration in g dm^{-3} = **concentration in mol dm^{-3} x M_r** (e.g. H_2SO_4 0.10 mol dm^{-3} → 0.10 x 98 dm^3 = 9.8 g dm^{-3})
- In many titrations, a standard solution of one the reagents is made (typically 250 cm^3 in a volumetric flask) and 25 cm^3 portions of this standard solution are used in each titration.

type of acid	what it means	examples	reacting ratio with NaOH
monoprotic	one H^+ ion per unit	HCl, HNO_3, CH_3COOH	1 : 1 (acid : NaOH)
diprotic	two H^+ ions per unit	H_2SO_4	1 : 2 (acid : NaOH)
triprotic	three H^+ ions per unit	H_3PO_4	1 : 3 (acid : NaOH)

E.g. 1: 25.0 cm^3 of 0.020 mol dm^{-3} sulfuric acid neutralises 18.6 cm^3 of sodium hydroxide solution.

$$H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow Na_2SO_4(s) + 2\,H_2O(l)$$

a) Find the concentration of the sodium hydroxide solution in mol dm^{-3}

moles of H_2SO_4 = conc x vol (dm^3) = $0.020\,x\,\frac{25.0}{1000}$ = 0.000500

moles of NaOH = conc x vol (dm^3) = 2 x moles H_2SO_4 = 0.000500 x 2 = 0.00100

concentration of NaOH = $\frac{\text{moles}}{\text{volume (dm}^3)}$ = $\frac{0.00100}{\frac{18.6}{1000}}$ = **0.0538 mol dm^{-3}**

b) Find the concentration of the sodium hydroxide solution in g dm^{-3}

M_r of NaOH = 23.0 + 16.0 + 1.0 = 40.0

mass of NaOH in 1 dm^3 = M_r x moles = 40.0 x 0.0538 = 2.15 g

concentration = **2.15 g dm^{-3}**

E.g. 2: Crystals of citric acid contain water of crystallisation ($C_6H_8O_7.nH_2O$). Citric acid is a triprotic acid. 1.52 g of the citric acid was made up to 250 cm^3 solution. 25 cm^3 portions of this solution required 21.80 cm^3 of 0.100 mol dm^{-3} sodium hydroxide for neutralisation. Calculate the M_r and the value of n.

Moles of NaOH = conc x vol (dm^3) = $0.100\,x\,\frac{21.80}{1000}$ = 0.00218

Moles of $C_6H_8O_7.nH_2O$ in each titration = $\frac{0.00218}{3}$ = 0.000727 (1 mol of acid reacts with 3 mol of NaOH)

Moles of $C_6H_8O_7.nH_2O$ in 250 cm^3 solution = 0.000727 x 10 = 0.00727

M_r of $C_6H_8O_7.nH_2O$ = <u>mass</u> = $\frac{1.52}{0.00727}$ = **209.2**

M_r of nH_2O = 209.2 – 192.0 = 17.2 n = $\frac{17.2}{18.0}$ = 0.954 = **1 (n is a whole number)**

TASK 18 – SOLUTION CALCULATIONS

1) Calculate the number of moles in the following.

a) 2 dm^3 of 0.05 mol dm^{-3} HCl
b) 50 litres of 5 mol dm^{-3} H_2SO_4
c) 10 cm^3 of 0.25 mol dm^{-3} KOH

2) Calculate the concentration of the following in **both** mol dm^{-3} and g dm^{-3}

a) 0.400 moles of HCl in 2.00 litres of solution
b) 12.5 moles of H_2SO_4 in 5.00 dm^3 of solution
c) 1.05 g of NaOH in 500 cm^3 of solution

3) Calculate the volume of each solution that contains the following number of moles.

a) 0.00500 moles of NaOH from 0.100 mol dm^{-3} solution
b) 1.00 x 10^{-5} moles of HCl from 0.0100 mol dm^{-3} solution

4) 25.0 cm^3 of 0.020 mol dm^{-3} sulfuric acid neutralises 18.6 cm^3 of barium hydroxide solution.

$$H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2\,H_2O$$

a) Find the concentration of the barium hydroxide solution in mol dm^{-3}
b) Find the concentration of the barium hydroxide solution in g dm^{-3}

5) 25.0 cm^3 of a solution of sodium hydroxide required 18.8 cm^3 of 0.0500 mol dm^{-3} H_2SO_4

$$H_2SO_4 + 2\,NaOH \rightarrow Na_2SO_4 + 2\,H_2O$$

a) Find the concentration of the sodium hydroxide solution in mol dm^{-3}
b) Find the concentration of the sodium hydroxide solution in g dm^{-3}

6) Calculate the volume of 0.05 mol dm^{-3} KOH is required to neutralise 25.0 cm^3 of 0.0150 mol dm^{-3} HNO_3.

$$HNO_3 + KOH \rightarrow KNO_3 + H_2O$$

7) 25.0 cm^3 of arsenic acid, H_3AsO_4, required 37.5 cm^3 of 0.100 mol dm^{-3} sodium hydroxide for neutralisation.

$$3\,NaOH(aq) + H_3AsO_4(aq) \rightarrow Na_3AsO_4(aq) + 3\,H_2O(l)$$

a) Find the concentration of the acid in mol dm^{-3}
b) Find the concentration of the acid in g dm^{-3}

8) A 250 cm^3 solution of NaOH was prepared. 25.0 cm^3 of this solution required 28.2 cm^3 of 0.100 mol dm^{-3} HCl for neutralisation. Calculate what mass of NaOH was dissolved to make up the original 250 cm^3 solution.

$$HCl + NaOH \rightarrow NaCl + H_2O$$

9) What volume of 5.00 mol dm^{-3} HCl is required to neutralise 20.0 kg of $CaCO_3$?

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

10) 3.88 g of a monoprotic acid was dissolved in water and the solution made up to 250 cm^3. 25.0 cm^3 of this solution was titrated with 0.095 mol dm^{-3} NaOH solution, requiring 46.5 cm^3. Calculate the relative molecular mass of the acid.

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TASK 1 – Writing formulas of ionic compounds

1	AgBr	2	Na_2CO_3	3	K_2O	4	Fe_2O_3	5	$CrCl_3$	6	$Ca(OH)_2$
7	$Al(NO_3)_3$	8	Na_2SO_4	9	PbO	10	Na_3PO_4	11	$Zn(HCO_3)_2$	12	$(NH_4)_2SO_4$
13	$Ga(OH)_3$	14	SrSe	15	$RaSO_4$	16	Na_3N				

TASK 2 – Writing formulas 1

1	PbO_2	2	Cu	3	Na	4	NH_4Cl	5	NH_3	6	S_8
7	H_2SO_4	8	Ne	9	SiO_2	10	Si	11	$Ba(OH)_2$	12	$SnCl_4$
13	$AgNO_3$	14	I_2	15	Ni	16	H_2S	17	TiO_2	18	Pb
19	$SrSO_4$	20	Li								

TASK 3 – Writing formulas 2

1	Ag_2CO_3	2	Au	3	PtF_2	4	HNO_3	5	NH_3	6	SiH_4
7	P_4	8	C	9	V_2O_5	10	$Co(OH)_2$	11	$Ba(OH)_2$	12	NH_3
13	HCl	14	F_2	15	Si	16	$Ca_3(PO_4)_2$	17	Rb	18	GeO_2
19	$MgAt_2$	20	NO								

TASK 4 – Writing balanced equations 1

1 a $Mg + 2\ HNO_3 \rightarrow Mg(NO_3)_2 + H_2$
b $CuCl_2 + 2\ NaOH \rightarrow Cu(OH)_2 + 2\ NaCl$
c $2\ SO_2 + O_2 \rightarrow 2\ SO_3$
d $C_4H_{10} + 6\frac{1}{2}\ O_2 \rightarrow 4\ CO_2 + 5\ H_2O$ or $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$

2 a $4\ Na + O_2 \rightarrow 2\ Na_2O$
b $2\ Al + 3\ Cl_2 \rightarrow 2\ AlCl_3$
c $Ca + 2\ HCl \rightarrow CaCl_2 + H_2$
d $2\ NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$

TASK 5 – Writing balanced equations 2

1 $4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3$
2 $C_6H_{14} + 9\frac{1}{2}\ O_2 \rightarrow 6\ CO_2 + 7\ H_2O$ or $2\ C_6H_{14} + 19\ O_2 \rightarrow 12\ CO_2 + 14\ H_2O$
3 $CH_3CH_2SH + 4\frac{1}{2}\ O_2 \rightarrow 2\ CO_2 + SO_2 + 3\ H_2O$ or $2\ CH_3CH_2SH + 9\ O_2 \rightarrow 4\ CO_2 + 2\ SO_2 + 6\ H_2O$
4 $2\ Li + 2\ H_2O \rightarrow 2\ LiOH + H_2$
5 $CaCO_3 + 2\ HNO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2$
6 $Li_2CO_3 \rightarrow Li_2O + CO_2$

7 $NH_3 + HNO_3 \rightarrow NH_4NO_3$
8 $K_2O + H_2SO_4 \rightarrow K_2SO_4 + H_2O$
9 $Ca(OH)_2 + 2\ HCl \rightarrow CaCl_2 + 2\ H_2O$
10 $3\ Zn + 2\ H_3PO_4 \rightarrow Zn_3(PO_4)_2 + 3\ H_2$
11 $2\ NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2\ H_2O + 2\ CO_2$
12 $2\ KOH + H_2SO_4 \rightarrow K_2SO_4 + 2\ H_2O$

TASK 6 – Ionic equations

1 HCl, LiOH, 1:1; H_2SO_4, $NaHCO_3$, 1:2; HNO_3, NH_3, 1:1; H_2SO_4, K_2CO_3, 1:1, HNO_3, $Sr(OH)_2$, 2:1

2 a $H^+ + OH^- \rightarrow H_2O$
b $Ag^+ + I^- \rightarrow AgI$
c $2\ H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$
d $Ca^{2+} + 2\ OH^- \rightarrow Ca(OH)_2$
e $NH_3 + H^+ \rightarrow NH_4^+$
f $H^+ + HCO_3^- \rightarrow H_2O + CO_2$
g $Ca^{2+} + SO_4^{2-} \rightarrow CaSO_4$
h $Pb^{2+} + 2\ Cl^- \rightarrow PbCl_2$
i $H^+ + OH^- \rightarrow H_2O$

TASK 7 – Significant figures & standard form

1	a 345800	b 297000	c 0.0790	d 6.10	e 0.00156	f 0.01040
2	a 2350000 (3sf)	b 0.25 (2sf)	c 13.7	d 300 (2sf)	e 0.00198 (3sf)	f 0.00031 (2sf)
3	a 0.0015	b 0.00046	c 357500	d 534	e 1030000	f 0.00835
4	a 1.64×10^{-4}	b 5.24×10^{-2}	c 1.5×10^{-8}	d 3.45×10^{4}	e 6.2×10^{-1}	f 8.7×10^{7}
5	a 0.021 (2sf)	b 6.1×10^{-5} (2sf)	c 4.0×10^{8}	d 2400	e 0.0610	f 8.00×10^{-7} (3sf)

TASK 8 – Relative formula mass

1	38.0	2	55.8	3	98.1	4	102.0	5	58.3	6	213.0
7	132.1	8	123.5	9	169.9	10	80.0	11	249.5	12	24.3
13	32.0	14	102.9	15	78.1	16	174.3	17	71.0	18	399.9

TASK 9A – Moles

1	a	2.96	b	50.3	c	0.500	d	17100	e	0.000107
2	a	355 g	b	20.4 g	c	1.08 g	d	0.264 g	e	85.8 g
3	a	0.250	b	0.250	c	0.500				
4	a	0.0500	b	0.100	c	0.150				
5	176									

TASK 9B – Avogadro number

1 1.99×10^{-23} g
2 3.011×10^{23}
3 1.81×10^{24}

TASK 10 – What equations mean

1 12 mol Na + 3 mol $O_2 \rightarrow$ 6 mol Na_2O; 0.1 mol Na + 0.025 mol $O_2 \rightarrow$ 0.05 mol Na_2O
2 5 mol Al + 7.5 mol $Cl_2 \rightarrow$ 5 mol $AlCl_3$; 0.1 mol Al + 0.15 mol $Cl_2 \rightarrow$ 0.1 mol $AlCl_3$
3 0.5 mol C_4H_{10} + 3.25 mol $O_2 \rightarrow$ 2 mol CO_2 + 2.5 mol H_2O; 20 mol C_4H_{10} + 130 mol $O_2 \rightarrow$ 80 mol CO_2 + 100 mol H_2O
4 0.5 mol NH_3 + 0.375 mol $O_2 \rightarrow$ 0.25 mol N_2 + 0.75 mol H_2O; 10 mol NH_3 + 7.5 mol $O_2 \rightarrow$ 5 mol N_2 + 15 mol H_2O

TASK 11 – Reacting mass calculations 1

1	1.0 g	2	126 g	3	120 g	4	253000 g	5	17.6 g	6	12.0 g
7	7	8	6	9	9780 g	10	1560000 g	11	0.00940 g	12	1.11 g
13	115 g	14	1650000 g	15	64.0 g	16	89.3 g				

TASK 12A – Limiting reagents 1

1	a	2 mol	b	8 mol	c	0.4 mol
2	a	2 mol	b	4 mol	c	0.4 mol
3	a	2 mol	b	10 mol	c	20 mol
4	a	1, 4 mol	b	2, 8 mol	c	0.25, 1 mol
5	a	20, 30 mol	b	0.5, 0.75 mol	c	4, 6 mol
6		4 mol	b	0.6 mol	c	12 mol
7		4 mol	b	3 mol	c	0.25 mol

TASK 12B – Limiting reagents 2

1	13.2 g	2	5.75 g	3	1.60 g	4	4.20 g	5	4.48 g
6	53.4 g								

TASK 12C – Reacting mass calculations 2

1	7.88 g	2	2690 g	3	303000 g	4	98.6 g	5	1210 g
6	1250 g	7	42.9 g						

CHALLENGE 1

1 $NaHCO_3$ = 3.51 g, Na_2CO_3 6.49 g 2 $CaCO_3$ = 40.3%, $MgCO_3$ = 59.7% 3 C_4H_8 4 26.6%

TASK 13 – Percentage yield

1	a	120 g	b	74.9%	c	reversible, product lost on isolation, other reactions take place			
2	a	701000 g	b	92.7%	3	a	510 g	b	30.0%
4	a	25.2 g	b	79.4%	5	a	529 g	b	94.4%
6	a	330 g	b	90.8%	7	a	2.40 g	b	88.4%

TASK 14 – Atom economy

1	39.3%	2	1.5%	3	45.8%	4	56.0%	5	100%	6	47.1%

7 a 320 g b 87.5% c 29.5%

d % yield compares the amount produced compared to the amount you should get, atom economy is the proportion of the mass of all the products that is the desired product

TASK 15 – Empirical & molecular formulas

1	a	CH_3	b	P_2O_3	c	SO_2	d	CH_2				
	e	CH_2O	f	C_2H_7N	g	B_3H_5	h	$C_{12}H_{22}O_{11}$				
2	a	N_2H_4	b	C_4H_{10}	c	C_5H_{10}	d	PH_3	e	C_6H_6	f	C_3H_6
3	a	3:2	b	1:2	c	5:1	d	2:5				
	e	3:4	f	5:3	g	4:5	h	4:7				
4	a	$CaBr_2$	b	$Na_2S_2O_3$	c	C_2H_7N	d	CO_2	e	NO_2		
5		$FeCl_3$	6	K_2SO_4	7	P_2O_3, P_4O_6	8	CH_2O, $C_2H_4O_2$				
9		$C_5H_{10}O$, $C_5H_{10}O$			10	x = 4, y = 2						

TASK 16 – Ideal gas equation

1	a 473 K	b	98000 Pa	c	50×10^{-6} m^3	d	223 K	e	100000 Pa	f	3.2×10^{-3} m^3
2	1.24×10^{-3} m^3	3	0.786	4	104000 Pa	5	155 K	6	71.0	7	0.00380 m^3
8	7.75×10^{-4} m^3	9	3.36 g	10	0.000538 m^3	11	4.53 m^3	12	64.1	13	483 K
14	126000 Pa										

TASK 17 – Reacting gas volumes

1 a O_2 2 dm^3, CO_2 1 dm^3 b O_2 120 cm^3, CO_2 80 cm^3
c O_2 1250 cm^3, CO_2 1000 cm^3 d O_2 5625 cm^3, CO_2 4500 cm^{3}q

2 20 cm^3 HBr left at end

3 300 cm^3 O_2, 100 cm^3 CO_2, total 400 cm^3 gas at end

4 4 dm^3 O_2, 4 dm^3 H_2O, 4 dm^3 SO_2, total 12 dm^3 gas

5 C_8H_{18}

CHALLENGE 2

1 44.0 2 3.21 : 1, 130.5 g 3 NS 4 C_2H_4 5 515 ms^{-1} 6 C_3H_8

Calculations CHECK-UP

1 a $Zn(NO_3)_2$ b Pb c Cr_2O_3 d $(NH_4)_2SO_4$
e P_4 f N_2 g $Ba(OH)_2$ h $Al_2(SO_4)_3$

2 H_2SO_4, KOH, 1:2; HCl, $KHCO_3$, 1:1; HNO_3, NH_3, 1:1; HCl, $ZnCO_3$, 2:1

3 a $H^+ + OH^- \rightarrow H_2O$ b $Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4$
c $H^+ + NH_3 \rightarrow NH_4^+$ d $H^+ + HCO_3^- \rightarrow H_2O + CO_2$

4 a average mass of an atom, relative to 1/12th mass of ^{12}C atom b it is the agreed standard
c mixture of other isotopes

5 a H_2, NH_3 = 3.33 b H_2, NH_3 = 3.33
c N_2, NH_3 = 20.0 d H_2, NH_3 = 0.033

6 3.10×10^{-4} m^3 7 8.21×10^{-3}

8 a volume of CO_2 = 57.1 cm^3, total = 128.5 cm^3 b volume of CO_2 = 200 cm^3, total = 350 cm^3
c volume of CO_2 = 229 cm^3, total = 314 cm^3

9 2.00×10^{-3} m^3 10 1.64 m^3

11 a 40, 60 b 40.0, 20.0 c 5.84, 8.76

12 193.5 g 13 9.39 g

14 a 1250 g b i 96% ii reversible, product lost on isolation, other reactions iii 100%

15 a 529 g b 94.5% c 52.9% 16 7

17 a 0.05, 0.05, 1.22×10^{-3} m^3, 4.07 g b 1.30 g, 2.77 g

TASK 18 – Solution calculations

1 a 0.1 b 250 c 0.0025

2 a 0.2 mol dm^{-3}, 7.3 g dm^{-3} b 2.5 mol dm^{-3}, 245.3 g dm^{-3} c 0.0525 mol dm^{-3}, 2.10 g dm^{-3}

3 a 0.05 dm^3 b 0.001 dm^3

4 0.0269 mol dm^{-3}, 4.61 g dm^{-3} 5 0.0752 mol dm^{-3}, 3.01 g dm^{-3} 6 0.0075 dm^3

7 0.0500 mol dm^{-3}, 7.10 g dm^{-3} 8 1.13 g 9 79.9 dm^3

10 87.8 11 2 12 A_r = 39.1, K

TASK 19 – Back titration calculations

1 87.7% 2 90.8% 3 0.05 mol, 4.22 g

4 0.606 g 5 4.68 g

CHALLENGE 3

1	9.67%	2	**A** Si_2OCl_6

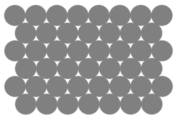
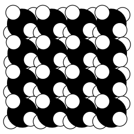
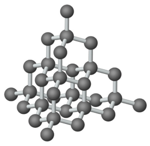
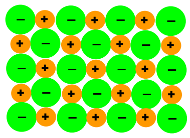
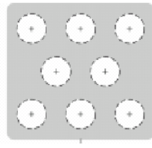
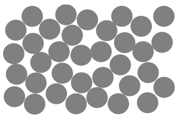
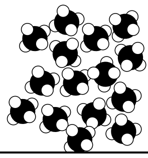
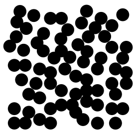
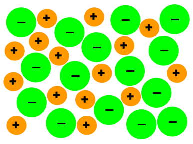
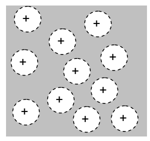
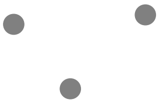
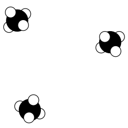
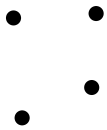
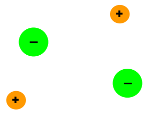
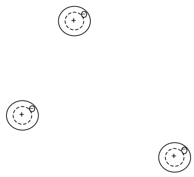
Calculation Allsorts

1	$C_5H_{11}NO$	2	$C_{11}H_{14}O_2$, $C_{11}H_{14}O_2$	3	526 g	4	2.71 g	5	5.00 dm^3
6	0.0241 mol dm^{-3}	7	234.9	8	3.21%	9	55.0%		

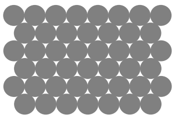
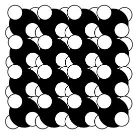
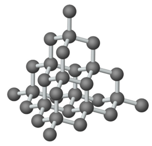
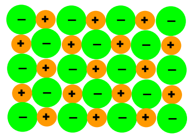
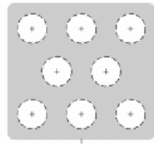
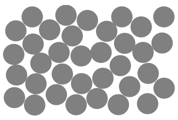
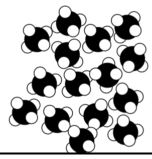
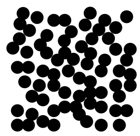
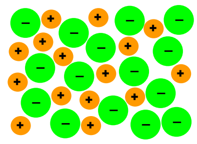
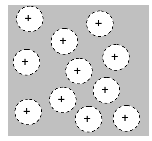
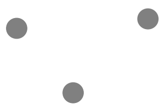
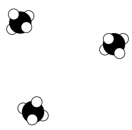
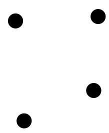
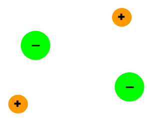
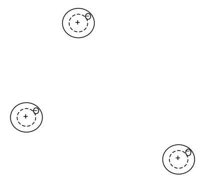
10 $10\ Al + 6\ NH_4ClO_4 \rightarrow 3\ N_2 + 9\ H_2O + 6\ HCl + 5\ Al_2O_3$

Structure and Bonding

STRUCTURE TYPES

	Monatomic	Simple molecular	Giant covalent	Ionic	Metallic
Substances	Group 0 elements	Elements: H_2 O_2 N_2 F_2 Cl_2 Br_2 I_2 S_8 P_4 Compounds: non-metal with non-metal	*sometimes incorrectly called macromolecular* Elements: Si, diamond, graphite, graphene Compounds: SiO_2	Compounds: metal with non-metal	Elements: metals
What the structure is	Individual atoms with very weak forces between them.	Lots of individual molecules with weak forces between the molecules. (the atoms within molecules are joined by covalent bonds)	Lattice structure in which all atoms are joined together in a giant network by covalent bonds.	Lattice structure of positive and negative ions. The ions are held together by the strong attraction between the + and – ions (this +/- attraction is known as ionic bonding, although it is just an electrostatic attractive force).	Lattice structure of metal atoms where the outer shell electrons from each atom are delocalised. There is a strong attraction between the positive nucleus of the atoms and the cloud of negative delocalised electrons (this is known as metallic bonding)
Bonding					
Solid					
Liquid					
Gas					
Formula (molecular)		Gives number of atoms of each type in one molecule: e.g. glucose $C_6H_{12}O_6$ each molecule contains 6C, 12H and 6O atoms			
Formula (empirical)	Just the symbol e.g. Ar	Gives ratio of atoms in substance e.g. glucose CH_2O ratio of C:H:O atoms is 1:2:1	Gives ratio of atoms in substance e.g. SiO_2 ratio of Si:O atoms is 1:2	Gives ratio of ions in substance: e.g. $MgCl_2$ ratio of Mg^{2+}:Cl^- ions is 1:2	Just the symbol e.g. Fe
Melting and boiling points					
	Higher melting / boiling points occur when:	*Higher melting / boiling points occur when:*	*Higher melting / boiling points occur when:*	*Higher melting / boiling points occur when:*	*Higher melting / boiling points occur when:*
Conductivity					
Solubility (aq)					

STRUCTURE TYPES

	Monatomic	Simple molecular	Giant covalent	Ionic	Metallic
Substances	Group 0 elements	Elements: H_2 O_2 N_2 F_2 Cl_2 Br_2 I_2 S_8 P_4 Compounds: non-metal with non-metal	*sometimes incorrectly called macromolecular* Elements: Si, diamond, graphite, graphene Compounds: SiO_2	Compounds: metal with non-metal	Elements: metals
What the structure is	Individual atoms with very weak forces between them.	Lots of individual molecules with weak forces between the molecules. (the atoms within molecules are joined by covalent bonds)	Lattice structure in which all atoms are joined together in a giant network by covalent bonds.	Lattice structure of positive and negative ions. The ions are held together by the strong attraction between the + and – ions (this +/- attraction is known as ionic bonding, although it is just an electrostatic attractive force).	Lattice structure of metal atoms where the outer shell electrons from each atom are delocalised. There is a strong attraction between the positive nucleus of the atoms and the cloud of negative delocalised electrons (this is known as metallic bonding)
Bonding	none	covalent (within molecules)	covalent	ionic	metallic
Solid					
Liquid					
Gas					
Formula (molecular)		Gives number of atoms of each type in one molecule: e.g. glucose $C_6H_{12}O_6$ each molecule contains 6C, 12H and 6O atoms			
Formula (empirical)	Just the symbol e.g. Ar	Gives ratio of atoms in substance e.g. glucose CH_2O ratio of C:H:O atoms is 1:2:1	Gives ratio of atoms in substance e.g. SiO_2 ratio of Si:O atoms is 1:2	Gives ratio of ions in substance: e.g. $MgCl_2$ ratio of Mg^{2+}:Cl^- ions is 1:2	Just the symbol e.g. Fe
Melting and boiling points	VERY LOW Very weak forces between atoms	LOW Weak forces between the molecules (Note – the atoms within the molecules are held together by strong covalent bonds, but these DO NOT break when molecules melt/boil)	VERY HIGH Need to break many strong covalent bonds	HIGH Strong electrostatic attraction between positive and negative ions	HIGH Strong electrostatic attraction between positive metal ions and delocalised negative electron
	Higher melting / boiling points occur when: the heavier the atoms, the stronger the forces between the atoms	*Higher melting / boiling points occur when:* the stronger the intermolecular forces	*Higher melting / boiling points occur when:* the stronger the covalent bonds	*Higher melting / boiling points occur when:* the smaller the ions and the higher the charge on the ions, the stronger the attraction between the positive and negative ions	*Higher melting / boiling points occur when:* the smaller the ions, the higher the charge on the ions, and/or the more delocalised electrons, the stronger the metallic bonding
Conductivity	Do not conduct contain no mobile ions or electrons	Do not conduct contain no mobile ions or electrons	*Diamond, Si, SiO_2* Do not conduct contain no mobile ions or electrons *Graphite, graphene* Conduct as delocalised electrons carry charge through structure	*Solids* Do not conduct as ions are not mobile *Liquids and solutions* Conduct mobile ions carry charge through structure	*Conduct* as delocalised electrons carry charge through structure
Solubility (aq)	Insoluble	Insoluble (usually)	Insoluble	Soluble (usually)	Insoluble

THE NATURE OF BONDS

1) TYPES OF BOND

Type of bonding	Ionic	Covalent	Metallic
Nature of bonding	Electrostatic attraction between positive and negative ions	Shared pair of electrons between atoms	Attraction in a lattice between the positive nuclei of the metal atoms and the negative delocalised outer shell electrons.
Types of structure which have this type of bonding	IONIC	SIMPLE MOLECULAR GIANT COVALENT	METALLIC
Strength of bonds	The smaller the ions and the greater the charge on the ions, the stronger the attraction between the positive and negative ions (usually).	The shorter the bond, the stronger the bond (usually). Double bonds are stronger than single bonds, while triple bonds are stronger than double bonds.	Stronger metallic bonding is caused by: • smaller atoms • more delocalised electrons

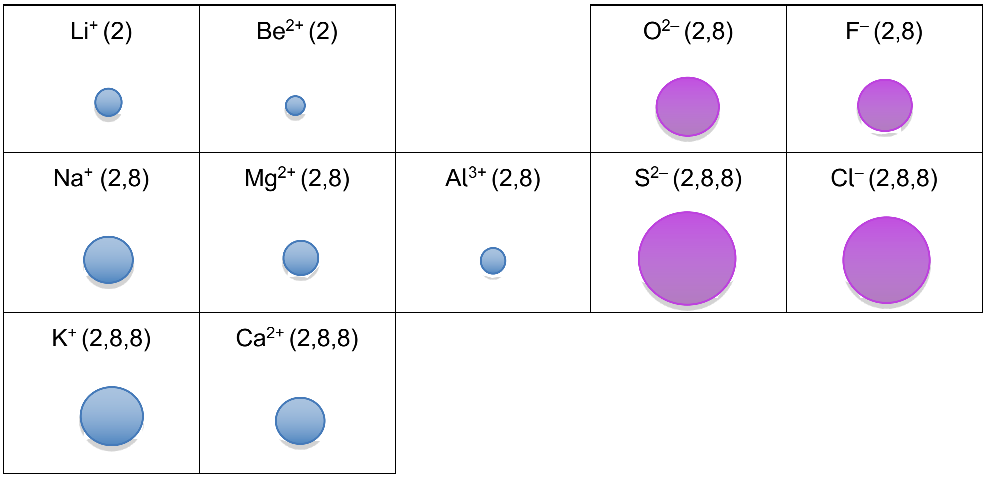
There are also three types of forces between molecules (van der Waals' forces, dipole-dipole attractions and hydrogen 'bonds'). These are **NOT** bonds because they are too weak.

2) IONIC BONDING

Strength of ionic bonding

The greater the charge on the ions, the stronger the ionic bonding. The smaller the ions, the stronger the ionic bonding (remember that ions get bigger down a group).

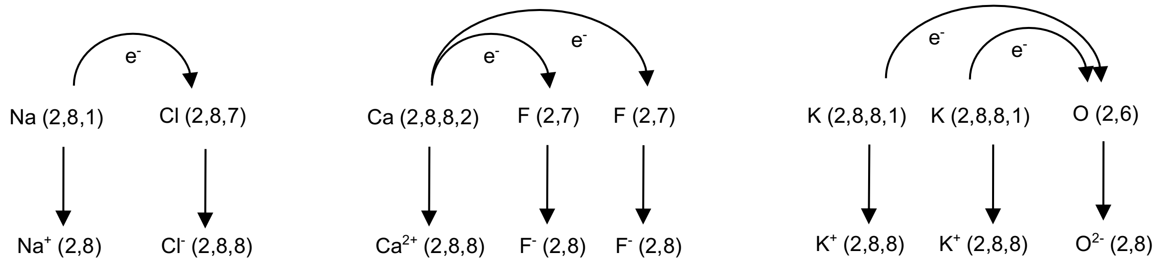
Relative size of some common ions:



TASK 1 Predict (with reasons) which one of each pair of ionic compounds will have the higher melting point by considering size and charge of ions.	
a) sodium chloride v potassium chloride	
b) sodium fluoride v magnesium fluoride	
c) aluminium oxide v sodium oxide	

Formation of ions

Ions can be formed when a **metal** reacts with a **non-metal**. The metal atoms lose electrons to form positive ions while non-metal atoms gain electrons to form negative ions (both obtaining full outer shells). For example (in *over-simplified* GCSE terms)

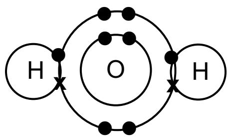
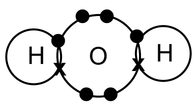
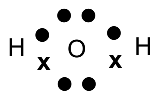


makes sodium chloride (NaCl) makes calcium fluoride (CaF_2) makes potassium oxide (K_2O)

Remember that ionic bonds are the attraction between positive and negative ions (and has nothing to do with the transfer of electrons!).

3) COVALENT BONDS

- Covalent bonds can be formed when a **non-metal** reacts with a **non-metal**.
- The atoms share electrons to obtain stable electron structures ("full outer shells").
- Two shared electrons make a **single bond**, four shared electrons make a **double bond**, and six shared electrons makes a **triple bond**.
- For example in water:

all the electrons	**dot-cross diagrams**		**stick diagram**
	outer shell electrons only	outer shell electrons only (without shell circles)	
			H—O—H

Drawing "dot-cross" diagrams

1) Draw a stick diagram (use this table to give some help).

Atoms	Group 4 atoms	Group 5 atoms	Group 6 atoms	Group 7 atoms	H
Common* number of covalent bonds	4	3	2	1	1

* Note that elements in period 3 and beyond can fit more than 8 electrons in their outer shell (due to availability of d orbitals in the shell). For example, P and S atoms can have more than electrons in the third shell (extra electrons occupy 3d orbitals).

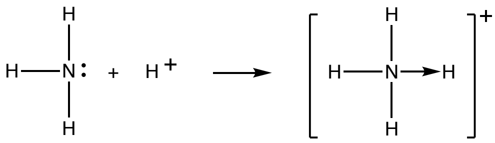
2) Re-draw the molecule without the sticks.
3) Draw a dot and a cross instead of each stick (i.e. ●**x** for a single bond, ●**x**●**x** for a double bond, etc.). Dots and crosses represent electrons from different atoms.
4) Work out how many electrons there are in the outer shell of each atom (e.g. atoms in Group 7 have 7 electrons in their outer shell), and then add in any of these electrons that are not already drawn on in the covalent bonds. (If the species has an electric charge, then that means that there are additional electrons (negative charge) or fewer electrons (positive charge)

TASK 2 Draw stick and then dot-cross diagrams for each of the following molecules and ions.	
1) F_2	11) NH_4^+
2) NH_3	12) $AlCl_4^-$
3) O_2	13) BCl_3
4) H_2S	14) SF_6
5) CH_4	15) $BeCl_2$
6) H_2	16) $BeCl_4^{2-}$
7) PF_3	17) IF_4^+
8) PF_5	18) C_2H_2
9) SO_2	19) SO_4^{2-}
10) CO_2	20) NO_3^-

Co-ordinate (dative covalent) bonds

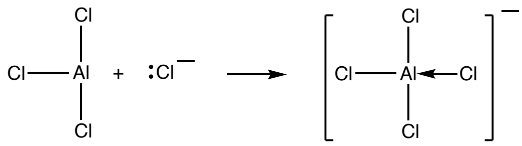
- A co-ordinate bond (dative covalent bond) is one where both the electrons come from the same species.
- They are often drawn with → rather than –, with the arrow showing the direction in which the electrons are donated.
- However, once formed, co-ordinate bonds are identical to other covalent bonds. For example, all four bonds in NH_4^+ are identical (although one was formed as a co-ordinate bond with both electrons being donated from the lone pair on the N to the H^+).

e.g. NH_4^+ $NH_3 + H^+ \longrightarrow NH_4^+$



lone pair of electrons donated from N of NH_3 to H^+

e.g. $AlCl_4^-$ $AlCl_3 + Cl^- \longrightarrow AlCl_4^-$



lone pair of electrons donated from Cl^- to Al of $AlCl_3$

TASK 3 Draw a diagram to show the formation of a co-ordinate bond in the following. In each case, state how the bond forms.	
a) BF_3 with F^- to form BF_4^-	
b) PH_3 with H+ to form PH_4^+	

4) METALLIC BONDS

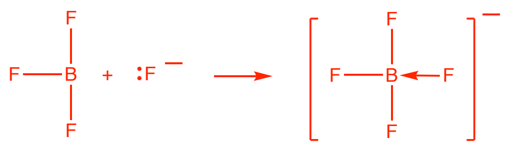
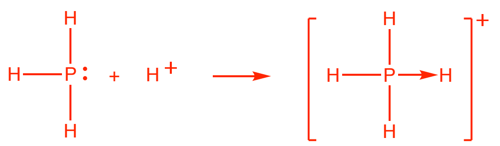
The stronger the attraction between the nucleus of the metal atoms and the delocalised electrons, the stronger the metallic bonding. The smaller the metal atoms and the more delocalised outer shell electrons there are, the stronger the metallic bonding.

TASK 4 Predict (with reasons) which one of each pair of metals will have the higher melting point.	
a) sodium v potassium	
b) sodium v magnesium	
c) potassium v aluminium	

THE NATURE OF BONDS

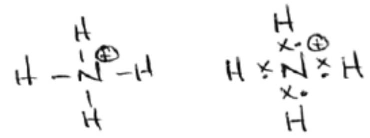
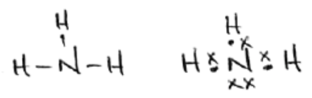
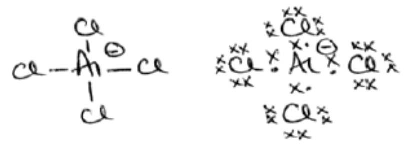
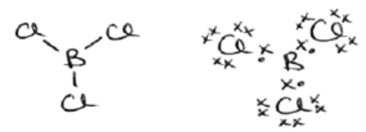
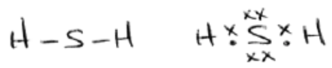
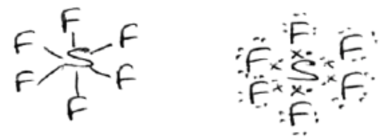
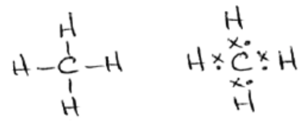
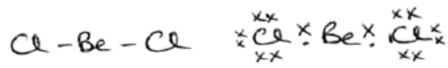
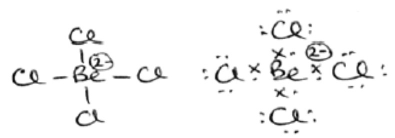
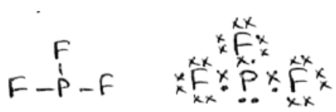
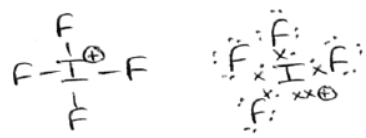
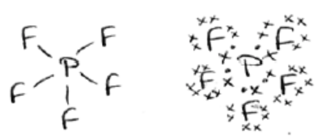
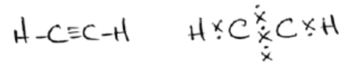
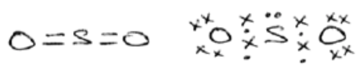
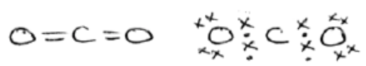
TASK 1 Predict (with reasons) which one of each pair of ionic compounds will have the higher melting point by considering size and charge of ions.	
a) sodium chloride v potassium chloride	sodium chloride • Na^+ ions are smaller than K^+ ions • so there will be a stronger attraction between positive and negative ions
b) sodium fluoride v magnesium fluoride	magnesium fluoride • Mg^{2+} ions are smaller <u>and</u> more highly charged than Na^+ ions • so there will be a stronger attraction between positive and negative ions
c) aluminium oxide v sodium oxide	aluminium oxide • Al^{3+} ions are smaller <u>and</u> more highly charged than Na^+ ions • so there will be a stronger attraction between positive and negative ions

TASK 2	See over the page

TASK 3 Draw a diagram to show the formation of a co-ordinate bond in the following. In each case, state how the bond forms.	
a) BF_3 with F^- to form BF_4^-	 lone pair of electrons is donated from F^- to B on BF_3
b) PH_3 with H+ to form PH_4^+	 lone pair of electrons is donated from P on PH_3 to H^+

TASK 4 Predict (with reasons) which one of each pair of metals will have the higher melting point.	
a) sodium v potassium	sodium • Na atoms are smaller than K atoms • so there will be a stronger metallic bonding
b) sodium v magnesium	magnesium • Mg atoms are smaller than Na atoms • magnesium has more delocalised electrons than sodium • so there will be a stronger metallic bonding
c) potassium v aluminium	aluminium • Al atoms are smaller than K atoms • aluminium has more delocalised electrons than potassium • so there will be a stronger metallic bonding

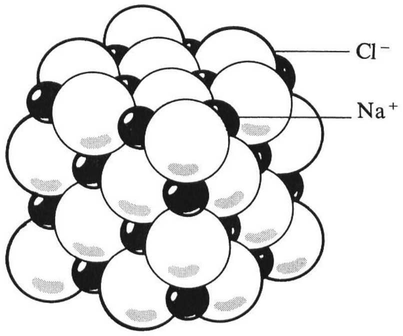
TASK 2 Draw stick and then dot-cross diagrams for each of the following molecules and ions.

1) F_2		11) NH_4^+	
2) NH_3		12) $AlCl_4^-$	
3) O_2		13) BCl_3	
4) H_2S		14) SF_6	
5) CH_4		15) $BeCl_2$	
6) H_2		16) $BeCl_4^{2-}$	
7) PF_3		17) IF_4^+	
8) PF_5		18) C_2H_2	
9) SO_2		19) SO_4^{2-}	
10) CO_2		20) NO_3^-	

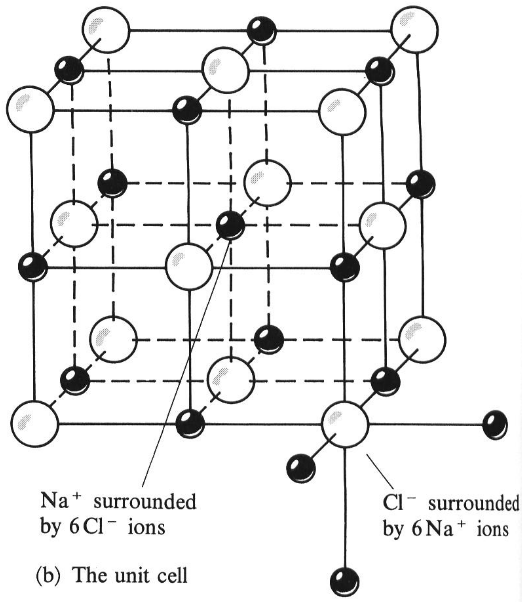
There are five types of structure: monatomic, simple molecular, ionic, giant covalent and metallic.

Monatomic substances are all gases at room temperature, but there are examples of substances with each of the other four types of structure that are solids. This looks closely at one example of a solid with each of the four remaining types of structure to reinforce and test your ideas of structure and bonding. Each of the four examples occur as crystalline solids at room temperature – hence this refers to a discussion of the four types of crystal.

1) IONIC (e.g. sodium chloride)



(a) The arrangement of ions



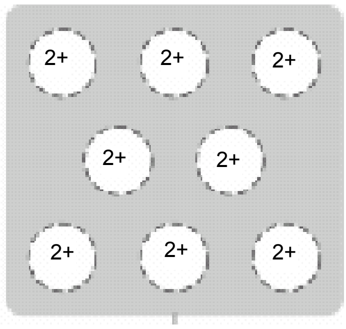
(b) The unit cell

This shows very well how the ions are packed close together

Note that this diagram only shows the centre of ions to help us see how the ions are arranged in a cubic shape. Although it may look like it, there are NO covalent bonds between the ions.

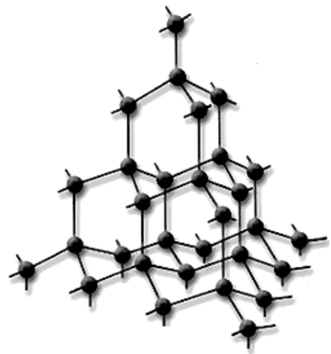
	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Strength		
Solubility		

2) METALLIC (e.g. magnesium)



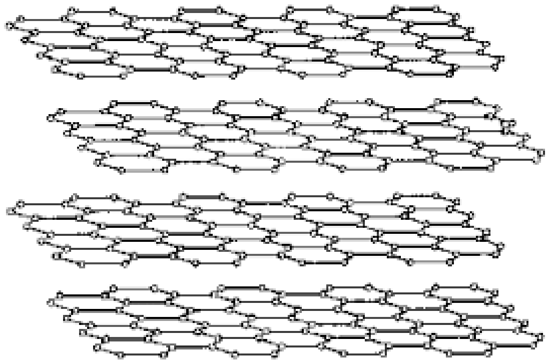
	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Strength		
Solubility		

3) GIANT COVALENT (e.g. diamond)



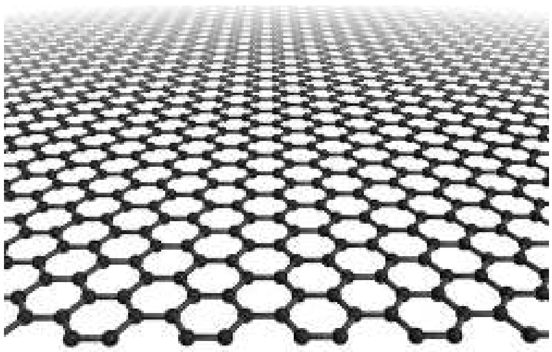
	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Strength		
Solubility		

graphite



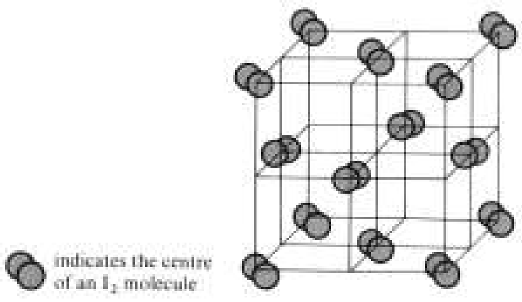
	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Strength		
Solubility		

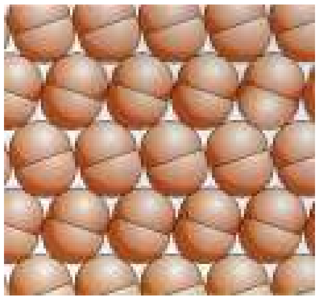
graphene



	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Strength		
Solubility		

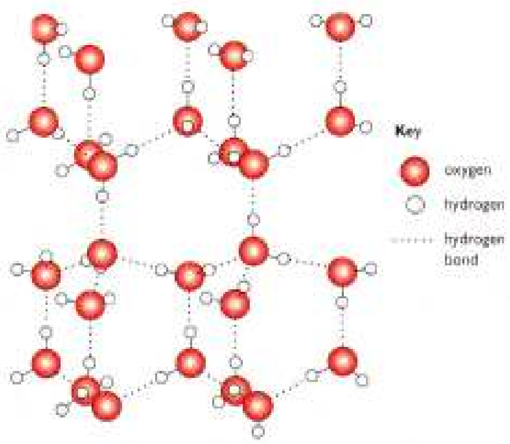
4) SIMPLE MOLECULAR (e.g. iodine)





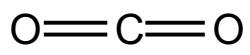
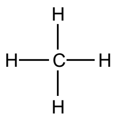
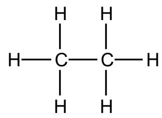
	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Strength		
Solubility		

e.g. ice



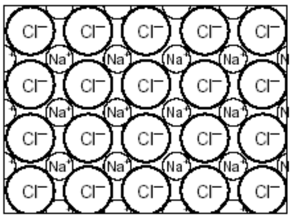
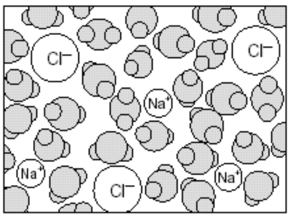
	Property	Explanation
Melting & boiling points		
Electrical conductivity		
Density of ice v water		

THE TRUTH ABOUT STRUCTURE & BONDING

SIMPLE MOLECULAR substances			*If false, what is wrong?*
 molecule of carbon dioxide, CO_2	 molecule of methane, CH_4	 molecule of ethane, C_2H_6	
1	T F	Methane is a gas at room temperature because the bonds between the atoms are weak.	
2	T F	Ethane has a higher boiling point than methane because there are more bonds to break.	
3	T F	Carbon dioxide has a higher boiling point than methane because its atoms are held together by double bonds rather than single bonds.	

GIANT COVALENT substances			*If false, what is wrong?*
4	T F	Diamond has a high melting point because there are strong covalent bonds between its molecules	
5	T F	Diamond has a high melting point because the atoms are all joined by covalent bonds in a lattice	

METALLIC structures			*If false, what is wrong?*
6	T F	Copper has a high melting point because there are strong forces of attraction between the nucleus of the copper atoms and the delocalised outer shell electrons	
7	T F	Copper has a high melting point because there are strong forces of attraction between the nucleus of each copper atom and its electrons	
8	T F	The metal conducts electricity because there is a delocalised electron	
9	T F	Copper has a high melting point because there are lots of strong covalent bonds to break	
10	T F	Copper can be bent because the layers of copper atoms can slide relative to each other	

IONIC structures			*If false, what is wrong?*
Ionic structures as a SOLID		 Sodium chloride as a solid, NaCl(s)	
11	T F	Each molecule of sodium chloride contains one sodium ion and one chloride ion.	
12	T F	Each sodium ion is attracted to one chloride ion.	
13	T F	The ions exist in pairs containing one sodium ion and one chloride ion.	
14	T F	Each sodium ion is bonded ionically to one chloride ion, and then to others by attractive forces.	
15	T F	There is a bond between the ions in each molecule, but no bonds between molecules.	
16	T F	There are no molecules shown in the diagram.	
17	T F	An ionic bond is when one atom donates an electron to another atom.	
18	T F	A sodium ion can only form one ionic bond because it only has one electron in its outer shell.	
19	T F	The sodium ions and chloride ions are not joined to each other, but are attracted to each other by electrostatic attraction.	
20	T F	Each sodium ion is attracted to all the chloride ions surrounding it.	
Ionic structures as a solution		 Sodium chloride as a solution, NaCl(aq)	
21	T F	The ions are separated	
22	T F	The sodium chloride molecules break apart when they dissolve	
23	T F	The sodium and chloride ions move around in Na^+ Cl^- pairs.	
24	T F	The solution conducts electricity because electrons can pass through the solution	

THE TRUTH ABOUT STRUCTURE & BONDING 1

1	False	The bonds do not break when it changes state – it is a gas due to weak forces between the molecules
2	False	The bonds do not break when it changes state – ethane has a higher boiling point as there are stronger forces between the molecules
3	False	The bonds do not break when it changes state – CO_2 has a higher boiling point as there are stronger forces between the molecules
4	False	There are no molecules – it is a giant covalent lattice
5	True	
6	True	
7	False	The copper ions are all positive so repel! The attraction is between the positive copper ions and the delocalised outer shell electrons
8	True	but it is only the outer shell electrons that are free to move and there are lots of them from all the atoms (not just one electron!)
9	False	There are no covalent bonds – the attraction that is overcome is between the positive copper ions and delocalised outer shell electrons
10	True	
11	False	It has a lattice of positive and negative ions – there are no molecules!
12	False	Each ion is attracted to all the nearby ions of opposite charge
13	False	It has a lattice of positive and negative ions
14	False	It has a lattice of positive and negative ions – ions of opposite charge are attracted to each other – this attraction is what ionic bonds are!
15	False	It has a lattice of positive and negative ions – there are no molecules!
16	True	
17	False	Ions can be formed by electron transfer – this is NOT an ionic bond. Ionic bonds are the attraction between positive and negative ions
18	False	Each ion is attracted to all the nearby ions of opposite charge
19	True	
20	True	
21	True	
22	False	It has made of positive and negative ions – there are no molecules!
23	False	The ions separate – they actually form new attractions to the water molecules instead of each other
24	False	It is because ions can move (not electrons)