Year 12 Preparation: Support

Year 12 Chemistry Students: This is the support document to help with the question booklet.

This must be brought to your first Chemistry lesson in September

Structure:

Information

Example

Answers

Within the first two weeks back you will have a test on the below Topics;

The three types of bonding and the properties arising from them	Calculating percentage by mass
Electronic structure and configuration (s, p, d, etc.)	No. of protons, neutrons and electrons
Calculating Relative Atomic Mass (Shown as M_r at A level)	Yields
Balancing Equations	Manipulating equations regarding number of moles and Avogadro's Number
Calculating Empirical Formula	Calculating moles

Topics highlighted in red are A Level Chemistry and new content. Those in blue are covered in GCSE.

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Structure and bonding

Video links

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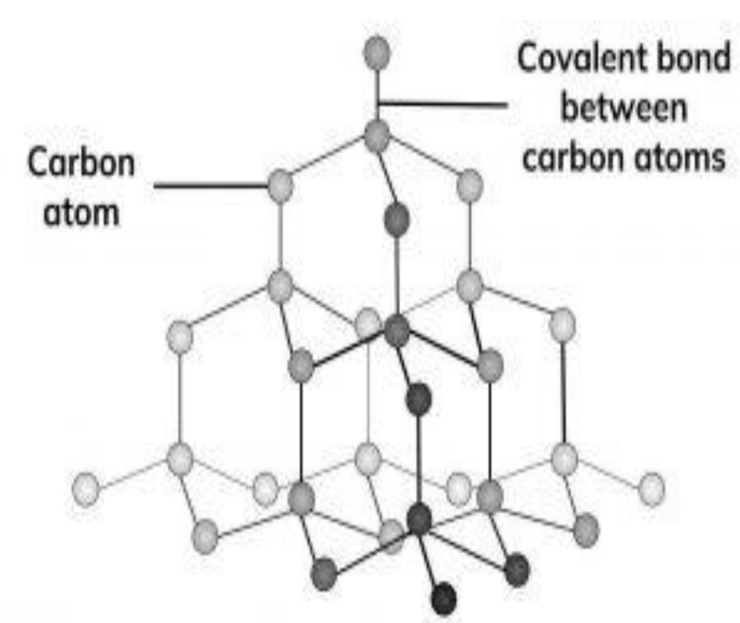
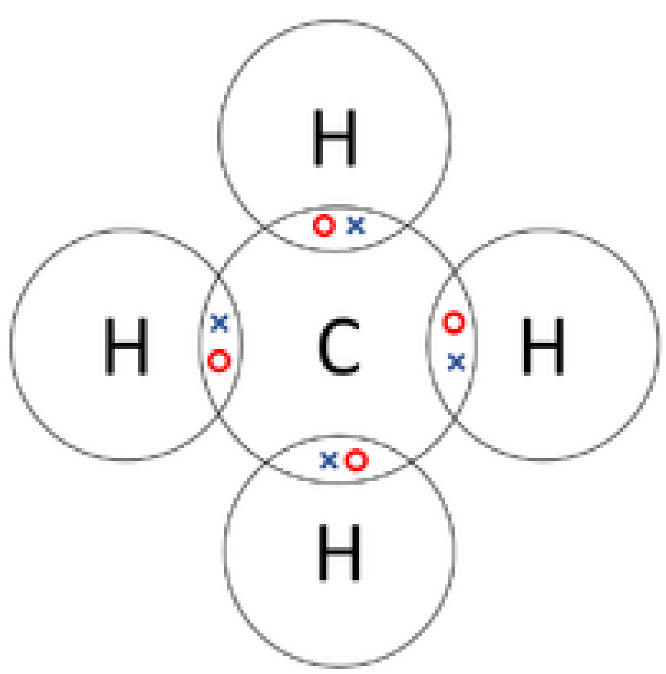
Bonding, structure and the properties of matter walkthrough

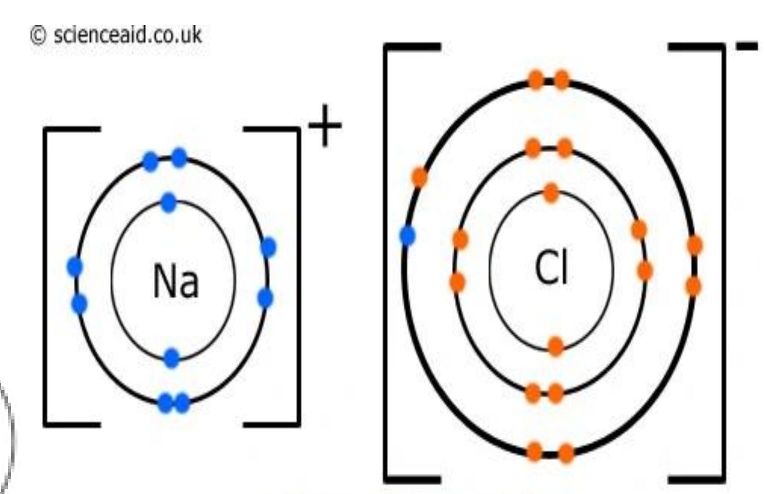
Identifying bonding and structure

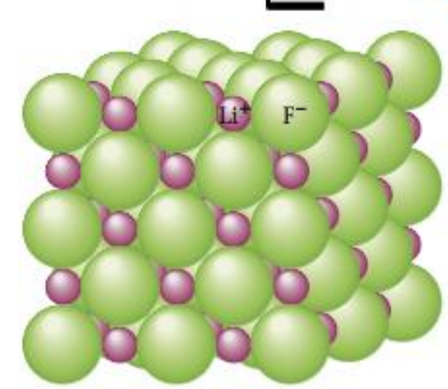
Does it contain a metal?

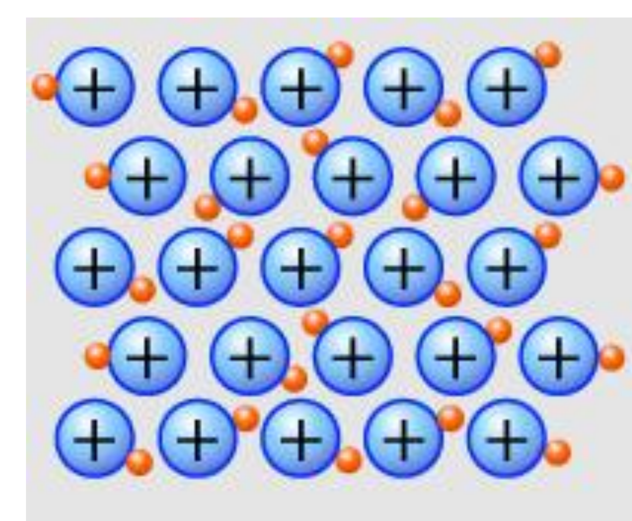
- No → Does it have a formula? E.g. O_2
  - No → Giant covalent bond e.g. diamond
  - Yes → Simple covalent bond e.g. Hydrogen
- Yes → Does it contain metal ONLY?
  - No → Ionic bond
  - Yes → Metallic bond









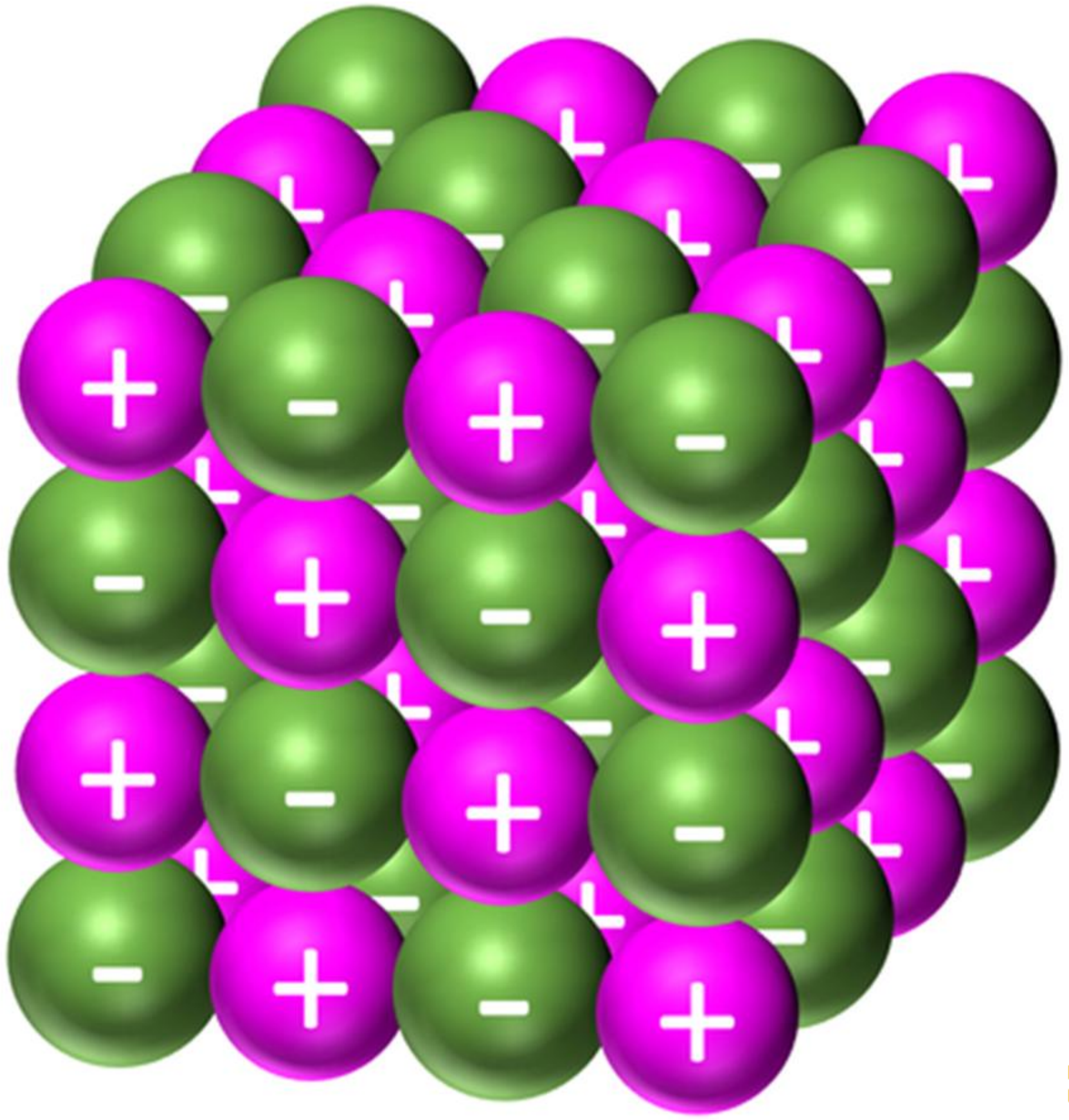


Giant Ionic

- Metal and Non Metal
- Giant uniform lattice of oppositely charged ions held together by STRONG ELECTROSTATIC FORCES of attraction acting in all directions.

Properties

- High melting/boiling points due to strong forces of attraction
- Can conduct ONLY when IONS can move freely (when a liquid)
- High density

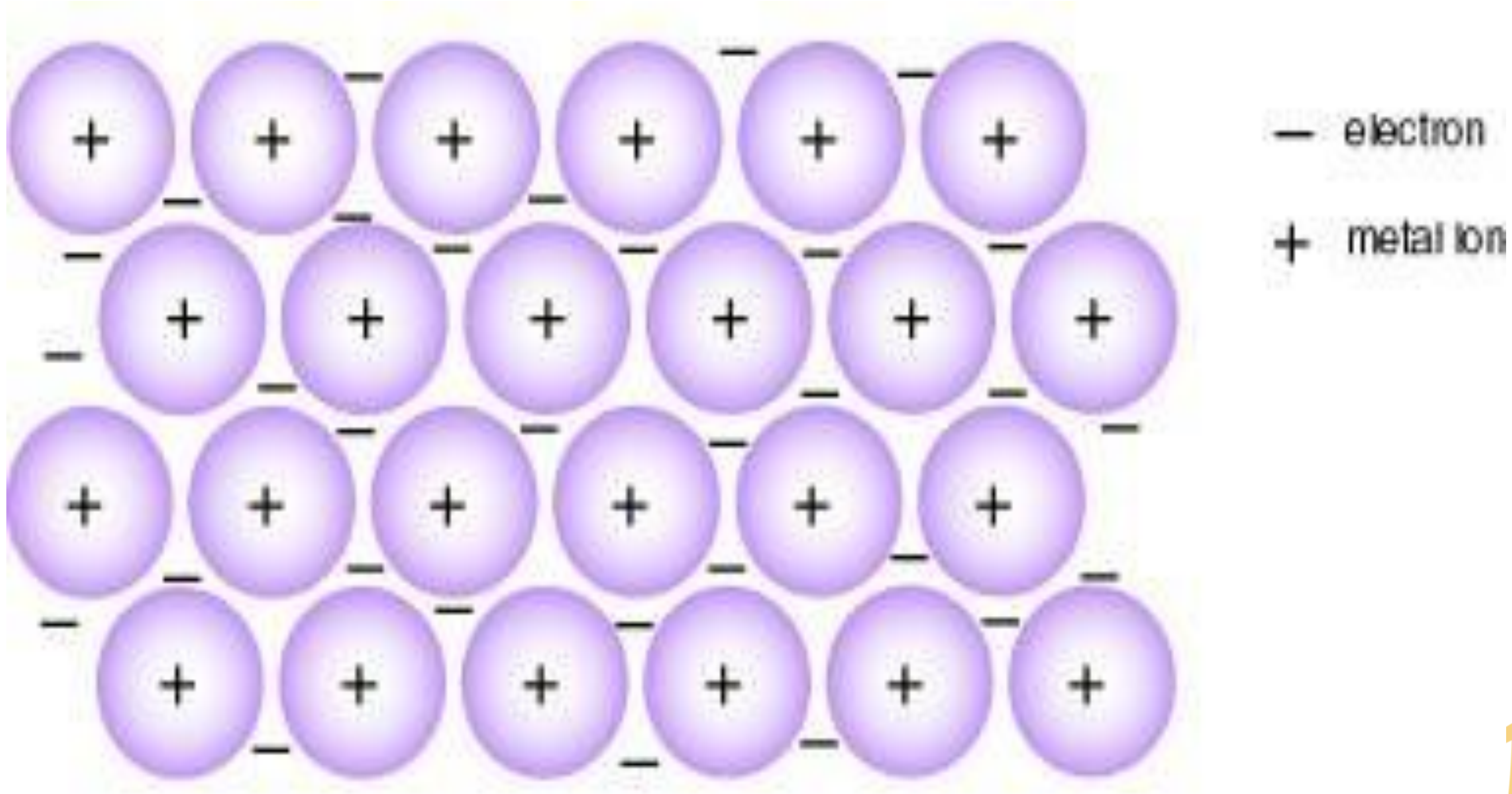


Giant Metallic

- Metals
- Giant uniform lattice of positively charged ions surrounded by a sea of delocalised electrons held together by STRONG ELECTROSTATIC FORCES of attraction

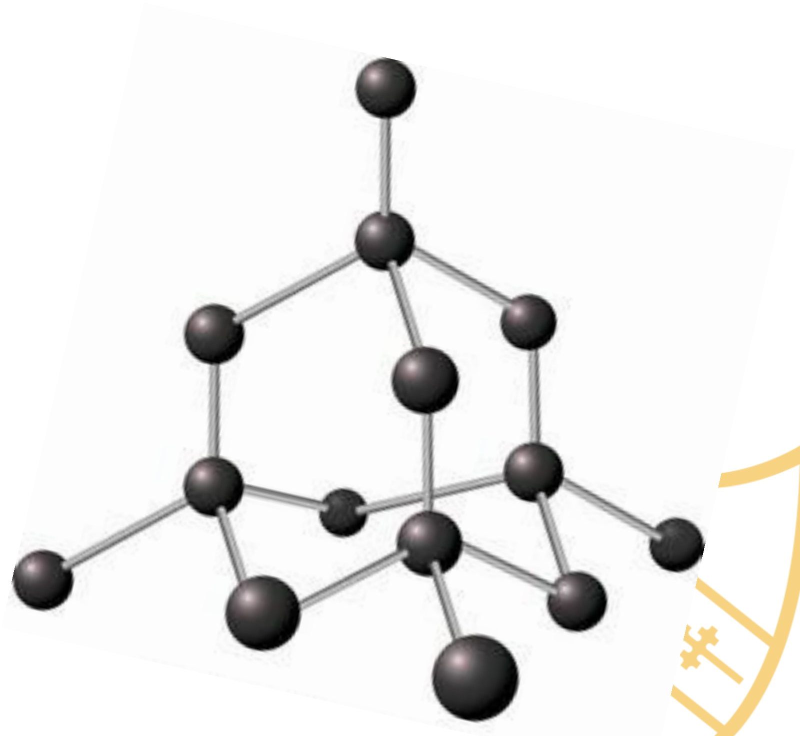
Properties

- High melting/boiling points due to strong forces of attraction
- Can conduct because ELECTRONS are the charge carriers and can move through the whole structure
- High density



Giant Covalent

- Non Metals e.g. C (diamond and graphite) or Si (SiO_2)
- Giant uniform lattice of atoms covalently bonded together
- E.g. Diamond/Graphite/SiO_2
- High melting points as there are lots of strong covalent bonds that require a large amount of energy to overcome

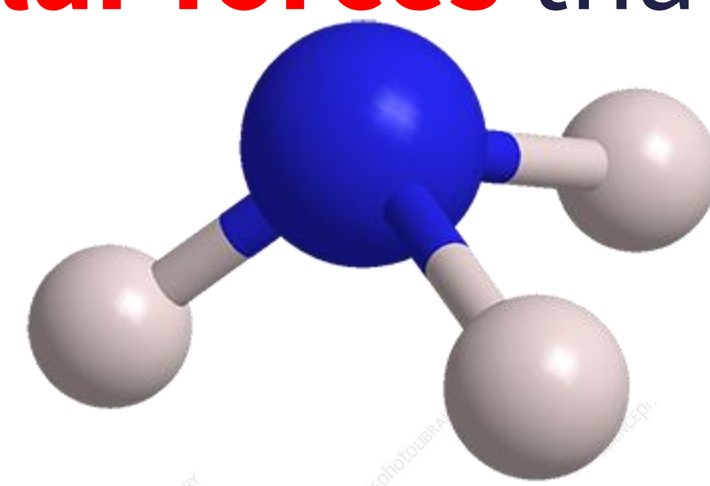


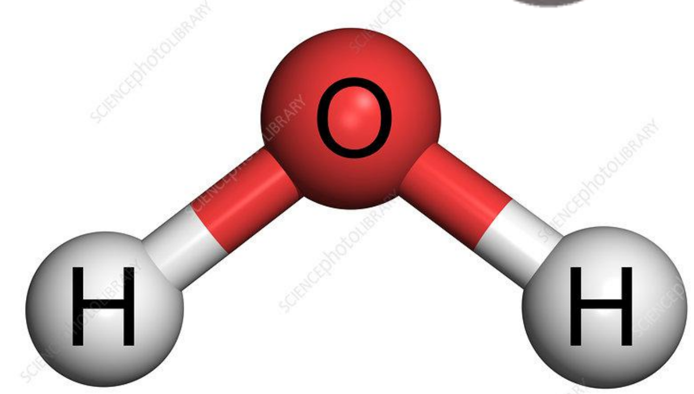
Simple Molecules

- Involves groups of non-metal atoms
- Has a discrete formula e.g. H_2O, CO_2, NH_3, H_2, C_2H_4

Properties

- Low melting a boiling point - They are **simple molecules** with **weak intermolecular forces** that are **easily overcome**
- **Don't** conduct electricity
- Have a low density



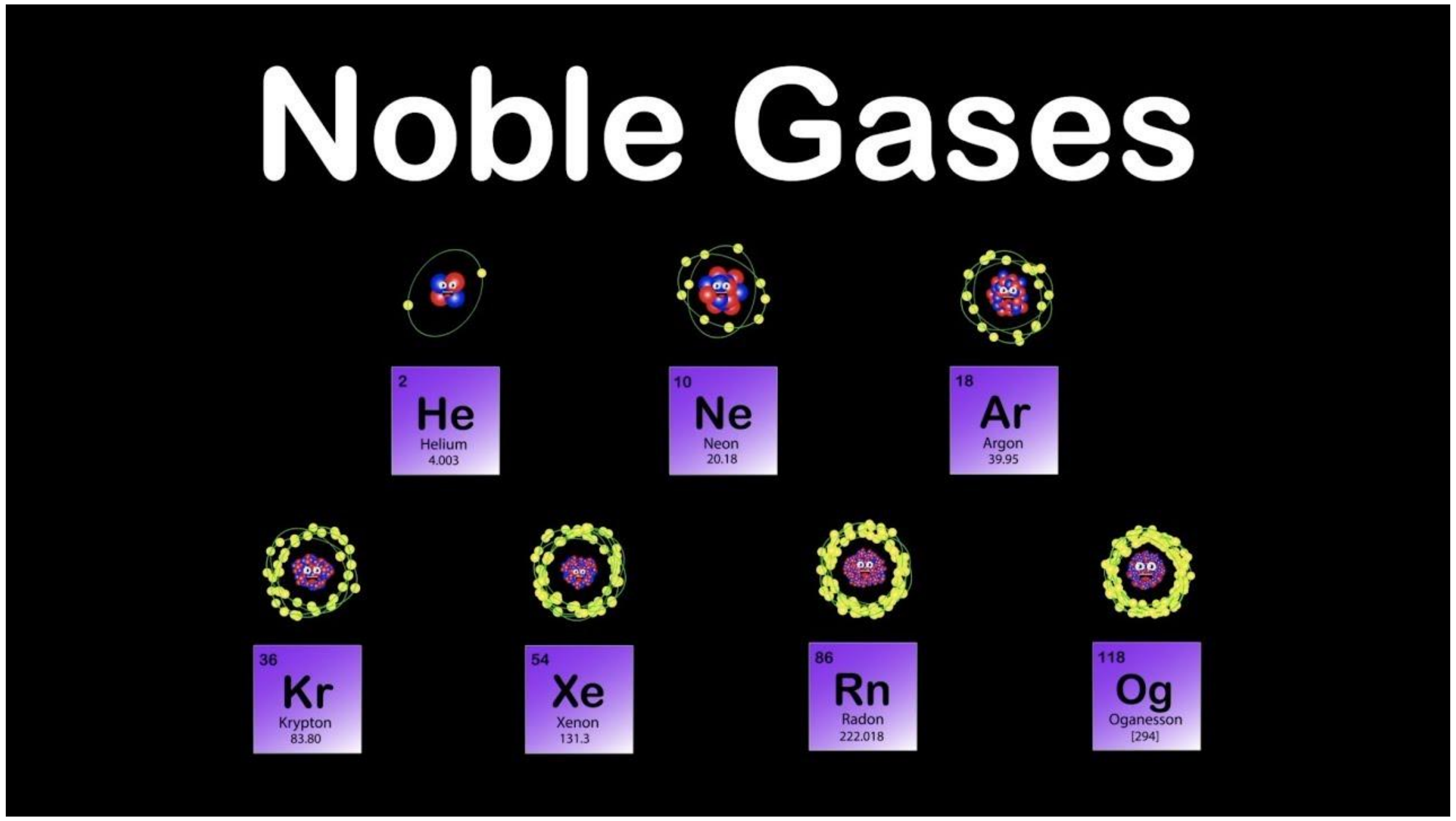


Monatomic: Noble Gases

- Group 0 elements
- Travel on their own
- Full outer shell

Properties

- Unreactive
- Don’t conduct electricity
- Low density



Proton, neutron and electron number

Video links

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Proton, neutron and electron number walkthrough

The periodic table is your *friend*! ***Use it***

Atomic number = smaller value

Atomic mass = larger value

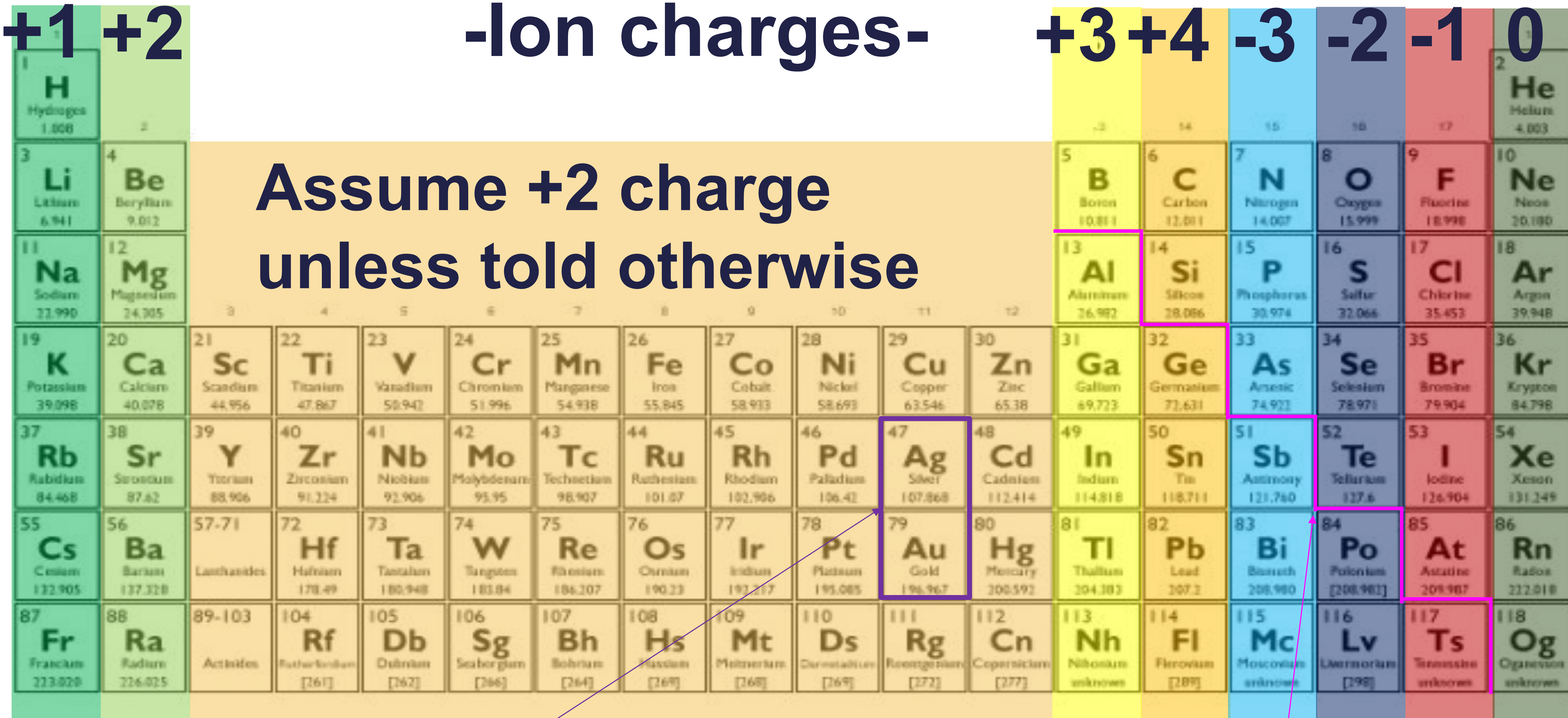
-Group-

1 2 3 4 5 6 7 0

-Ion charges-

+1 +2 +3 +4 -3 -2 -1 0

Assume +2 charge unless told otherwise



Has a charge of +1 only

Metal to non-metal dividing line

Atomic number and mass number

Lithium looks like this on the table, this is known as the 'nuclear symbol.'

This is the **mass number**. The mass number of an atom is the total number of **protons and neutrons** in the nucleus.

This is the **atomic number** for lithium. This number tells us how many **protons are in the nucleus**. No other element has the same atomic number as lithium.

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

We know that the 'Li' is the chemical symbol for lithium, but what do those numbers mean?

To calculate the number of neutrons in an element use this equation

Mass number – atomic number = neutron number

7 – 3 = 4 neutrons

Ions and isotopes

Isotopes – An atom of the **same element** with the **same number of protons** and **different numbers of neutrons**

Ions – An element with the **same number of protons** but different **numbers of electrons**. This means they will have a **positive or negative charge**

How to calculate proton, neutron and electron number of ions

Element	Protons	Neutrons	electrons
Ca	20	20	20
Ca^{+2}	20	20	18
Na	11	12	11
Na^{+1}	11	12	10
F	9	10	9
F^{-1}	9	10	8

These are **all ions** as they have **positive or negative charges**. When electrons are **lost** then the charge becomes **positive** and when **negative** electrons are **gained**

- +2 charge means two electrons are lost
- -1 charge means 1 electron is gained
- All elements reacts to gain a complete outer shell!

Ions and isotopes

How to calculate isotope abundance

A sample contains three different isotopes of oxygen.

Isotope	Isotope percentage
O^{15}	30
O^{16}	45
O^{17}	25

Calculate the relative atomic mass of oxygen in the sample.

Step 1 – Multiply mass by percentage and add them together

15 x 30 = 450 16 x 45 = 720 17 x 25 = 425

450 + 720 + 425= 1595

Step 2 – Divide the total mass by 100

1595 ÷ 100 = 15.95

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Calculating empirical formula

Video links

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Empirical formula walkthrough

Calculating empirical formula

Using the mass of each element in a compound

In an experiment 4.86g of magnesium was combined with 31.96g of bromine to form a compound. (A_r:Mg-24.3; Br-79.9)

Hint – Set it up in a table

Step 1 – Calculate the moles of each element

	Mg	**Br**
Mass	4.86	31.96
Mr	24.3	79.9
Moles	4.86÷24.3=0.2	31.96÷79.9=0.4

Step 2 - Find the smallest whole number ratio by **dividing the moles of each element** by the **smallest number of moles**

	Mg	**Br**
Moles	0.2	0.4
Molar ratio	0.2÷0.2=1	0.4÷0.2=2

Step 3 – Write out the formula

$MgBr_2$

Calculating empirical formula

The % by mass for the elements in a compounds are as follows: 26.6% Potassium, 35.4% Chromium and the rest is oxygen. Calculate the empirical formula for a compound.

Hint – Set it up in a table* & *assume the percentage is the mass

Step 1 – Calculate the percentage of oxygen

100 – (35.4 + 26.6) = 38% We can assume that the **percentages** are **masses**

Step 2 – Calculate the moles of each element

	K	**Cr**	**O**
Mass	26.6	35.4	38
Mr	39.1	52	16
Moles	26.6÷39.1=0.680	35.4÷52=0.681	38÷16=2.38

Step 3 - Find the smallest whole number ratio by **dividing the moles of each element** by the **smallest number of moles**

	K	**Cr**	**O**
Moles	26.6÷39.1=0.680	35.4÷52=0.681	38÷16=2.38
Molar ratio	0.680÷0.680= 1	0.681÷0.680= 1	2.38÷0.680=3.5

Step 4 – Write out the formula

$K_2Cr_2O_7$

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Using Avogadro’s number

Video links

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Moles and Avogadro's number

Using Avogadro's constant

Avogadro's constant – **6.02 x 10^{23}** This allows the number of atoms within a mole to be calculated. It also allows the number of molecules within mole to be calculated.

The basic rules

If calculating the number of **atoms** for an element

1. *Moles x Avogadro's constant*

e.g. How many atoms are there in 2 moles of Na
$2 \times 6.02X10^{23} = 1.204 \times 10^{24}$

If calculating the number of **molecules** for a **molecule**

1. *Moles x Avogadro's constant*

e.g. How many molecules are there in 3 moles of N_2
$3 \times 6.02X10^{23} = 1.806 \times 10^{24}$

Using Avogadro's constant

How to calculate the number of atoms/molecules in a sample:

Number of moles X ***Avogadro's constant = molecule/atom number***

Example 1 – A student has 5 moles of copper. How many atoms of copper does the student have in their sample?

$$5 \times (6.02 \times 10^{23}) = 3.01 \times 10^{24}$$

Example 2 – How many molecules of oxygen are there in 3.6 moles of oxygen (O_2)

$$3.6 \times (6.02 \times 10^{23}) = 2.1672 \times 10^{24}$$

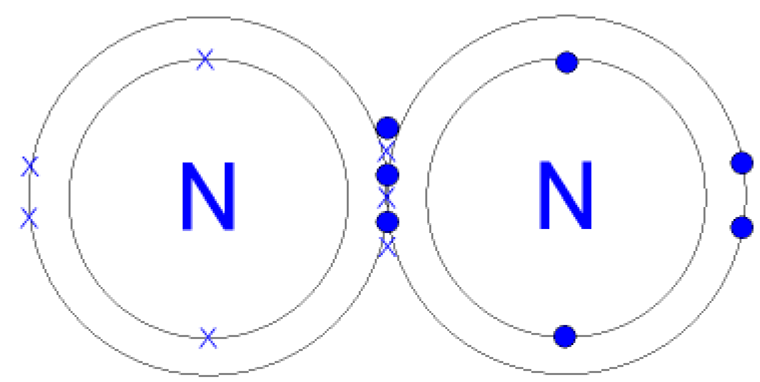
Using Avogadro's constant

The constant can also be used to also calculate the number of atoms in moles of a molecules

<u>The more complex rule</u>

If calculating the number of **atoms** for a **molecule**

N_2



Every nitrogen molecule contains 2 nitrogen ATOMS

1. *Moles x Avogadro's constant = Number of molecules*
2. *Number of molecules x number of atoms in molecule = number of atoms*

e.g. How many atoms are there in 3 moles of N_2

There are 3 moles in total $\quad 3 \times 6.02 \times 10^{23} = 1.806 \times 10^{24}$

There are 2 atoms per molecule $\quad 2 \times 1.806 \times 10^{24} = 3.612 \times 10^{24}$

Using Avogadro's constant

How to calculate the number of atoms/molecules in a sample:

Number of moles X ***Avogadro's constant = molecule/atom number***

One more – A student has 2.6 moles of oxygen molecules (O_2). How many atoms of oxygen does the student have?

Remember oxygen (O_2) has TWO ATOMS per molecule

Answer: $2.6 \times (6.02 \times 10^{23}) = 1.5652 \times 10^{24}$ Number of molecules

$1.5652 \times 10^{24} \times 2 = 3.1304 \times 10^{24}$

Number of atoms

A little more complex - A student has 1.5 moles of ethene molecules (C_2H_4). How many atoms does the student have?

Remember that ethene has 6 ATOMS per molecule (2 carbon and 4 hydrogen)

Answer: $1.5 \times (6.02 \times 10^{23}) = 9.03 \times 10^{23}$ Number of molecules

$9.03 \times 10^{23} \times 6 = 5.418 \times 10^{24}$

Number of atoms

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Balancing equations

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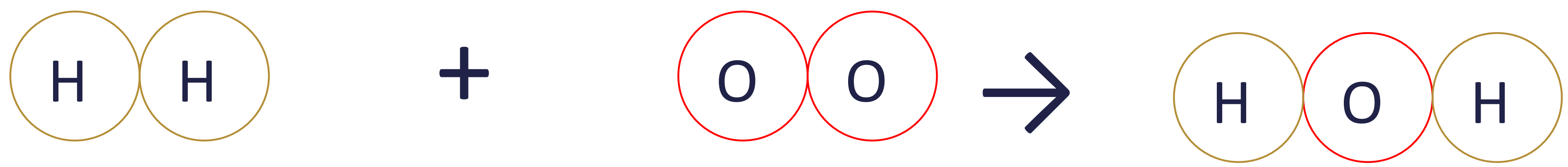
Balancing chemical equation walkthrough

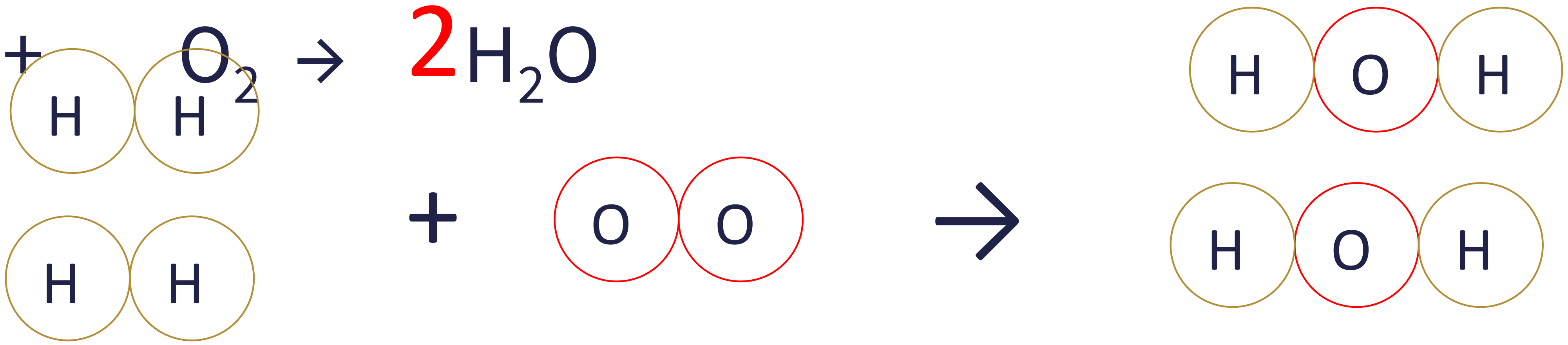
Rules for balancing equations

1. **You can only change the coefficients, the large font numbers in front of formula** (you can have as many particles as you need)
2. **You cannot change the small font numbers in a formula** (This would change the chemicals).

- Example $H_2 + O_2 \rightarrow H_2O$



- The number of oxygen atoms on each side of the arrow are not the same. We need the same number of atoms each side of the arrow to be balanced.
- There are 2 oxygen atoms on one side, and one on the other… We need two H_2O molecules.
- Now we have two hydrogen atoms on the left, and four on the right…… So we need to put a 2 in front of H_2.
- $2H_2 + O_2 \rightarrow 2H_2O$



How to balance equations

1. Write out the unbalanced symbol equation (leave spaces for coefficients).
2. One element at a time, check and balance changing coefficients to balance numbers.
3. Elements balancing order. Metals, non-metals, finally C, H, then O
4. If you leave O_2 to last… and you can finish the balancing by writing a half value, like $1.5O_2$, you can, but then you have to double the whole equation to complete it.
5. When balancing an element, if the number is odd one side, and even the other, just double the odd number first. Then continue to balance that element.
6. **Do a final double check to balance all elements.**

Balancing equations

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

Step 1 – Find out how many of each element there is on the reactants and products side of the equation

Element	Reactant atom number	Product atom number
C	1	1
H	4	2
O	2	3

Balancing equations

$$CH_4 + O_2 \rightarrow CO_2 + 2H_2O$$

Step 2 – Balance one of the elements by adding big numbers in front of different compounds. *This might worse before it gets better*

Element	Reactant atom number	Product atom number
C	1	1
H	4	~~2~~ 4
O	2	~~3~~ 4

This has changed as we now have 2X1 oxygen from water (H_2O) and we already had 2 oxygen from carbon dioxide (CO_2)

This has changed as we now have 2 x the number of hydrogen (2X2 = 4)

Balancing equations

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

Step 3 – Balance the other side of the question to make the atom number on both reactant and product side the same.

Element	Reactant atom number	Product atom number
C	1	1
H	4	~~2~~ 4
O	~~2~~ 4	~~3~~ 4

This has changed as we now have
2 x the number of oxygen (2X2=4)

Balancing equations

$$C_3H_8 \quad + \quad O_2 \quad \rightarrow CO_2 \quad + \quad H_2O$$

Step 1 – Find out how many of each element there is on the reactants and products side of the equation

Element	Reactant atom number	Product atom number
C	3	1
H	8	2
O	2	3

Balancing equations

$$C_3H_8 \ + \ O_2 \ \rightarrow \ 3CO_2 \ + \ H_2O$$

Step 1 – Find out how many of each element there is on the reactants and products side of the equation

Element	Reactant atom number	Product atom number
C	3	~~1~~ 3
H	8	2
O	2	~~3~~ 7

Balancing equations

$$C_3H_8 \quad + \quad O_2 \quad \rightarrow 3CO_2 \quad + \quad 4H_2O$$

Step 1 – Find out how many of each element there is on the reactants and products side of the equation

Element	Reactant atom number	Product atom number
C	3	~~1~~ 3
H	8	~~2~~ 8
O	2	~~3~~ ~~7~~ 10

Balancing equations

$$C_3H_8 \ + \ 5O_2 \rightarrow 3CO_2 \ + \ 4H_2O$$

Step 1 – Find out how many of each element there is on the reactants and products side of the equation

Element	Reactant atom number	Product atom number
C	3	~~1~~ 3
H	8	~~2~~ 8
O	~~2~~ 10	~~3~~ ~~7~~ 10

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Relative atomic mass

Video links

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Calculating moles walkthrough

Relative Atomic Mass

Relative Atomic Mass (A_r) is just the mass number from the periodic table (the large number protons + neutrons)

Element	Symbol	R.A.M (A_r)
Cobalt	Co	58.9
Rubidium	Rb	85.5
Silicon	Si	28.1
Neon	Ne	20.2
Boron	B	10.8

Relative Formula Mass (Mr)

- Is the mass of a substance made up of more than one atom
- To calculate this we simply add together the **atomic masses** of all the atoms shown in the formula, being careful to make sure that we look at the formula.

Substance	Formula	Formula Mass
Ammonia	NH_3	14 + 3 = 17
Sodium oxide	Na_2O	23 + 23 +16 = 62
Magnesium hydroxide	$Mg(OH)_2$	24.3 + [(16 + 1) x 2] = 58.3
Calcium carbonate	$CaCO_3$	40.1 + 12 + 3x16 = 100.1

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Calculating moles

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Calculating moles walkthrough

Vital equations

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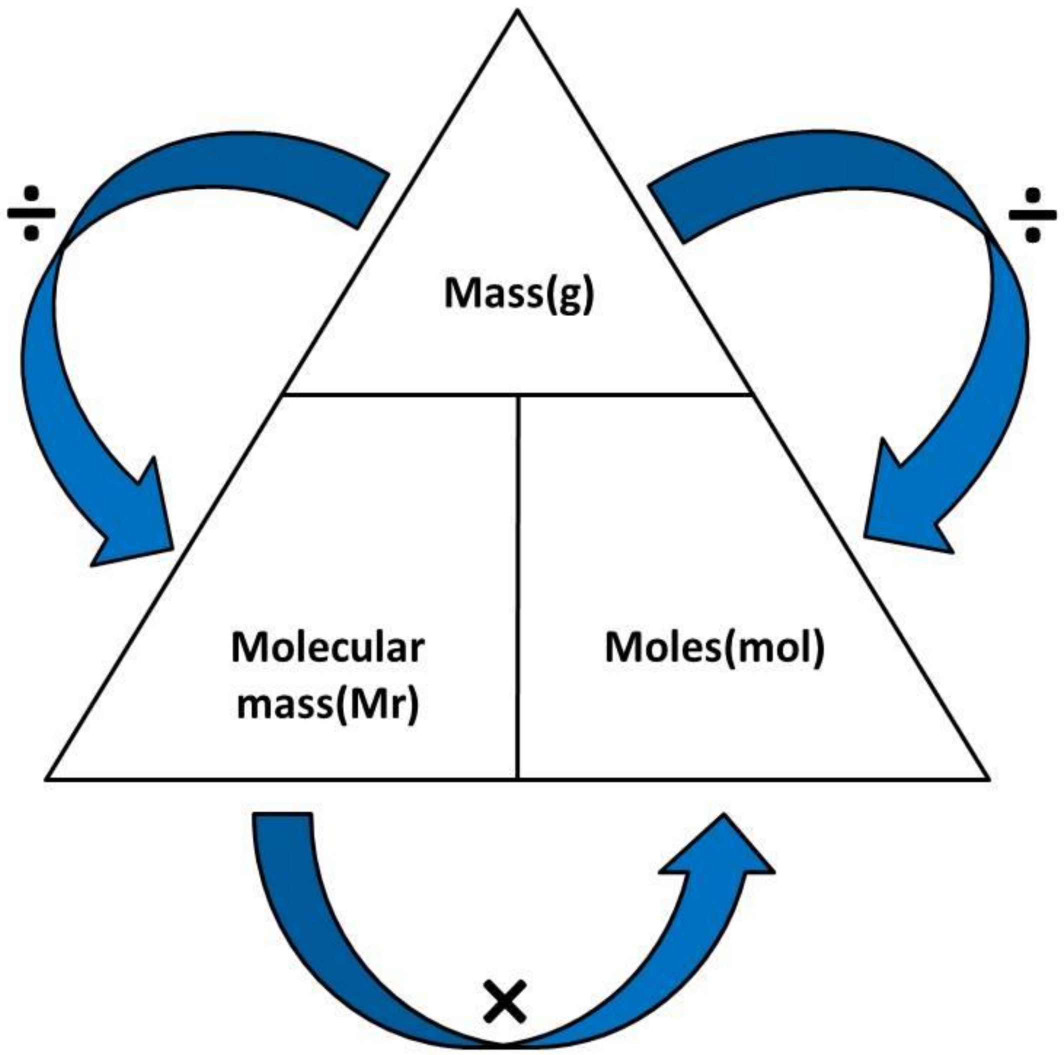
- Mass (g) ÷ Mr (g/mol) = Moles
- Volume (dm^3) x Concentration (g/dm^3) = Mass
- Volume (dm^3) x Concentration (mol/dm^3) = Moles
- Moles X 24 = Volume (dm^3)

Note *– Concentration can be written as mol/dm^3 or $moldm^{-3}$ as these are the same unit!*

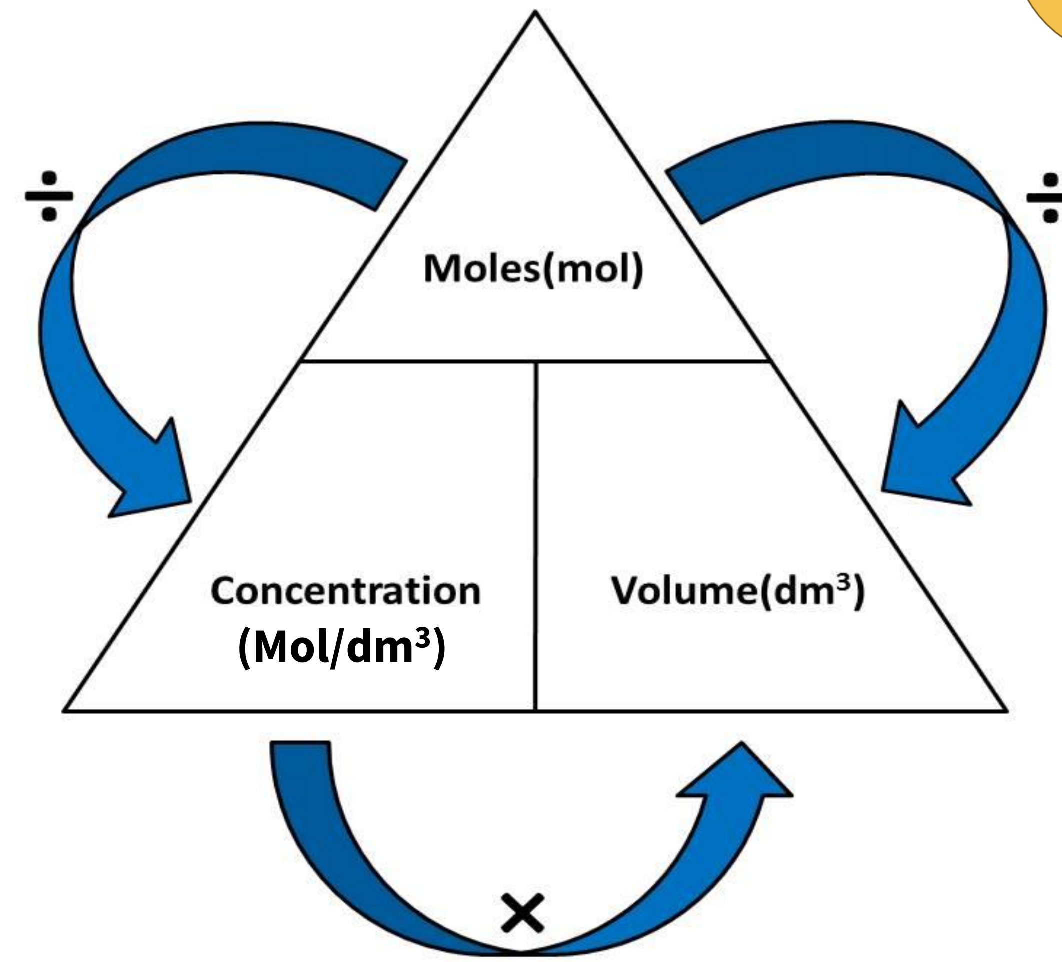
Units and conversations

- Converting cm^3 to dm^3 ÷ 1000
- g/dm^3 is the same as gdm^{-3}

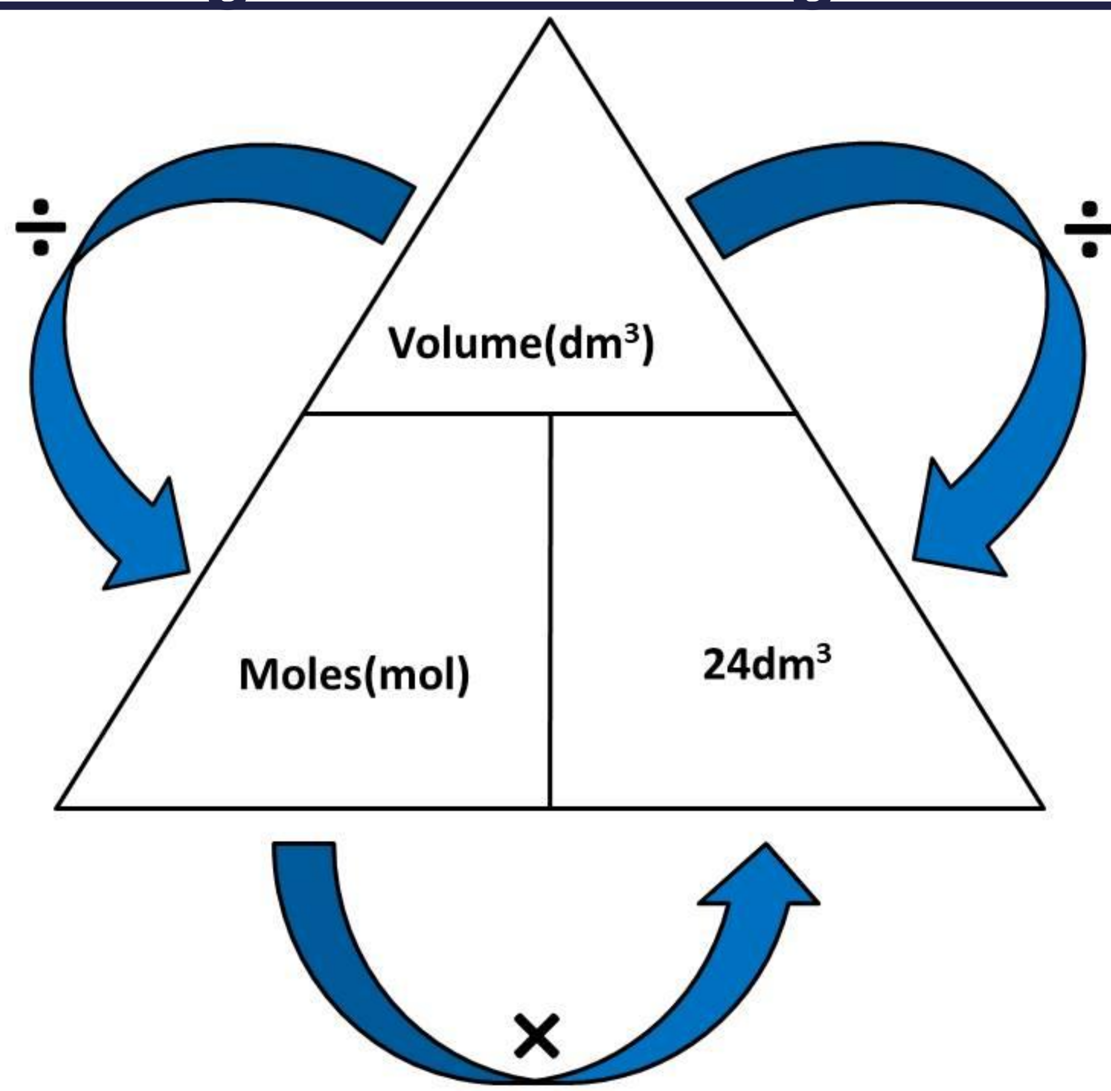
Calculating moles from mass

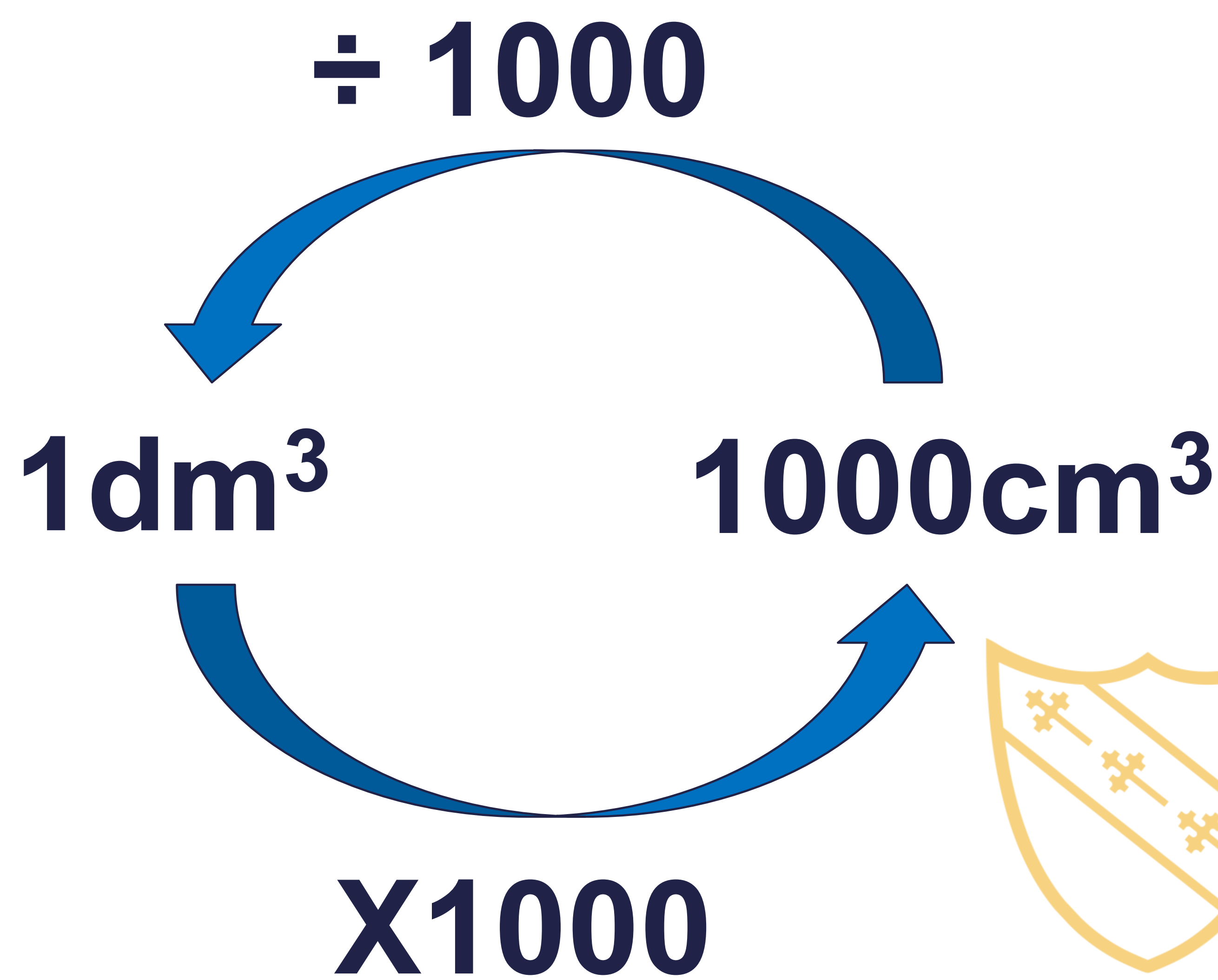


Calculating moles from volume and concentration



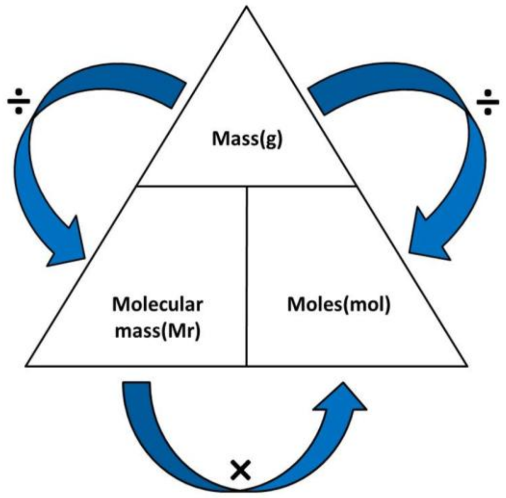
Calculating moles from a gas volume





Moles & Mass

1. How many moles are in 24g of carbon
   Equation - Mass ÷ Mr = Mole
   24 ÷ 12 = 2 mol
2. How many grams are in 5 moles of magnesium
   Equation - Mass ÷ Mr = Mole
   Rearrange – Mole x Mr = Mass
   5 x 24.3 = 121.5g
3. How many grams are in 7 moles of sodium hydroxide (NaOH)
   Equation - Mass ÷ Mr = Mole
   Rearrange – Mole x Mr = Mass
   7 X (23+16+1) = 280



Calculating Mass From Moles

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Calculate the mass of iron (Fe) that can be extracted from 8kg of Fe_2O_3 in the reaction with carbon (4 marks)

Step 1 – Draw the mass, Mr mol table

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

Mass		
Mr		
Mol		
Molar ratio		

Calculating Mass From Moles

11/06/2026

Calculate the mass of iron (Fe) that can be extracted from 8kg of Fe_2O_3 in the reaction with carbon (4 marks)

Step 2 – Identify the known and unknown compounds

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

	Fe	**Fe_2O_3**
Mass		
Mr		
Mol		
Molar ratio		

Calculating Mass From Moles

Calculate the mass of iron (Fe) that can be extracted from 8kg of Fe_2O_3 in the reaction with carbon (4 marks)

Step 3 – Calculate the mass of the known

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

	Fe	**Fe_2O_3**
Mass		8000
Mr		159.6
Mol		50
Molar ratio		

Calculating Mass From Moles

Calculate the mass of iron (Fe) that can be extracted from 8kg of Fe_2O_3 in the reaction with carbon (4 marks)

Step 4 – Write the ratio (big numbers in front of the compounds) to calculate the moles of the unknown

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

	Fe	**Fe_2O_3**
Mass		8000
Mr		159.6
Mol	100	50
Molar ratio	4	2

If the ratio doubles, then so does the moles!

Calculating Mass From Moles

Calculate the mass of iron (Fe) that can be extracted from 8kg of Fe_2O_3 in the reaction with carbon (4 marks)

Step 5 – Calculate the mass of the unknown

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

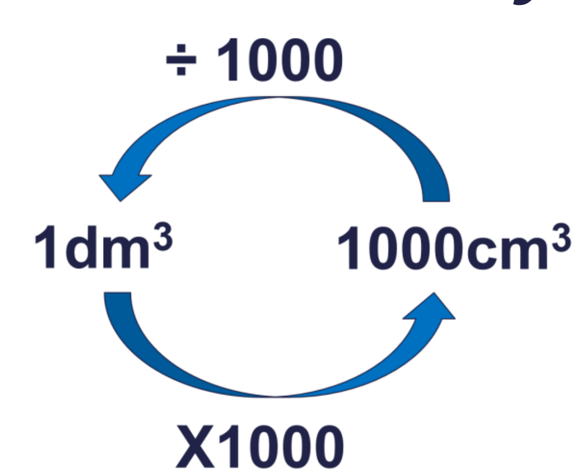
	Fe	**Fe_2O_3**
Mass	5580	8000
Mr	55.8	159.6
Mol	100	50
Molar ratio	4	2

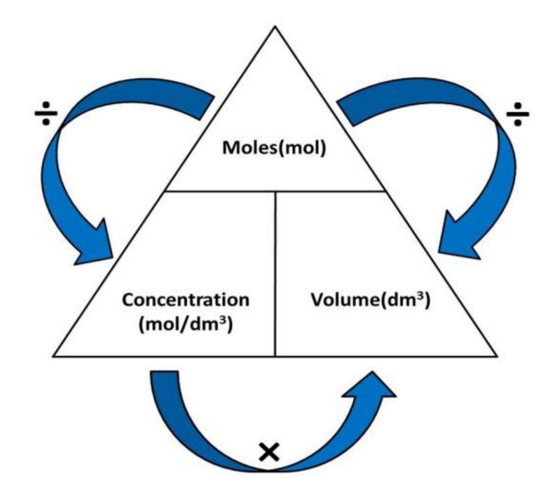
Moles and Concentration

1. How many moles are in $50cm^3$ of $0.5mol/dm^3$ HCl
   Equation – Volume x Concentration = Moles
   Volume = $50cm^3$ This must be in dm^3 so divide by 1000
   Concentration = 0.5 mol/dm^3
   0.05 x 0.5 = 0.025 mol

Concentration units are mol/dm^3 or $moldm^{-3}$. They can also be g/dm^3 or gdm^{-3} which can be converted into $mols/dm^3$ by ÷ Mr of compound e.g. $50g/dm^3$ of NaOH = 50 ÷ 40 = $1.25mol/dm^3$

2. What is the concentration of $100cm^3$ of 2 mols of MgO
   Equation – Volume x Concentration = Moles
   Rearrange – Moles ÷ Volume = Concentration
   Volume = $100cm^3$ This must be in dm^3 so divide by 1000
   Moles = 2 mol
   2 ÷ 0.1 = 20 mol/dm^3



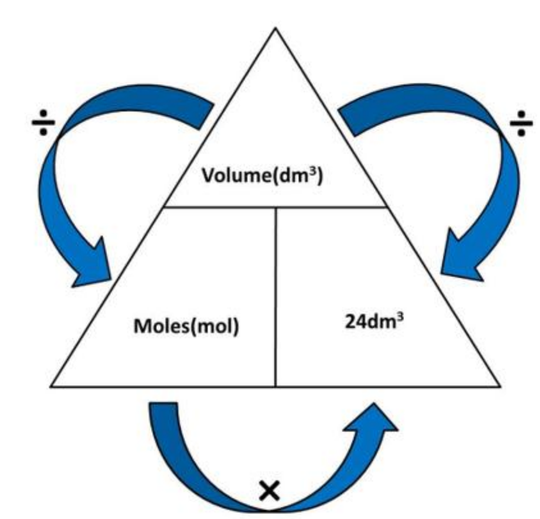


Moles & Volume of Gas

1. How many moles are there in $48dm^3$ of oxygen gas?
   Equation – Moles x 24 = Volume
   Rearrange – Volume ÷ 24 = Moles
   48 ÷ 24 = 2 Moles

2. How many moles are there in $100dm^3$ of Carbon dioxide gas to 2.d.p?
   Equation – Moles x 24 = Volume
   Rearrange – Volume ÷ 24 = Moles
   100 ÷ 24 = 4.17

3. What volume of hydrogen gas is in 5 moles? Put your answer in cm^3.
   Equation – Moles x 24 = Volume
   5 X 24 = $120dm^3$
   $120dm^3$ x 1000 = $120{,}000cm^3$



Moles & Volume of Gas – Complex Example

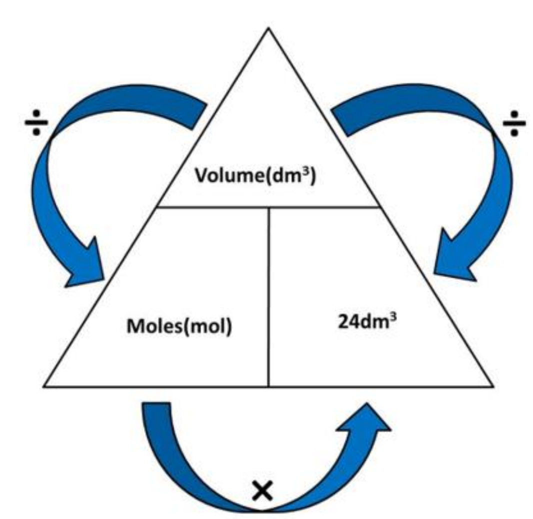
1. 4dm^3 of carbon reacts with oxygen to form carbon monoxide. Calculate the volume of oxygen needed for the reaction.

$$2C_{(g)} + O_{2(g)} \rightarrow 2CO_{(g)}$$

When dealing with ONLY GASES you **don't need** to convert volume into moles. You can just **directly compare the volumes of gas**.

Step 1 – Work out the ration 2:1

	Carbon	**Oxygen**
Volume (dm^3)	4	
Ratio	2	1



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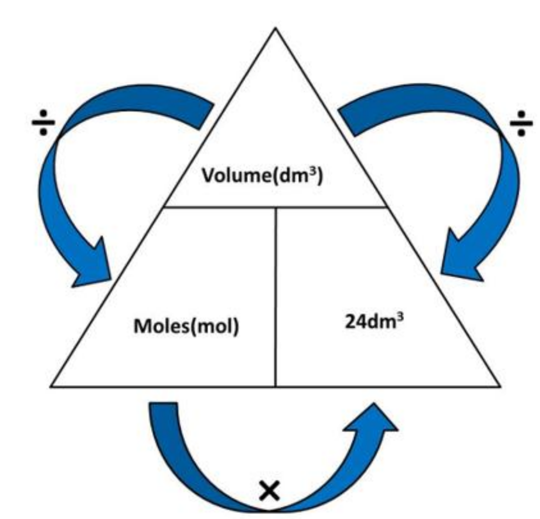
When dealing with ONLY GASES you **don't need** to convert volume into moles. You can just **directly compare the volumes of gas**.

Step 2 – Divide the volume by the ratio

	Carbon	**Oxygen**
Volume (dm^3)	4	4 ÷ 2 = **2**
Ratio	2	1



Moles & Volume of Gas – Complex Example

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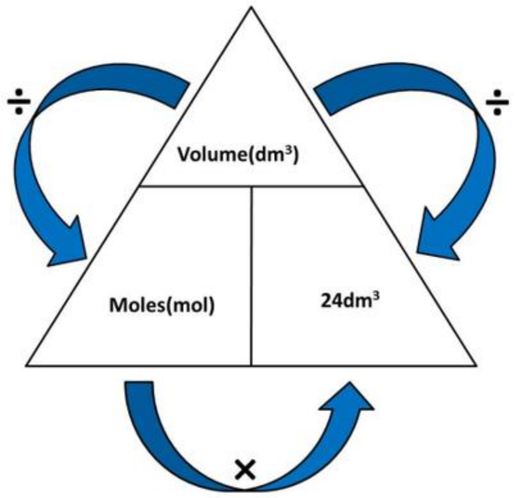
2. 5.0 dm^3 of oxygen reacts with 3.0dm^3 of hydrogen. Calculate the volume of water produced

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

Step 1 – Identify which element will run out first. Divide the volume by the ratio to get the comparison volumes

	Hydrogen	Oxygen	Water
Volume (dm^3)	3.0	5.0	
Ratio	2	1	
Comparison volume (dm^3)	3 ÷ 2 = 1.5	5 ÷ 1 = 5	

The comparison volume that is smaller will **run out first.** This is called the **limiting reactant.**



Moles & Volume of Gas – Complex Example

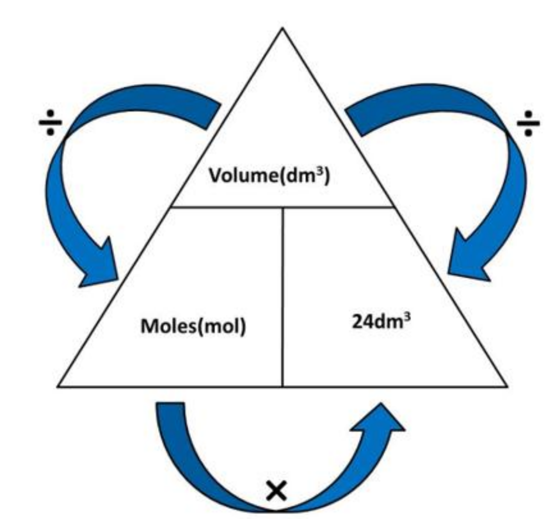
2. 5.0 dm³ of oxygen reacts with 3.0dm³ of hydrogen. Calculate the volume of water produced

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

Step 2 – Use the volume of the limiting reactant to calculate the volume of the product.

	Hydrogen	Oxygen	Water
Volume (dm³)	3.0	5.0	3.0
Ratio	2	1	2
Comparison volume (dm³)	3 ÷ 2 = 1.5	5 ÷ 1 = 5	--------------------------

As the **ratio is the same** for hydrogen (the limiting reactant) and water (product) then the **volumes are the same**.



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Percentage yield

Video links

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Percentage yield

Percentage Yield

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Even though no atoms are gained or lost in a chemical reaction, it is not always possible to obtain the calculated amount of a product because:

1. The reaction may not go to completion because it is reversible
2. Some of the product may be lost when it is separated from the reaction mixture
3. Some of the reactants may react in ways different to the expected reaction

To try and prevent the loss or reactants during a reaction we use a process called **REFLUX**

Percentage yield calculation

Percentage YIELD - The amount of product you ***actually*** make as a percentage (%) of the amount you should **theoretically** make

Actual mass ÷ Theoretical mass X 100 = Percentage yield

Worked example 1

In a reaction Mr Maidens obtained 4.6g of sodium but 5.8g of sodium was expected to be obtained. Calculate the percentage yield of the reaction to 2.d.p

- 4.6 is the actual mass
- 5.8 is the theoretical mass

Answer - (4.6 ÷ 5.8) X 100 = 79.31 (2.d.p)

Percentage yield calculation

Worked example 2

In a reaction Miss Swan obtained 6.5g of potassium but 8.9g of potassium was expected to be obtained. Calculate the percentage yield of the reaction to 2.d.p

- 6.5 is the actual mass
- 8.9 is the theoretical mass

Answer - $(6.5 \div 8.9) \times 100 = 73.03$ (2.d.p)

Percentage yield calculation

More complex examples

A student completes the following reaction

$$Ca + 2HCl \rightarrow CaCl_2 + H_2$$

The student starts with 8g calcium (Ca). They obtained 20.6g of calcium chloride ($CaCl_2$). Calculate the percentage yield of the reaction.

Note – 20.6g is the actual mass of calcium chloride obtained

Step 1 – Calculate the theoretical mass of calcium chloride

	Ca	**$CaCl_2$**
Mass	8	22.2
Mr	40	111
Mole	0.2	0.2
Molar ratio	1	1

Percentage yield calculation

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<u>More complex example 1</u>

A student completes the following reaction

$$Ca + 2HCl \rightarrow CaCl_2 + H_2$$

The student starts with 8g calcium (Ca). They obtained 20.6g of calcium chloride ($CaCl_2$). Calculate the percentage yield of the reaction.

Step 1 – Theoretical mass of calcium chloride = 22.2g

Step 2 – Actual mass ÷ Theoretical mass X 100 = Percentage yield

(20.6 ÷ 22.2) x 100 = 92.8%

Percentage yield calculation

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More complex example 2

A student completes the following reaction

$$2Na + 2HCl \rightarrow 2NaCl + H_2$$

The student starts with 5g sodium (Na). They obtained 6.6g of sodium chloride (NaCl). Calculate the percentage yield of the reaction to 3.s.f.

	Na	NaCl
Mass	5	12.71…
Mr	23	58.5
Mole	0.2173…	0.2173…
Molar ratio	1	1

5 ÷ 12.71 x 100 = 39.3% (3.s.f)

Percentage yield calculation

Working the opposite direction 1

A student completes the following reaction

$$CH_4 + 2O_2 \rightarrow + CO_2 + 2H_2O$$

Calculate the mass of methane (CH_4) needed to obtain 55g of carbon dioxide (CO_2) assuming there is a 70% yield of carbon dioxide.

Step 1 – Calculate 100% yield of carbon dioxide

- 55g is the actual mass of CO_2

55 ÷ Theoretical mass x 100 = 70%

55 ÷ Theoretical mass = 0.7

55 ÷ 0.7 = Theoretical mass

Theoretical mass = 78.571…g

Percentage yield calculation

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Working the opposite direction 1

A student completes the following reaction

$$CH_4 + 2O_2 \rightarrow + CO_2 + 2H_2O$$

Calculate the mass of methane (CH_4) needed to obtain 55g of carbon dioxide (CO_2) assuming there is a 70% yield of carbon dioxide.

Step 1 – Theoretical mass = 78.571…g

Step 2 – Calculate the mass of methane (CH_4) using theoretical mass

	CO_2	CH_4
Mass	78.571…	28.57
Mr	44	16
Mole	1.785…	1.785…
Molar ratio	1	1

Percentage yield calculation

Working the opposite direction 2

A student completes the following reaction

$$CaCO_3 \rightarrow + CaO + CO_2$$

Calculate the mass of calcium carbonate ($CaCO_3$) needed to obtain 40g of calcium oxide (CaO) assuming there is a 50% yield of calcium oxide to 1.d.p.

Step 1 – Calculate 100% yield of calcium oxide

$40 \div 0.5 = 80g$

Step 2 – Calculate the mass of calcium carbonate ($CaCO_3$) using theoretical mass

Percentage yield calculation

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Working the opposite direction 2

A student completes the following reaction

$$CaCO_3 \rightarrow + CaO + CO_2$$

Calculate the mass of calcium carbonate ($CaCO_3$) needed to obtain 40g of calcium oxide (CaO) assuming there is a 50% yield of calcium oxide to 1.d.p.

Step 1 – 40 ÷ 0.5 = 80g

Step 2 – Calculate the mass of calcium carbonate ($CaCO_3$) using theoretical mass

	CaO	**$CaCO_3$**
Mass	80	142.7
Mr	56.1	100.1
Mole	1.426…	1.426…
Molar ratio	1	1

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Electron configuration

Video links

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Electronic configuration 1

Electronic configuration 2

Electron configuration

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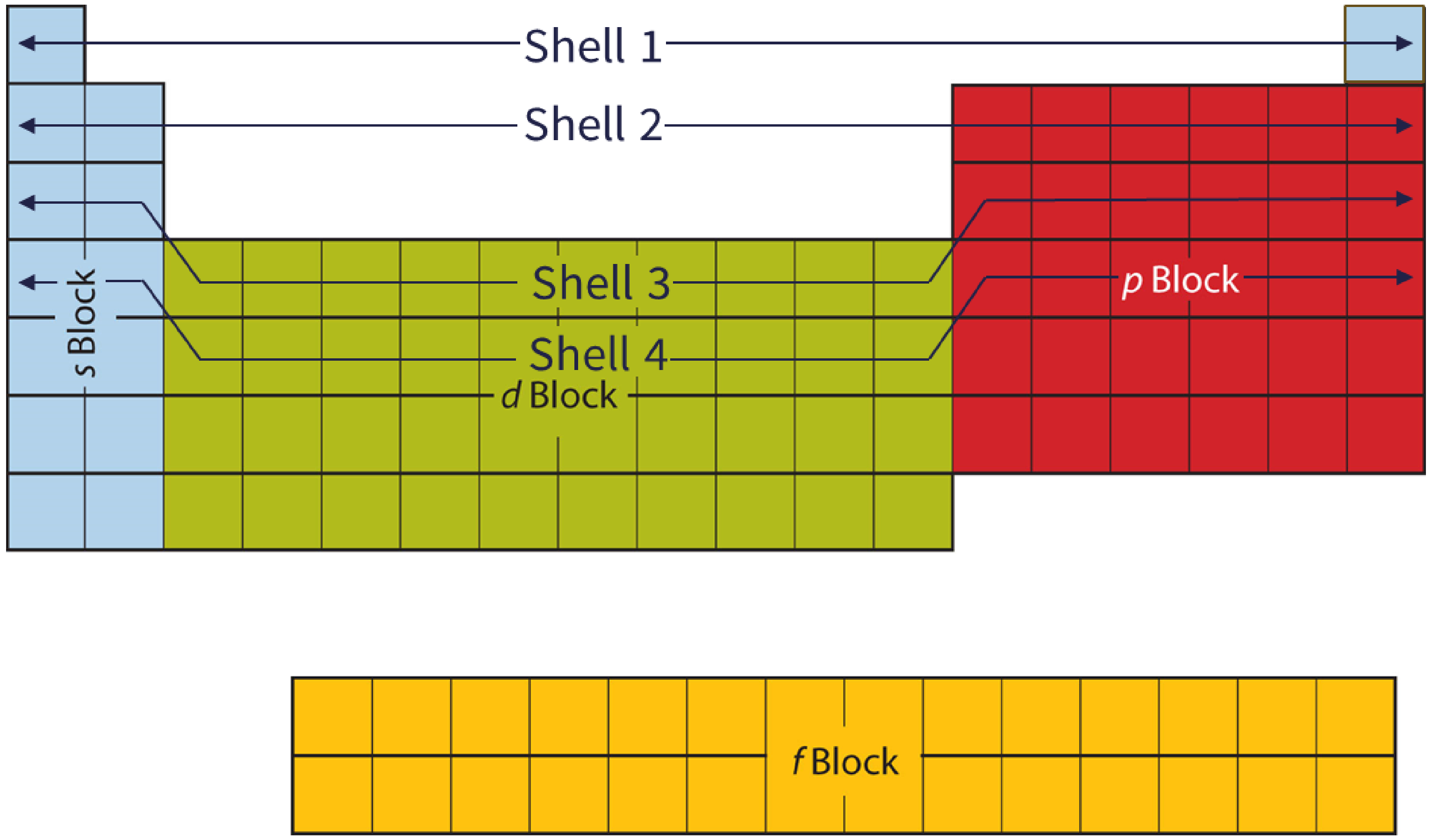

In GCSE we had a simplified electrons arrangement!

Electrons are arranged into energy levels called shells. Each row on the periodic table is a shell.

- The first shell of electrons can contain 2 electrons in total
- The second shell can contain 8 electrons in total
- The third shell can contain 18 electrons in total

This can happen because each energy level is divided into sub-shells. These are called:

The **s** sub-shell, The **p** sub-shell, The **d** sub-shell, The **f** sub-shell

Sub-shells and orbitals

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Each sub-shell can be further split into orbitals.

An orbital is a region in a sub-shell that can contain a maximum of 2 electrons.

The s Sub-shell

This contains 1 orbital and so each **s** sub-shell can have a MAXIMUM of 2 electrons within it.

The p sub-shell

This contains 3 orbitals and so each **p** sub-shell can have a MAXIMUM of 6 electrons within it.

The d sub-shell

This contains 5 orbitals and so each **d** sub-shell can have a MAXIMUM of 10 electrons within it.

The f sub-shell

This contains 7 orbitals and so each **f** sub-shell can have a MAXIMUM of 14 electrons within it.

Sub-shells and orbitals

Each energy level/shell can contain multiple sub-shells. So this means that the total number of electrons in that energy level is determined by the total number of sub-shells within that energy level/shell.

For example:

- An element in shell 1 can have a maximum of 2 electrons in its outer shell/energy level ($1s^2$)
- An element in shell 2 can have a maximum of 8 electrons in its outer shell/energy level ($2s^2$,$2p^6$)
- An element in shell 3 can have a maximum of 18 electrons in its outer shell/energy level ($3s^2$,$3p^6$,$3d^{10}$)

Writing electron configurations

We can write out electron configurations for elements to show how many electrons are within each shell and within each sub-shell

Example 1

Write the electron configuration for Boron

Boron has a total of 5 electrons. We fill the first energy level and then the second just like at GCSE

Number of electrons in sub-shell

$$B = 1s^2, 2s^2, 2p^1$$

Energy level

Sub-shell

Writing electron configurations

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We can write out electron configurations for elements to show how many electrons are within each shell and within each sub-shell

<u>Example 2</u>

Write the electron configuration for calcium

Calcium has a total of 20 electrons. We fill the first energy level and then the second just like at GCSE

Number of electrons in sub-shell

$$Ca = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$$

Energy level   Sub-shell

Note *– The 4s sub-shell fills BEFORE the 3d sub-shell as it required less energy and is easier to fill*!

Writing electron configurations

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We can write out electron configurations for elements to show how many electrons are within each shell and within each sub-shell

<u>Example 3</u>

Write the electron configuration for vanadium

Vanadium has a total of 23 electrons. We fill the first energy level and then the second just like at GCSE

Number of electrons in sub-shell

$$V = 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^3$$

Energy level   Sub-shell

Note *– The 4s sub-shell fills BEFORE the 3d sub-shell as it required less energy and is easier to fill*!