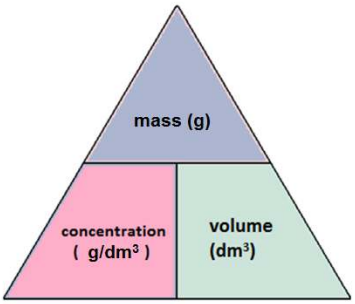
Plymstock School

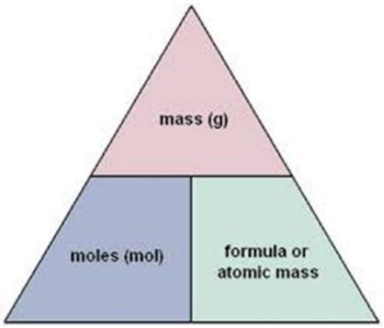
Chemistry

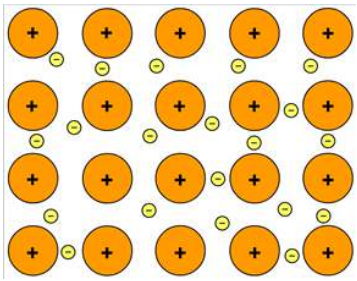
Bridging the Gap

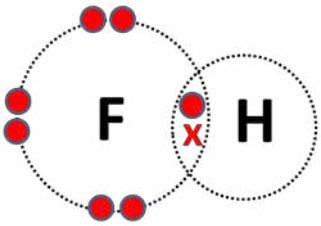
GCSE to A level Chemistry

Name: _____









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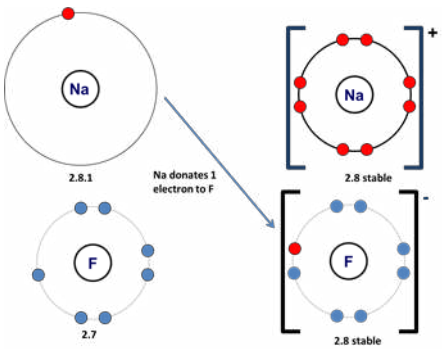


Welcome

Welcome to AS-Level Chemistry! This work is designed to help you revise your GCSE Chemistry over the summer to prepare you for starting AS-Level Chemistry in September. The aim is for you to practise your Chemistry and identify your strengths and weaknesses in the subject. This booklet focuses on two really important topics in Chemistry; Bonding and Structures and Quantitative Chemistry.

If you would like further work, or an insight into the wonderful world of AS-Level Chemistry and beyond, there are some further reading suggestions at the bottom of the contents pages. There is also a list of websites you will undoubtedly find useful throughout the course and may need to use to complete this task.

Good luck and happy Chemis-trying!

From the Chemistry Team

AQA

The Periodic Table of Elements

Key

relative atomic mass
atomic symbol
name
atomic (proton) number

1	2											3	4	5	6	7	0
							1 **H** hydrogen 1										4 **He** helium 2
7 **Li** lithium 3	9 **Be** beryllium 4											11 **B** boron 5	12 **C** carbon 6	14 **N** nitrogen 7	16 **O** oxygen 8	19 **F** fluorine 9	20 **Ne** neon 10
23 **Na** sodium 11	24 **Mg** magnesium 12											27 **Al** aluminium 13	28 **Si** silicon 14	31 **P** phosphorus 15	32 **S** sulfur 16	35.5 **Cl** chlorine 17	40 **Ar** argon 18
39 **K** potassium 19	40 **Ca** calcium 20	45 **Sc** scandium 21	48 **Ti** titanium 22	51 **V** vanadium 23	52 **Cr** chromium 24	55 **Mn** manganese 25	56 **Fe** iron 26	59 **Co** cobalt 27	59 **Ni** nickel 28	63.5 **Cu** copper 29	65 **Zn** zinc 30	70 **Ga** gallium 31	73 **Ge** germanium 32	75 **As** arsenic 33	79 **Se** selenium 34	80 **Br** bromine 35	84 **Kr** krypton 36
85 **Rb** rubidium 37	88 **Sr** strontium 38	89 **Y** yttrium 39	91 **Zr** zirconium 40	93 **Nb** niobium 41	96 **Mo** molybdenum 42	[98] **Tc** technetium 43	101 **Ru** ruthenium 44	103 **Rh** rhodium 45	106 **Pd** palladium 46	108 **Ag** silver 47	112 **Cd** cadmium 48	115 **In** indium 49	119 **Sn** tin 50	122 **Sb** antimony 51	128 **Te** tellurium 52	127 **I** iodine 53	131 **Xe** xenon 54
133 **Cs** caesium 55	137 **Ba** barium 56	139 **La*** lanthanum 57	178 **Hf** hafnium 72	181 **Ta** tantalum 73	184 **W** tungsten 74	186 **Re** rhenium 75	190 **Os** osmium 76	192 **Ir** iridium 77	195 **Pt** platinum 78	197 **Au** gold 79	201 **Hg** mercury 80	204 **Tl** thallium 81	207 **Pb** lead 82	209 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89	[261] **Rf** rutherfordium 104	[262] **Db** dubnium 105	[266] **Sg** seaborgium 106	[264] **Bh** bohrium 107	[277] **Hs** hassium 108	[268] **Mt** meitnerium 109	[271] **Ds** darmstadtium 110	[272] **Rg** roentgenium 111	[285] **Cn** copernicium 112	[286] **Nh** nihonium 113	[289] **Fl** flerovium 114	[289] **Mc** moscovium 115	[293] **Lv** livermorium 116	[294] **Ts** tennessine 117	[294] **Og** oganesson 118

* The Lanthanides (atomic numbers 58 – 71) and the Actinides (atomic numbers 90 – 103) have been omitted.

Relative atomic masses for **Cu** and **Cl** have not been rounded to the nearest whole number.

Contents

5.2 Bonding and structure

Topic	Lesson Content	Additional resources
1.Chemical bonds *Pages 7-9*	• Describe the three main types of bonding. • Explain how bonding and properties are linked. • Explain how positive and negative ions form.	**1.** Youtube video: https://www.youtube.com/watch?v=5aBbDxQEE3M 8 mins 35 secs
2.Ionic bonding *Pages 10-16*	• Represent an ionic bond with a diagram. • Draw dot and cross diagrams for ionic compounds.	**2.** Cognito science video– https://www.youtube.com/watch?v=6DtrrWA5nkE Freesciencelessons videos - https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/ionic-bonding-1/ https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/ionic-bonding-2/
3. Ionic compound properties *Pages 17-19*	• Identify ionic compounds from structures. • Work out the empirical formula of an ionic compound. • Describe and explain the properties of ionic compounds.	**3.** Cognito science video - https://www.youtube.com/watch?v=kShlfIsvWbQ Freesciencelessons videos - https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/properties-of-ionic-compounds/
4. Covalent bonding *Pages 20-24*	• Recognise substances made of small molecules from their formula. • Draw dot and cross diagrams for small molecules. • Deduce molecular formulae from models and diagrams.	**4.** Cognito science video - https://www.youtube.com/watch?v=5I_1jRGSR9E Freesciencelessons videos – https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/covalent-bonding-1/ https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/covalent-bonding-2/ https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/covalent-bonding-3/
5. Properties of small molecules *Pages 25-30*	• Identify small molecules from formulae. • Explain the strength of covalent bonds. • Relate the intermolecular forces to the bulk properties of a substance.	**5.** Cognito science video - https://www.youtube.com/watch?v=d2ogZgGmMDY Freesciencelessons videos – https://www.freesciencelessons.co.uk/gcse-chemistry-paper-1/structure-and-bonding/properties-of-small-covalent-molecules/

5.3 Quantitative Chemistry

6. Moles (HT only) *Pages 31-41*	• Describe the measurement of amounts of substances in moles. • Calculate the number of moles in a given mass. • Calculate the mass of a given number of moles.	**6. Moles (HT only)** Cognito science - https://tinyurl.com/yxkfuw8f Kay science – https://tinyurl.com/yyonkmvv and https://tinyurl.com/y64uy735 and https://tinyurl.com/y63zra3l
7.Amounts of substances in equations (HT only) *Page 42-46*	• Calculate the masses of substances in a balanced symbol equation. • Define Avogadro's constant • Calculate the masses of reactants and products from balanced symbol equations. • Calculate the mass of a given reactant or product.	**7. Amount of substances in equations (HT only)** Freesciencelesson - https://www.youtube.com/watch?v=5zOpoeN0dV0&t=9s Exam QA – https://www.youtube.com/watch?v=Lbzi52gMAIc
8. Using masses to balance equations (HT only) *Page 47-52*	• Convert masses in grams to amounts in moles. • Balance an equation given the masses of reactants and products. • Change the subject of a mathematical equation.	**8. Using moles to balance equations (HT only)** Freesciencelesson - https://www.youtube.com/watch?v=4wTSLBBBMo0&t=10s The ScienceBreak - https://www.youtube.com/watch?v=sjAH78janT0
9.Concentration of solutions *Page 53-59*	• Relate mass, volume and concentration. • Calculate the mass of solute in solution. • Relate concentration in mol/dm3 to mass and volume.	**9. Concentration of solutions** Cognito science - https://tinyurl.com/y3486bm2 Kay science – https://tinyurl.com/yx9fldyt and https://tinyurl.com/y6h6edez
10. Limiting reactants (HT only) *Pages 60-64*	• Define the term limiting reactant. • Be able to explain the effect of a limiting quantity of a reactant on the amount of products	**10. Limiting reactants (HT only)** Cognito science - https://tinyurl.com/y2dak83x Tyler Dewitt - https://www.youtube.com/watch?v=nZOVR8EMwRU&t=669s –

Useful Websites

GCSE revision: https://www.bbc.co.uk/bitesize/examspecs/z8r997h

A level: Chemguide www.chemguide.co.uk

A level: MaChemGuy https://www.youtube.com/user/MaChemGuy

Further Reading

These are some textbooks which you may find interesting and useful before and during your AS-Level Chemistry course.

*Essential Maths Skills for AS/A-Level Chemistry By Nora Henry
Published by Philip Allan for Hodder Education ISBN 978 1 4718 6349 3

*A-Level Year 1, Chemistry, OCR A Complete Revision and Practice Published by CGP ISBN 9781782943402

Lesson 1 – Bonding introduction

Learning Objectives

- Identify the bonding type from a chemical name and formulae
- Describe the three main types of bonding.
- Explain how electrons are used in the three types of bonding.

Explain and Understand Information

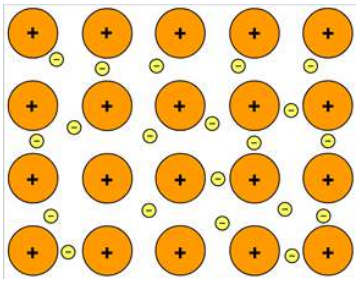
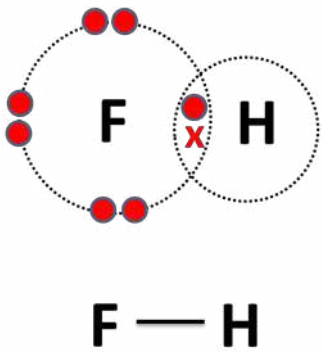
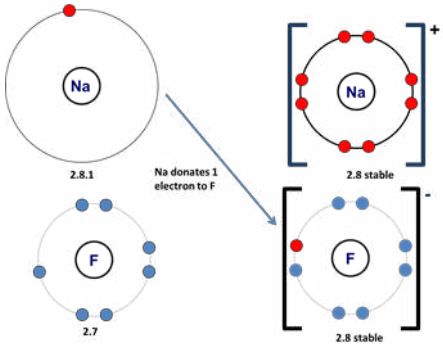
Atoms bond so that they can have a **full outer shell** of electrons, in order to be stable.

During chemical bonding only the sub-atomic particles **electrons** are involved.

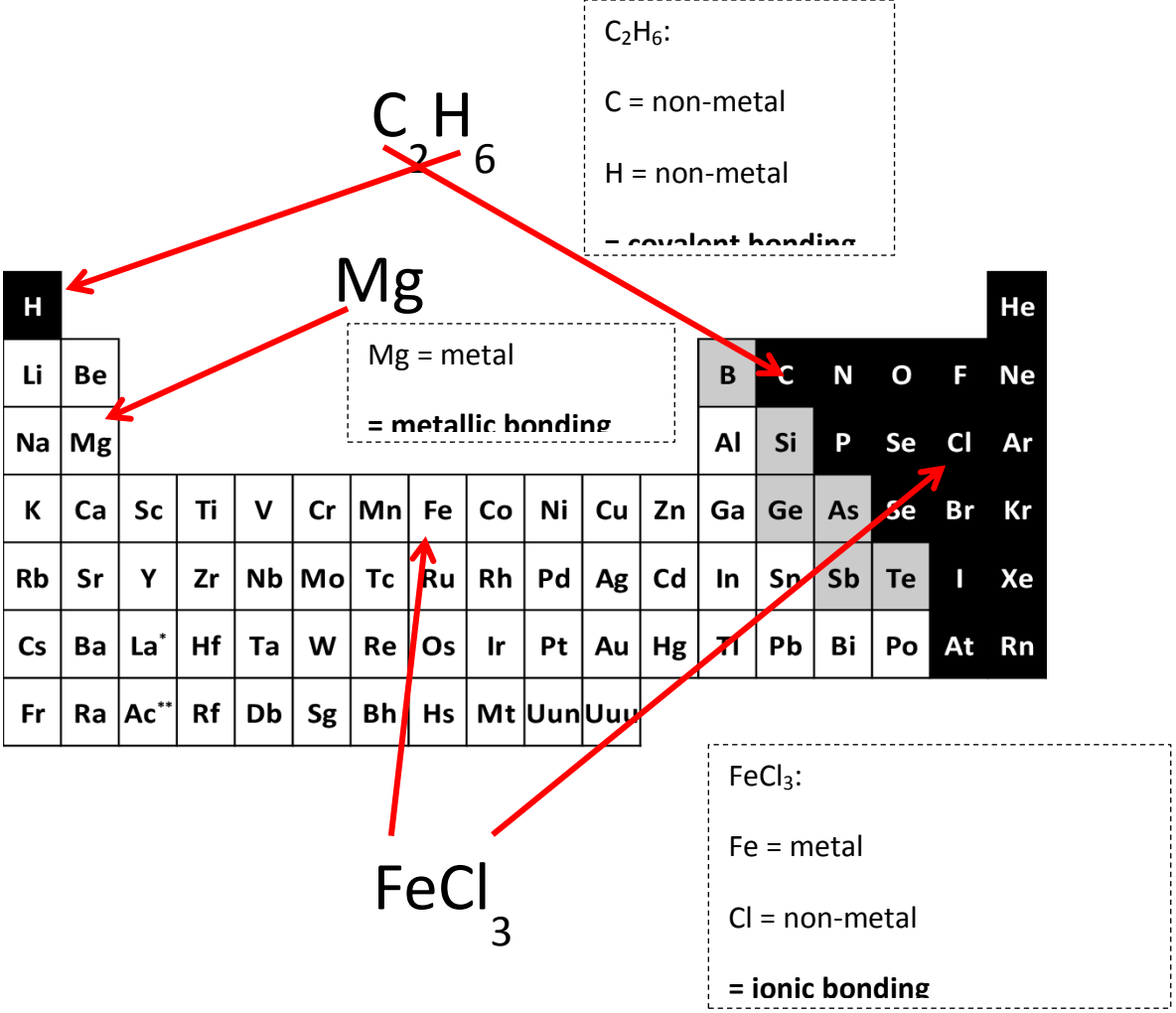
Group 0 (the Noble gases) do not form chemical bonds because they already have a full outer shell of electrons and so are stable.

There are **3 types of chemical bonds** that you are going to learn about – **metallic, covalent and ionic**.

Firstly, you need to be able to look at chemical formula and decide the type of bonding it has. The later lessons will look at each bonding in much greater detail. Below is a table summarising each bonding type.

Metallic	Covalent	Ionic
Element: Metals only. **Description**: The attraction between positive metal ions and delocalised electrons	**Elements:** Non-metals only. **Description:** Two non-metal atoms share a pair of electrons.	**Elements:** Metals and non-metals. **Description:** A metal atom transfers an electron to a non-metal atom to forms ions.
		

In order to determine which bonding type occurs in a chemical substance you need to use a periodic table to see if the elements are metals or non-metals. In the following periodic table, the white squares are the metals.



Use the information from above to help you complete the following tasks:

https://www.kayscience.com/chemistry.html C8.1

Explain and Understand Questions

1. Why do atoms form chemical bonds?

   _____

   _____

2. Which particle is involved with chemical bonding?

   _____

3. Which bonding type involves the sharing of a pair of electrons?

   _____

4. Determine the bonding in nickel (Ni). Use a periodic table to decide if the elements involved are metal or non-metal.

_____

5. Name the bonding type which occurs between a metal and a non-metal.

_____

6. Determine the bonding in carbon dioxide (CO_2). Use a periodic table to decide if each element is a metal or non-metal.

_____

_____

7. Determine the bonding in magnesium chloride ($MgCl_2$). Use a periodic table to decide if each element is a metal or non-metal.

_____

8. Describe metallic bonding.

_____

_____

9. Describe ionic bonding.

_____

_____

10. Describe covalent bonding.

_____

_____

Lesson 2 – Ionic bonding

Learning Objectives

- Work out the charge on the ions of metals and non-metals from the group number of the element (1, 2, 6 and 7).
- Represent an ionic bond with a diagram.
- Draw dot and cross diagrams for ionic compounds.

Explain and Understand Information

Forming ions

The number of **protons** that an atom has **cannot change**.

The number of **electrons** an atom has **can change** through a **chemical reaction**.

If an atom **gains or loses electrons** then the number of protons and electrons will no longer be equal.

Example: A lithium **atom** has three protons and three electrons. It has **no overall charge** because each proton is +1 and each electron is -1 so they **cancel each other out**.

	Protons	Electrons
	+1	-1
	+1	-1
	+1	-1
Total charge on each side:	+3	-3
	Overall: +3 + (-3) = 0	

However, if it **loses** one electron, then the charges become imbalanced:

	Protons	Electrons
	+1	-1
	+1	-1
	+1	
Total charge on each side:	+3	-2
	Overall: +3 + (-2) = +1	

So if a lithium atom **loses** an electron, it becomes a **positive ion**.

The opposite would happen if it **gained** an electron:

	Protons	Electrons
	+1	-1
	+1	-1
	+1	-1
		-1
Total charge on each side:	+3	-4
	Overall: +3 + (-4) = -1	

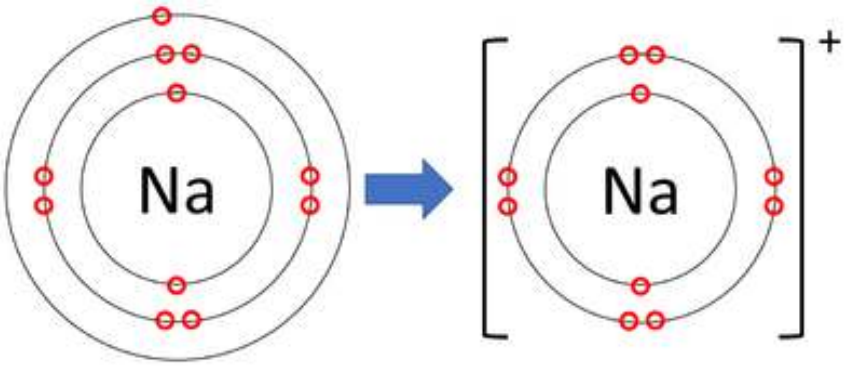
Summary of ions:

*If an atom **gains** electrons, it becomes a **negative** ion.*
*If an atom **loses** electrons, it becomes a **positive** ion.*

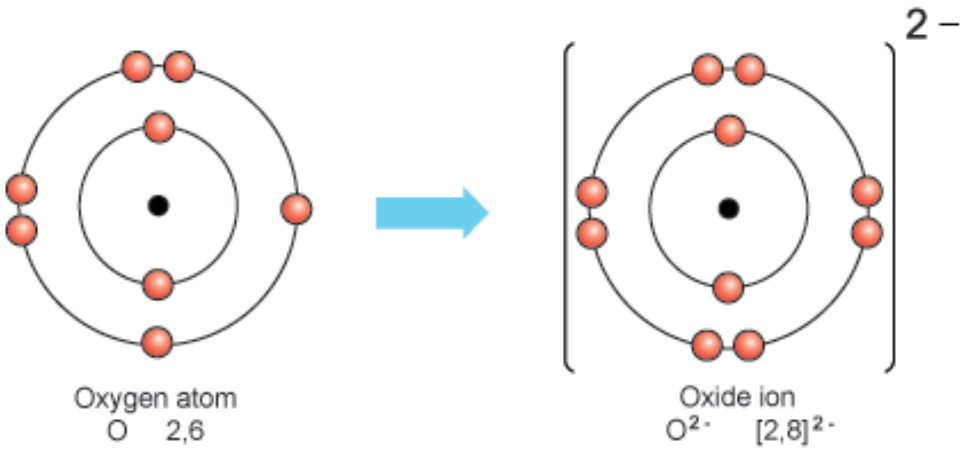
Transferring electrons to achieve a full outer shell

Different atoms have different numbers of electrons on their outer shell. In order to have a **full outer shell**, atoms can either **lose electrons or gain electrons**.

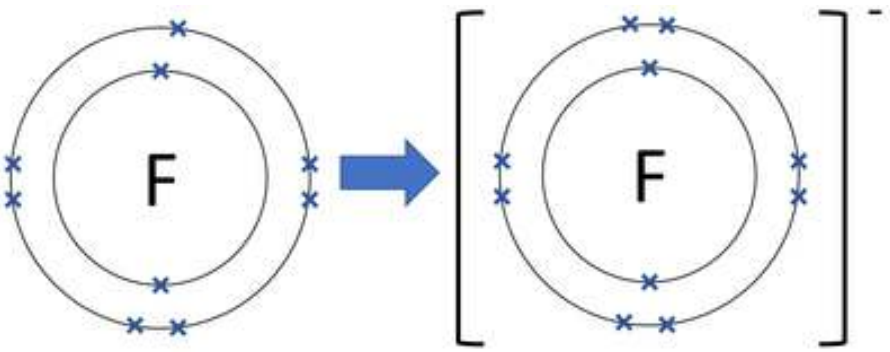
eg Sodium has one electron on its outer shell as it is in group 1 of the periodic table. It will **lose** one electron to become a 1+ ion.



Oxygen has 6 electrons on its outer shell as it is in group 6 of the periodic table. It will **gain** two electrons to become a 2- ion



Fluorine has 7 electrons on its outer shell as it is in group 7 of the periodic table. It will **gain** one electron to become a 1- ion.



Summary:

- *If the atom is in group 1, it will lose an electron to become a 1+ ion*
- *If the atom is in group 2, it will lose two electrons to become a 2+ ion*
- *If the atom is in group 3, it will lose three electrons to become a 3+ ion*
- *If the atom is in group 5, it will gain three electrons to become a 3- ion*
- *If the atom is in group 6, it will gain two electrons to become a 2- ion*
- *If the atom is in group 7, it will gain an electron to become a 1- ion*
- *Group 0 does not form ions as the atoms have full outer shells already*

<u>Forming ionic bonds = transfer of electrons between metal and non-metal atoms</u>

Electrons from the outer shell of the metal atom are **transferred** to the non-metal atom so that both the metal atom and the non-metal atom end up with full outer shells.

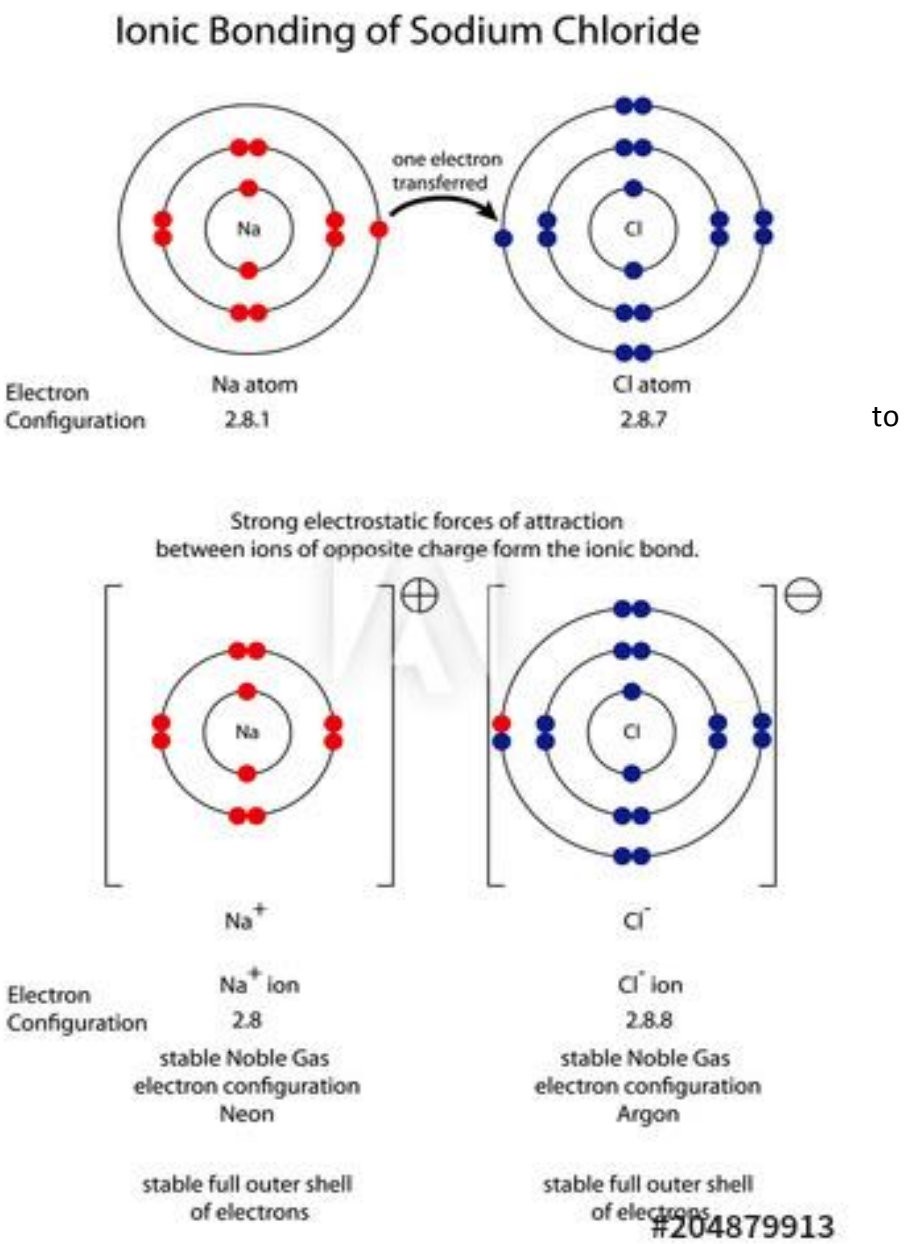
○ **Metal** atoms **lose electrons** to become **positively** charged **ions**

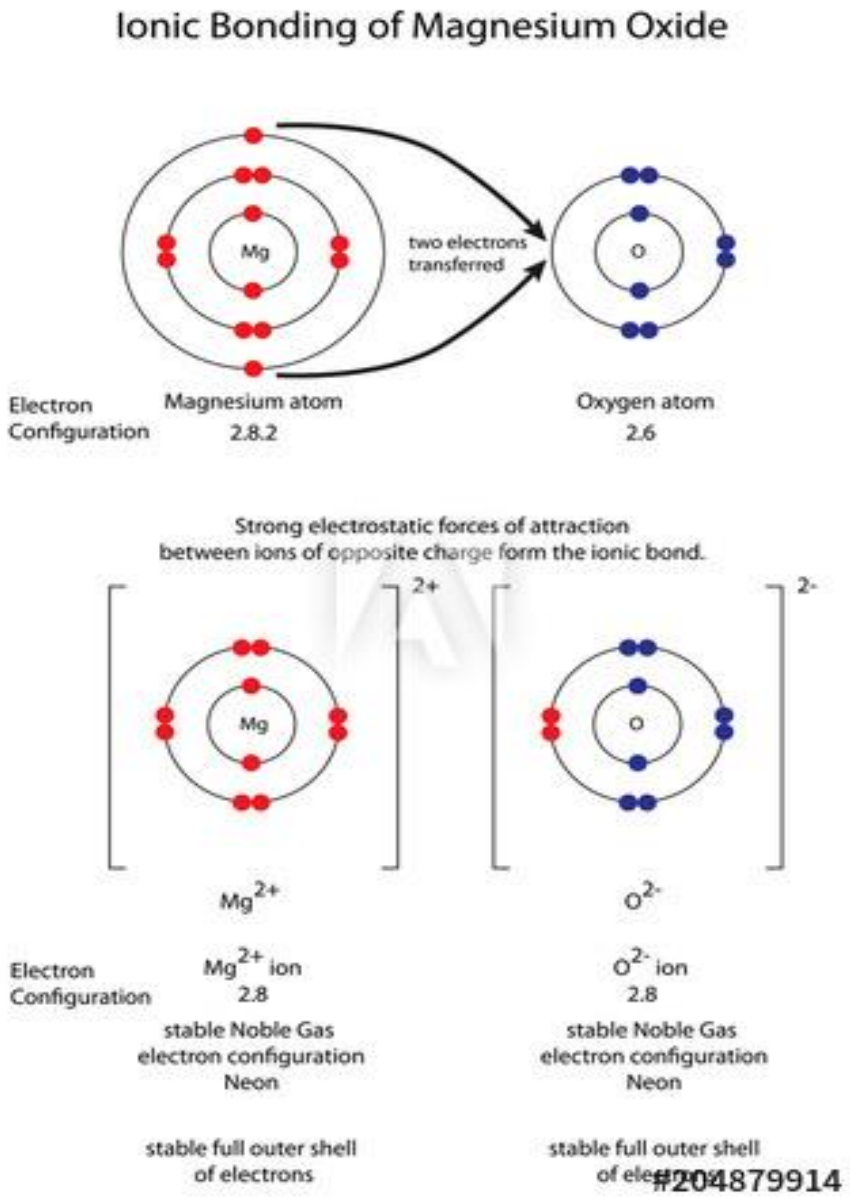
○ **Non-metal** atoms **gain electrons** to become **negatively** charged **ions**

Electron transfer during the formation of an ionic compound can be represented by a **dot and cross diagram**

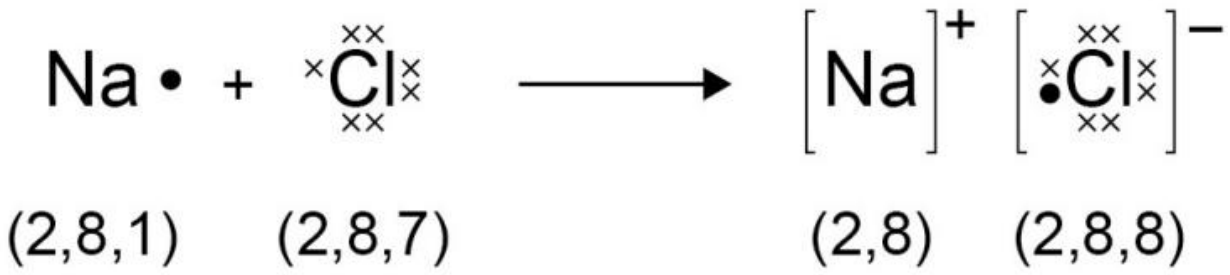
The attraction between the oppositely charged ions is called an electrostatic attraction.

This attraction is the ionic bond.





It is also possible to show **only the outer shell** in the **dot cross diagram**:



The attraction between the oppositely charged ions is called an electrostatic attraction.
This attraction is the ionic bond.

Use the information from above to help you complete the following tasks:

Explain and Understand Questions

1. If an atom loses a negative electron what charge will it then have?

   _____

2. If an atom gains a negative electron what charge will it then have?

   _____

3. Explain why atoms lose or gain electrons

   _____

   _____

4. How many electrons are on the outer shell of a group 1 metal atom?

   _____

5. How many electrons must a group 1 metal atom lose in order to get a full outer shell?

   _____

6. Draw an atom of lithium (top of group 1 in periodic table).

7. Will the lithium atom lose or gain electrons to get a full outer shell?

   _____

8. Draw a lithium ion, include its charge

9. Draw an atom of chlorine (group 7 of periodic table)

10. Draw a chloride ion, include its charge

11. Draw the ionic bonding (transfer of electrons) of lithium chloride

12. What is an ionic bond?

_____

_____

13. Describe, in terms of electrons, what happens when a lithium atom reacts with a chlorine atom to produce lithium chloride.

_____

_____

_____

_____

Extension: Describe, in terms of electrons, what happens when lithium atoms react with oxygen atoms to produce lithium oxide.

_____

_____

_____

_____

_____

_____

Lesson 3 – Ionic compounds and properties

Learning Objectives

- Identify ionic compounds from structures
- Describe the properties of ionic compounds
- Explain when ionic compounds can conduct electricity
- Work out the empirical formula of an ionic compound
- Relate melting points of ionic compounds to forces between ions

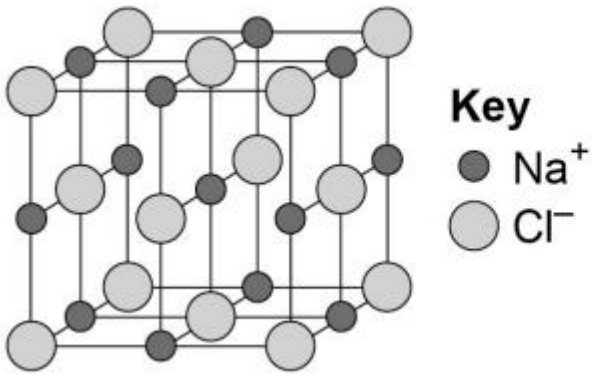
Explain and understand Information

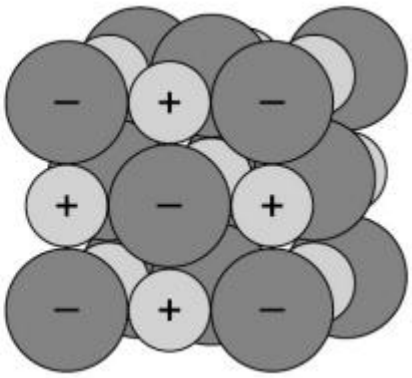
An ionic compound is a **giant structure of ions**.

Ionic compounds are held together by **strong electrostatic forces of attraction** between **oppositely charged ions**. Since the structure is **3D** these forces act in all directions in the **lattice.** This is strong ionic bonding.

The giant ionic lattice structure contains positively charged metal ions and negatively charged non-metal ions.

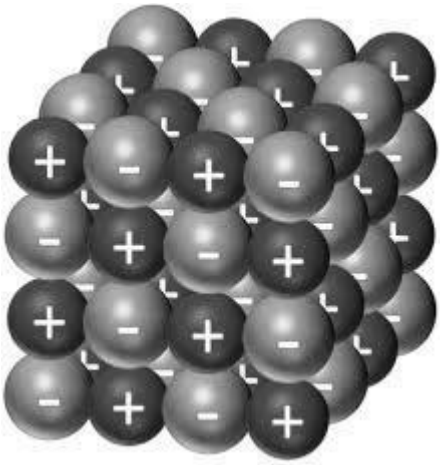
The structure of sodium chloride can be **represented** in the following forms:





To **describe** the structure and bonding in sodium chloride:

Sodium chloride has a giant lattice structure of Na^+ and Cl^- ions held together by electrostatic attraction



Properties of ionic compounds:

- giant structure/lattice structure
- crystalline / hard
- high melting point / solid
- high boiling point
- conductor (of electricity) when dissolved **in** water
- conductor (of electricity) when molten
- soluble in water

Ionic compounds have **high melting points and high boiling points** because of the **large amounts of energy** needed to **break the many strong bonds**.
When **melted or dissolved in water**, ionic compounds **conduct electricity** because the **ions** are free to **move** and so charge can flow.

Use the information from above to help you complete the following tasks:

Explain and Understand Questions

1. Describe the structure of an ionic compound

   _____

   _____

2. How are ionic compounds held together?

   _____

   _____

3. What does a giant ionic lattice contain?

   _____

   _____

4. Describe the structure and bonding in sodium chloride

   _____

_____

5. Name 3 properties of ionic compounds

_____

_____

_____

6. Explain why ionic compounds have high melting points

_____

_____

7. What has to be done to ionic compounds to enable them to conduct electricity?

_____

_____

8. What can the ions do that allows them to conduct electricity?

_____

_____

Lesson 4 – Covalent Bonding.

Learning Objectives

- Recognise substances made of small molecules from their formula.
- Draw dot and cross diagrams for covalent bonding in small molecules.
- Deduce molecular formulae from models and diagrams.

Explain and understand information

Atoms bond so that they can have a **full outer shell** of electrons, and be **stable**.

During chemical bonding only the **electrons** are involved.

Forming covalent bonds = sharing of electrons between non-metal atoms

- **Non-metals** react with other **non-metals**.
- When they react the atoms **share pairs of electrons** and they form **covalent bonds**.
- Covalent bonds between atoms are **strong**.
- The sharing of electrons ensures that each atoms has a full outer shell of electrons.

Covalently bonded substances may consist of **small molecules** for example:

CO_2 = carbon dioxide

H_2O = water

NH_3 = ammonia

CH_4 = methane

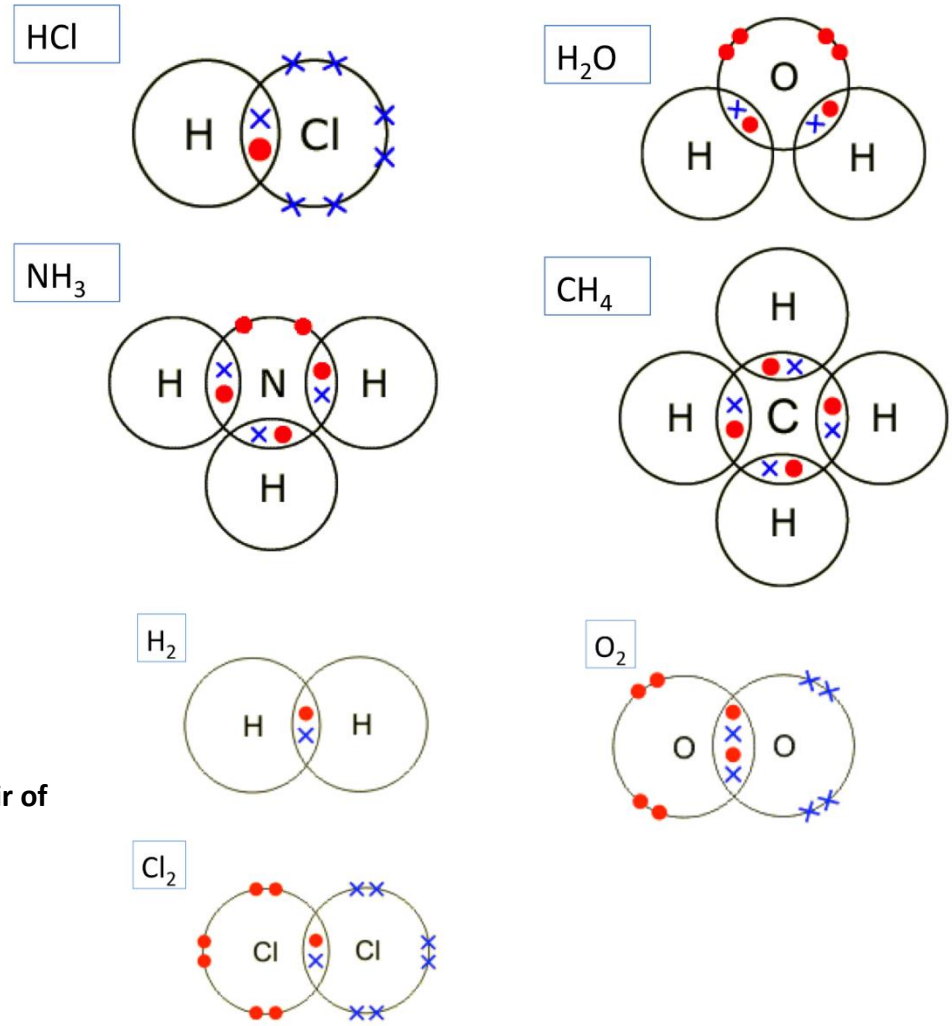
HCl = hydrochloric acid

Cl_2 = chlorine

H_2 = hydrogen

O_2 = oxygen

N_2 = nitrogen



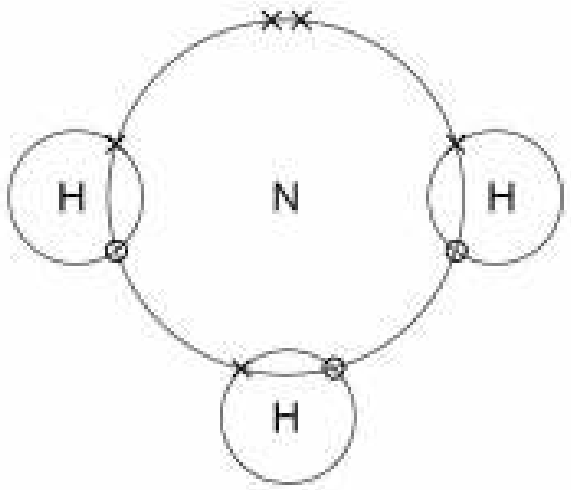
A covalent bond is a shared pair of electrons

The covalent bonds in molecules (and giant covalent structures) can be represented in the following forms:

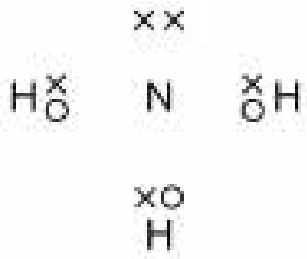
- Dot cross diagrams with electrons sitting on outer shells
- Dot cross diagrams with only electrons showing (no shells)
- A single short line (displayed formula)
- 3D stick and ball models

eg:

For ammonia (NH_3)



and/or



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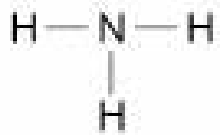

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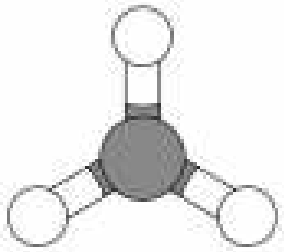


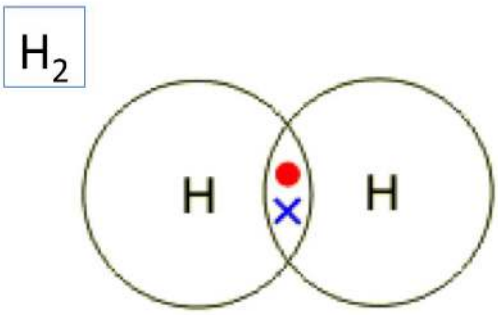
You need to be able to:

• draw dot and cross diagrams for the molecules of hydrogen, chlorine, hydrogen chloride, water, oxygen, nitrogen, ammonia and methane

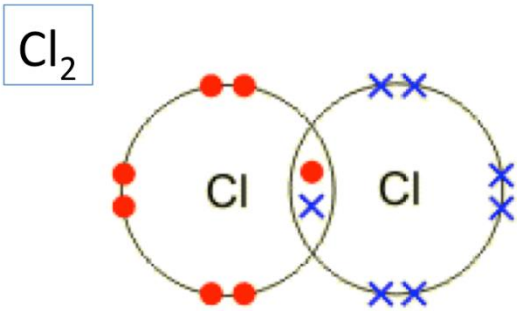
• represent the covalent bonds in small molecules using a line to represent a single bond

Why share electrons in covalent bonding? 3 examples:

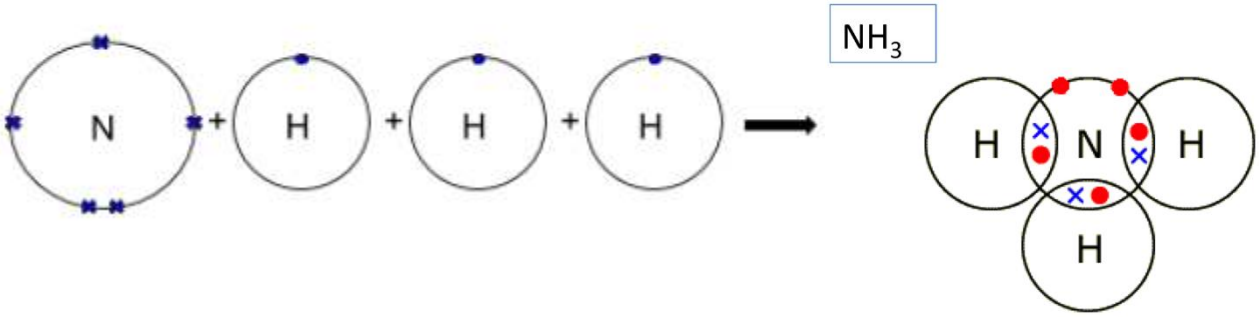
1. Hydrogen atoms prefer to travel around with another hydrogen atom. This is because both hydrogen atoms have only 1 electron in their outer shells. When the two hydrogen atoms form a covalent bond they each share their electron which means that they both now have 2 electrons in their outer shells. This outer shell is now full.



2. Chlorine atoms prefer to travel around with another chlorine atom. This is because both chlorine atoms have 7 electrons in their outer shells. When the two chlorine atoms form a covalent bond they each share an electron which means that they both now have 8 electrons in their outer shells and so both have full outer shells.

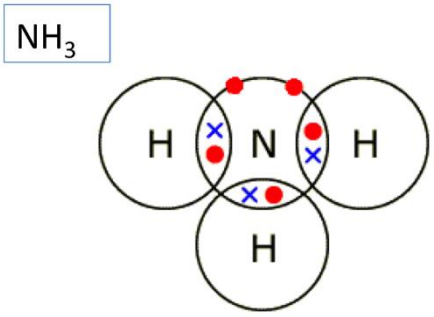


3. When nitrogen reacts with hydrogen to make ammonia. Electrons are shared. Nitrogen needs 3 more electrons to fill its outer shell. Hydrogen needs only 1 more electron to fill its outer shell. 3 separate hydrogen atoms each share their 1 electron with one of nitrogen's electrons. This way, nitrogen gains 3 electrons to have 8 electrons in its outer shell and each hydrogen gains one more to have 2 electrons in each of their outer shells.



Evaluate the different ways of representing covalent bonding

Pros of dot-cross diagrams:

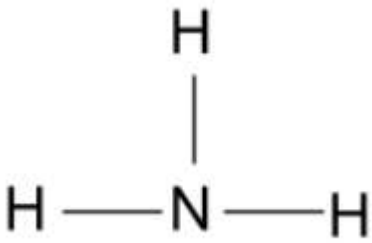


1. Can easily see where the electrons have come from
2. Easy to draw

Cons:

1. Don't show the relative sizes of the atoms
2. Can't see how the atoms are arranged in 3D space

Pros of displayed formula form:

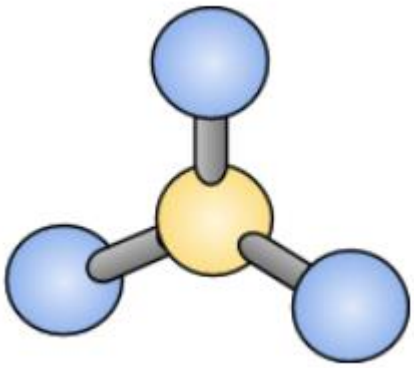


1. Quick and easy to draw (useful for large molecules)

Cons:

1. Can't see how the atoms are arranged in 3D space
2. Can't see where the shared electrons have come from

Pros of 3D stick and ball models:



1. They show the arrangement of molecules in 3D space

Cons:

1. difficult to draw and get confusing
2. don't know where the electrons have come from

A covalent bond is a shared pair of electrons

Explain and Understand Questions

1. What type of elements form covalent bonds?

   _____

2. What is a covalent bond?

   _____

   _____

3. Some small molecules have covalent bonds.

Draw a dot cross diagram to show the bonding in a molecule of hydrogen

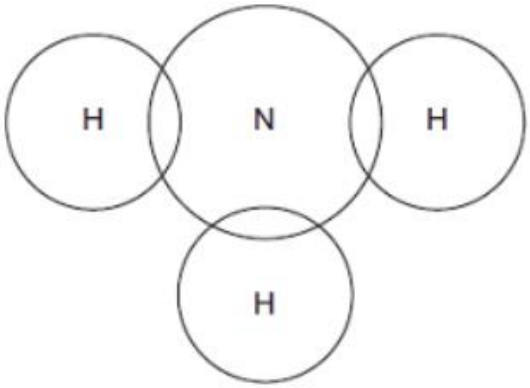
5. Draw a dot cross diagram to show the bonding in a molecule of water

6. What are the advantages of dot and cross diagrams to represent covalently bonded molecules?

_____

_____

7. Complete the dot cross bonding diagram for ammonia. Show **only** electrons in the outer energy level of each atom.



Lesson 5 – Properties of small molecules

Learning Objectives

- Identify small molecules from formulae.
- Explain the strength of covalent bonds.
- Relate the intermolecular forces to the bulk properties of a substance.

Explain and understand information

Properties of simple molecular substances.

Melting and boiling are **changes of state**.

Energy must be transferred to a substance to make it melt or boil. This energy overcomes the attractive forces between the molecules (intermolecular forces) in the substance:

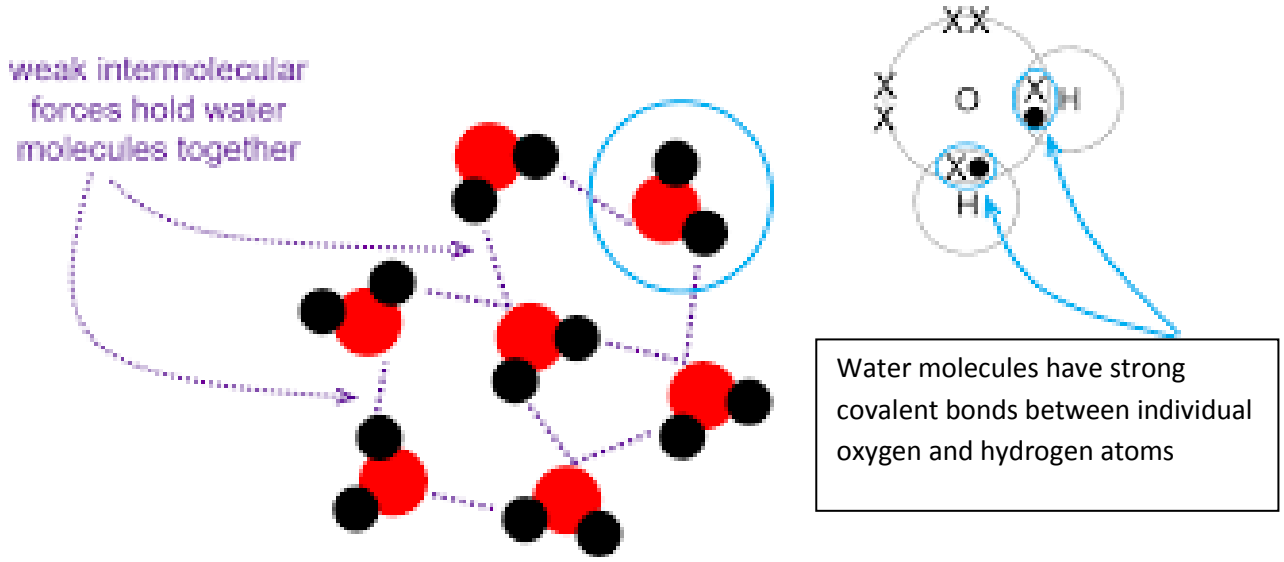
- some intermolecular forces of attraction are overcome during melting, allowing molecules to move over each other
- more of the intermolecular forces of attraction are overcome during boiling, allowing the molecules to move freely away from each other

The more energy that is needed, the higher the melting point or boiling point.

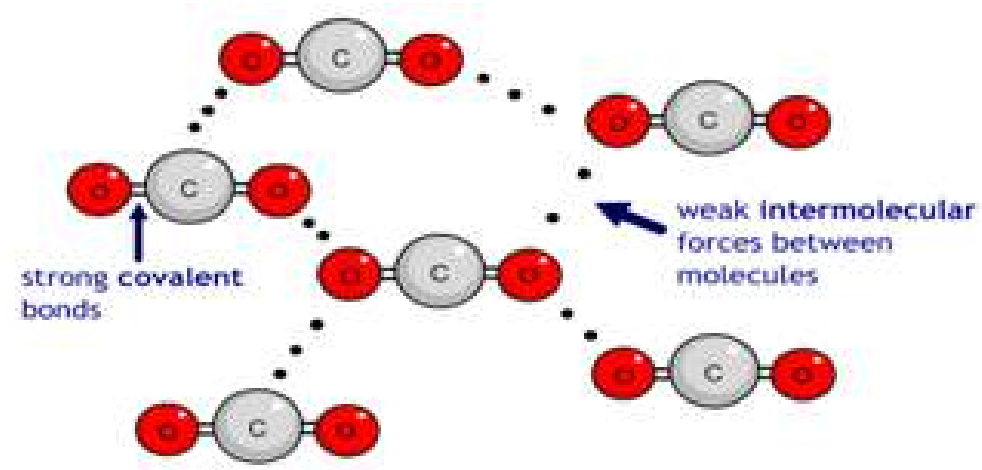
Substances that consist of **small molecules** are usually **gases or liquids** that have relatively **low melting points and boiling points.**

When something is melted or boiled, the molecules are separated from one another. For example, if you had a bowl of water (lots of H_2O molecules) and you boil the water, the H_2O molecules will all separate from one another.

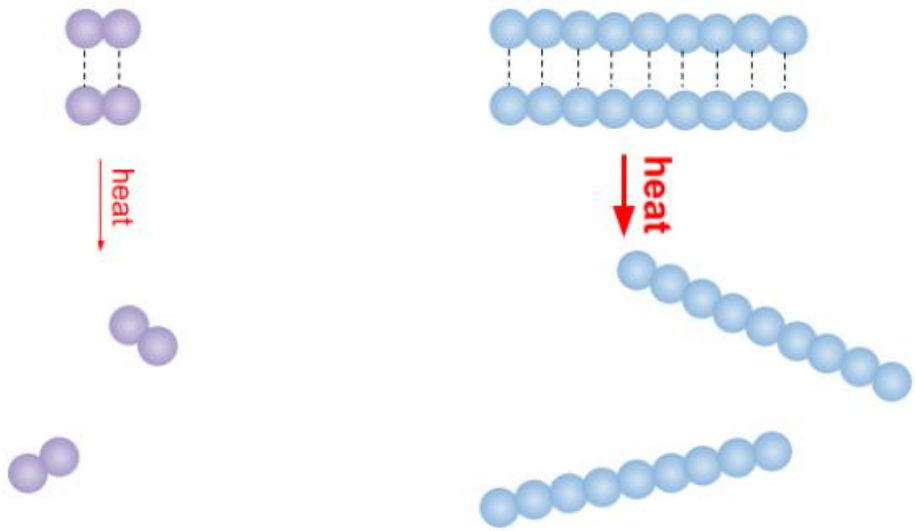
Simple molecular substances have only **weak forces between the molecules** (**intermolecular forces**).
It is these intermolecular forces that are overcome, not the covalent bonds that hold the atoms of the molecule together, when the substance melts or boils.



Carbon dioxide is a small covalent molecule. It has a low boiling point because it takes very little energy to separate the molecules from one another by breaking the weak intermolecular forces between molecules. The strong covalent bonds between atoms are not broken.



DO NOT get this confused with breaking covalent bonds between the atoms. The carbon dioxide molecules are still intact.



Trends in melting and boiling points of simple covalent substances

The bigger the molecule, the stronger the intermolecular forces, so the higher the melting and boiling points.

When something is heated, the **heat breaks** these **intermolecular forces**. The larger the molecules, the more intermolecular forces and the more heat is needed to break all of the intermolecular forces.

Therefore, **small molecular substances have low melting/boiling points** (need little heat energy to break the intermolecular forces) and **larger molecular substances have high melting/boiling points** (need lots of heat energy to break all of the intermolecular forces).

Electrical conductivity

A substance can conduct electricity if:

- it contains charged particles, and
- these particles are free to move from place to place

Small molecules have no overall electric charge, so they **cannot conduct electricity**, even when liquid or dissolved in water.

In covalent substances, there are **no charged particles** that are free to move around, therefore they **do not conduct electricity**!

Small covalently bonded molecules with low melting and boiling points include:

CO_2 = carbon dioxide

Cl_2 = chlorine

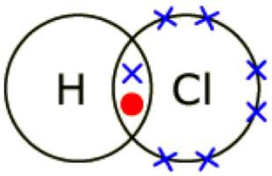
H_2 = hydrogen

O_2 = oxygen

N_2 = nitrogen

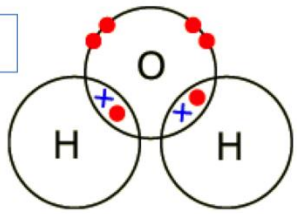
HCl

hydrochloric acid



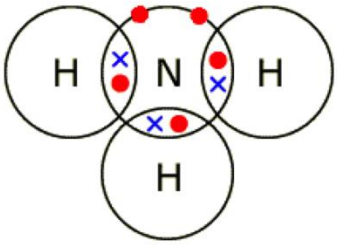
H_2O

water



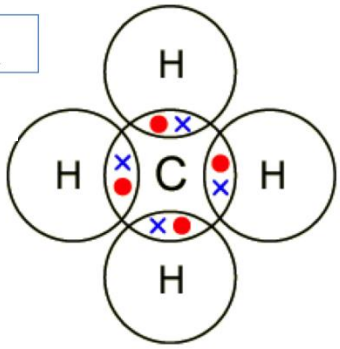
NH_3

ammonia



CH_4

methane



Use the information from above to help you complete the following tasks:

Explain and Understand Questions

1. Give examples of changing state

_____

2. What holds covalent molecules together?

_____

3. Why are simple covalent molecules usually gases or liquids?

_____

_____

4. When a molecular substance like water boils, what is broken…. the covalent bonds between the hydrogen and oxygen atoms or the intermolecular bonds between water molecules?

_____

5. Why does carbon dioxide have a low boiling point?

_____

_____

6. Describe the trend shown by small covalent molecules linking molecule size to melting point

_____

_____

7. With reference to heat energy and intermolecular forces, explain the link between size of molecule and melting point

_____

_____

8. Explain why covalent substances do not conduct electricity

_____

_____

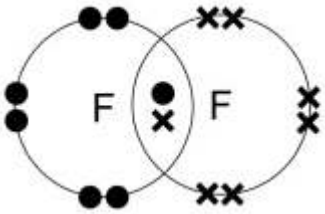
9. Name the following small molecules:

CH_4 _____

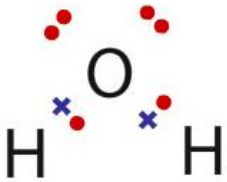
NH_3 _____

HCl _____

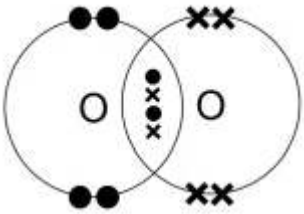
10. Name the following small molecules and write the formulae of each:



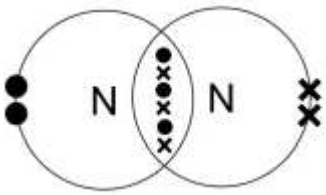
a) _____



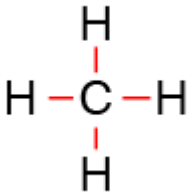
b) _____



c) _____



d) _____



e) _____

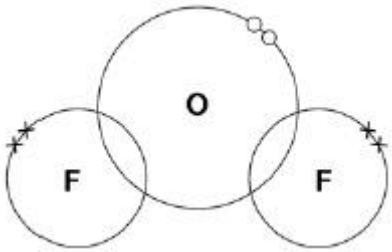
11.

This question is about oxygen.

(a) One oxygen atom shares one pair of electrons with each fluorine atom in oxygen difluoride (OF_2).

Complete the dot and cross diagram of oxygen difluoride below.

You should show only the electrons in the outer shells.



(2)

(b) Oxygen difluoride (OF_2) has a melting point of – 224 °C and a boiling point of –145 °C

What is the state of oxygen difluoride at room temperature?

Explain your answer in terms of structure and bonding.

_____

_____

_____

_____

_____

_____

_____

(4)

(Total 6 marks)

Lesson 6 – Moles (HT only)

Learning Objectives

- Describe the measurement of amounts of substances in moles.
- Calculate the number of moles in a given mass.
- Calculate the mass of a given number of moles.

Explain and Understand

It is often useful in chemistry to know the number of atoms, ions or molecules involved in a chemical reaction. The mole is the unit for amount of substance. It is abbreviated to mol.

1 mol is the amount of substance that contains the same number of particles as there are atoms in 12.0 g of carbon-12.

Since atoms are so very small and have very little mass, the number of atoms in 12.0 g of carbon-12 is huge. It is equal to the Avogadro constant:

The Avogadro constant = 6.022×10^{23} atoms per mole.

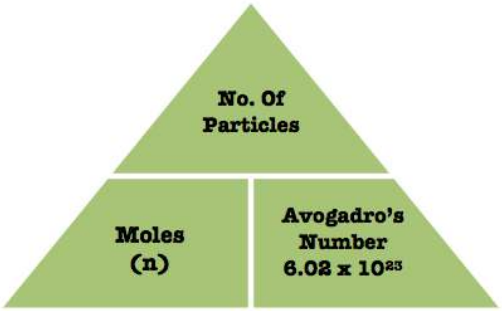
Atoms are too small and there are far too many to be counted so the mole is a very useful unit. To help you visualise the number of particles in a mole try to imagine the Earth covered in a mole of fizzy drink cans- how deep would it be???

200 miles deep!! That is nearly the same distance as Plymouth to London!!

Calculating the number of particles

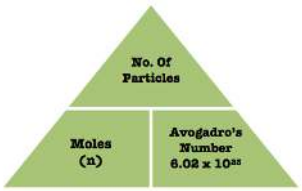
The number of particles of a substance can be calculated using:

- the Avogadro constant
- the amount of substance in mole

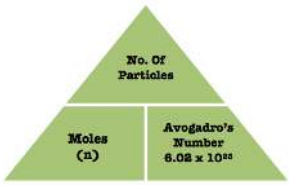


number of particles = Avogadro constant × amount (mol)

Teacher modelling – calculating the number of particles using Avogadro's constant and the number of moles. Complete the calculations.

	**I do!**	**We do!**	**You do!**
	How many molecules are there in 0.3 moles of carbon dioxide?	How many molecules are there in 3 moles of water?	How many atoms are there in 10 moles of copper?
1. Write out the equation for calculating the number of particles.	No. of particles = moles x N_A		
2. Substitute in the number of moles and Avogadro's constant.	$0.3 \times 6.02 \times 10^{23}$		
3. Write your answer.	1.806×10^{23}		

Teacher modelling – calculating the number of moles using Avogadro's constant and the number of particles. Complete the calculations.

	**I do!**	**We do!**	**You do!**
	How many moles are in 3.01×10^{23} water molecules?	How many moles are in 6.02×10^{24} water molecules?	How many moles are in 2.05×10^{22} water molecules?
1. Write out the equation for calculating the amount of moles.	Moles = No. of particles / N_A		
2. Substitute in the number of moles and Avogadro's constant.	$\frac{3.01 \times 10^{23}}{6.02 \times 10^{23}}$		
3. Write your answer.	= 0.5 mol		

Molar mass

The **molar mass** of a substance is the mass of one mole of that substance. It is equal to:

- the relative atomic mass or A_r **in grams** for atoms and metals
- the relative formula mass or M_r **in grams** for molecules and compounds

The table shows some examples.

Substance	Relative mass	Mass of 1 mole (give units)
H_2O	18	18 g
HCl	36.5	36.5 g
NaOH	40	40 g
CO_2	44	44g
MgO	40	40g

Masses from moles

The mass of a substance in grams can be calculated from its amount in moles and its molar mass:

mass (g) = molar mass (g/mol) × amount (mol)

Example

Calculate the mass of 0.25 mol of carbon dioxide molecules, CO_2. (Relative atomic masses: C = 12.0, O = 16.0)

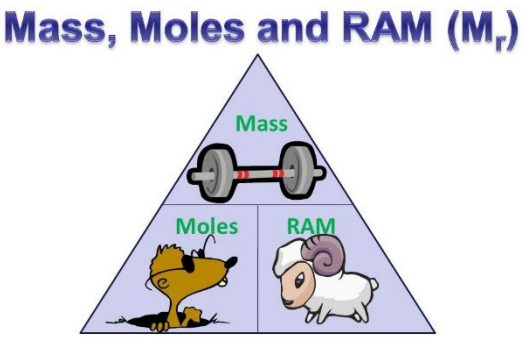
relative formula mass, M_r = 12.0 + (2 × 16.0) = 44.0

molar mass = 44.0 g/mol

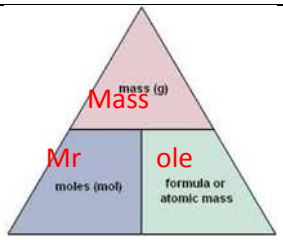
= 11.0 g

The calculation is the same if a substance is a metal or exists as separate atoms, but its A_r is used to work out the molar mass instead of an M_r.

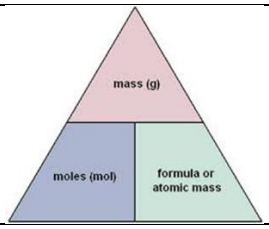
You can use the following formula triangle to support you with these calculations and always remember the '**Mr' 'Mole' lives under the 'Mass'.**



Teacher modelling – calculating the number of moles from molar mass. Complete the calculations.

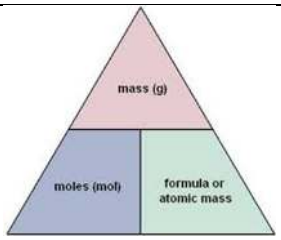
	**I do!**	**We do!**	**You do!**
	How many moles are in 180g of H_2O?	How many moles are in 10g of $CaCO_3$?	How many moles are in 10g of H_2O?
1. Write out the equation for calculating the amount of moles.	Moles = mass (g) molar mass (g/mol)		
2. Calculate the relative mass of the substance	O = 16 H = 2 x 1 =2 = 18		
3. Substitute the mass in grams and relative mass into the equation.	= 180g / 18		
4. Write your final answer.	10 mol		

Teacher modelling – calculating the molar mass from the number of moles and mass on grams. Complete the calculations

	I do!	**We do!**	**You do!**
	What is the molar mass if 5g of a substance is 0.1 moles?	What is the molar mass if 0.1g of a substance is 0.02 moles?	What is the molar mass if 2.8g of a substance is 0.5 moles?
1. Write out the equation for calculating the molar mass.	molar mass = mass/mol		
2. Substitute the mass in grams and moles into the equation.	5g / 0.1		
3. Write your final answer.	50 g/mol		

Teacher modelling – calculating the mass from the number of moles and relative mass (molar mass). .

Complete the calculations

	**I do!**	**We do!**	**You do!**
	What mass is 0.5 moles of H_2O?	What mass is 2 moles of CO_2?	What mass is 0.1 moles of H_2O?
1. Write out the equation for calculating the mass in grams.	mass = moles x molar mass		
2. Calculate the relative mass of the substance	O = 16 H = 2 x 1 =2 = 18		
3. Substitute the moles and relative mass into the equation.	= 0.5 X 18		
4. Write your final answer.	9 g		

Use the information from above to help you to complete the following tasks:

Explain and Understand Questions

Moles and molar mass

1. a) Why do chemists need a unit like the mole? _____

   b) What is the symbol for mole? _____

   c) How many particles are in one mole of a substance? _____

   d) What is this number called? _____

   e) Which of these contain the same number of particles. Circle your answers.

   1 mol copper; 0.5 mol copper oxide; 1 mol water; 2 mol chlorine gas; 1 mol hydrogen gas

2. One mole of any substance contains 6.02×10^{23} particles.

   a) Use your calculator to work out how many particles there are in:

   (i) 0.5 mol water _____

   (ii) 2 mol iron _____

   (iii) 10 mol copper(II) sulfate _____

(iv) 0.1 mol zinc _____

(v) 5 mol sodium chloride _____

3. a) Use the A_r values given below to calculate the molar mass of these elements (*Hint* – units are g/mol):

(i) sodium _____

(ii) phosphorus _____

(iii) chromium _____

(iv) lithium _____

A_r values: Na = 23; P = 31; Cr = 52; Li = 7

(b) Use the A_r values below to calculate the molar mass of:

(i) copper(II) chloride, $CuCl_2$ _____

(ii) sulfur dioxide, SO_2 _____

(iii) methane, CH_4 _____

(iv) calcium hydroxide, $Ca(OH)_2$ _____

(v) copper(II) sulfate crystals, $CuSO_4{\cdot}5H_2O$ _____

A_r values: Cu = 64; C = 12; H = 1; O = 16; Ca = 40; S = 32; Cl = 35.5

Converting between moles and mass

1. How many moles are there in each of the following amounts of substances?

a) 48 g magnesium _____

b) 6 g magnesium _____

c) 9.8 g sulfuric acid H_2SO_4 _____

d) 20 g calcium carbonate $CaCO_3$ _____

e) 400 g sodium hydroxide NaOH _____

f) 9.2 g ethanol, C_2H_5OH _____

g) 4 g hydrogen molecules H_2 _____

h) 4 g hydrogen atoms _____

i) 1 kg hydrogen atoms _____

j) 1 kg hydrogen molecules H_2 _____

A_r values: Mg = 24; H = 1; S = 32; O = 16; Ca = 40; C = 12; Na = 23

2. What is the mass in grams of:

a) 2 mol argon gas _____

b) 0.5 mol copper _____

c) 5 mol ethanol, C_2H_5OH _____

d) 0.1 mol calcium carbonate $CaCO_3$ _____

e) 0.25 mol calcium hydroxide $Ca(OH)_2$ _____

f) 0.01 mol sulfuric acid H_2SO_4 _____

g) 3 mol nitrogen atoms, N _____

h) 3 mol nitrogen molecules, N_2 _____

i) 0.5 mol chlorine molecules, Cl_2 _____

j) 10 mol lead _____

A_r values: Ar = 40; Cu = 64; C = 12; H = 1; O = 16; Ca = 40; S = 32; N = 14; Cl = 35.5; Pb = 207

Practice Questions

Q1. Iron is an essential part of the human diet. Iron(II) sulfate is sometimes added to white bread flour to provide some of the iron in a person's diet.

(a) The formula of iron(II) sulfate is $FeSO_4$

Calculate the relative formula mass (M_r) of $FeSO_4$

Relative atomic masses: O = 16; S = 32; Fe = 56.

_____

_____

The relative formula mass (M_r) = _____

(2)

(b) What is the mass of one mole of iron(II) sulfate? Remember to give the unit.

_____

(1)

Q3.(a) Calculate the mass of one atom of sodium.

Avogadro constant = 6.02×10^{23} per mole.

Give your answer to 3 significant figures.

_____

_____

_____

_____

_____

Mass of one atom of sodium = _____ g

(2)

Q4.(a) Ethene is a product of cracking and the relative formula mass (M_r) of ethene = 28

Calculate the number of moles of ethene (C_2H_4) in 50.4 kg

Give your answer in standard form.

_____

_____

_____

_____

_____

Numbers of moles = _____

(3)

Q5. (a) Calculate the **number of molecules** in 14 g of carbon dioxide CO_2.

Give your answer in standard form.

Relative atomic masses (A_r): C = 14; O = 16

_____

_____

_____

_____

_____

Answer = _____ molecules

(4)

Q6. Fertilisers are formulations.

(a) A bag of fertiliser contains 14.52 kg of ammonium nitrate (NH_4NO_3).

Relative formula mass (M_r): NH_4NO_3 = 80

Calculate the number of moles of ammonium nitrate in the bag of fertiliser.

Give your answer in standard form to 2 significant figures.

_____

_____

_____

_____

_____

_____

_____

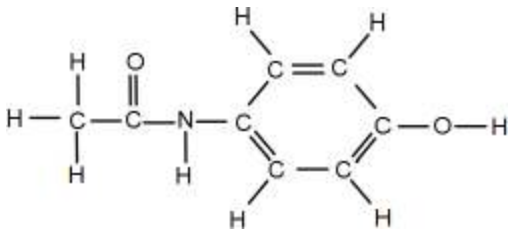
_____

Moles of ammonium nitrate = _____ mol

(4)

Q7.(a) The main ingredient in Aqamed is a painkiller called paracetamol.

The figure below represents a molecule of paracetamol.



Give the molecular formula of paracetamol.

Calculate its relative formula mass (M_r).

Relative atomic masses (A_r): H = 1; C = 12; N = 14; O = 16

Molecular formula _____

Relative formula mass _____

_____

M_r = _____

(2)

(b) Aspirin is a medicine for use by adults.

An aspirin tablet contains 300 mg of acetylsalicylic acid.

Calculate the number of moles of acetylsalicylic acid in one aspirin tablet.

Give your answer in standard form to three significant figures.

Relative formula mass (M_r) of aspirin = 180

_____

_____

_____

_____

_____

_____

Number of moles = _____

(4)

TOTAL = _____ / 24

Lesson 7 – Amounts of substances in equations (HT only)

Learning Objectives

- Calculate the masses of substances in a balanced symbol equation.
- Define Avogadro's constant
- Calculate the masses of reactants and products from balanced symbol equations.
- Calculate the mass of a given reactant or product.

Explain and Understand

The mass of product formed in a reaction depends upon the mass of the limiting reactant (the one which runs out first). This is because no more product can form when the limiting reactant is all used up.

Stoichiometry of a reaction is the ratio of the amounts of each substance in the balanced equation and this need to solve reacting masses calculations.

Reacting mass calculations

The maximum mass of product formed in a reaction can be calculated using:

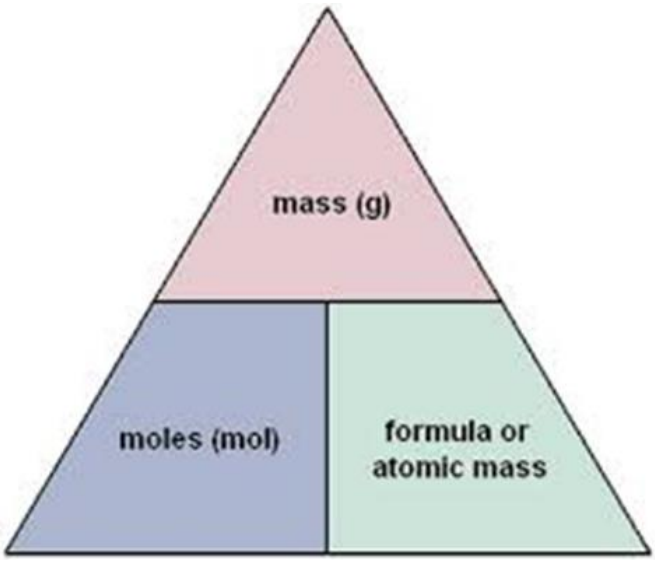
- the balanced equation
- the mass of the limiting reactant, and
- the A_r or M_r values of the limiting reactant and the product

Example

x3.0 g of carbon reacts completely with excess oxygen to form carbon dioxide:

$C(s) + O_2(g) \rightarrow CO_2(g)$

Calculate the maximum mass of carbon dioxide that can be produced. (A_r of C = 12.0, M_r of CO_2 = 44.0)



amount = mass x molar mass

amount of carbon = 3.0 / 12.0

= 0.25 mol

Looking at the equation, 1 mol of C forms 1 mol of CO_2, so 0.25 mol of C forms 0.25 mol of CO_2

mass = molar mass × amount

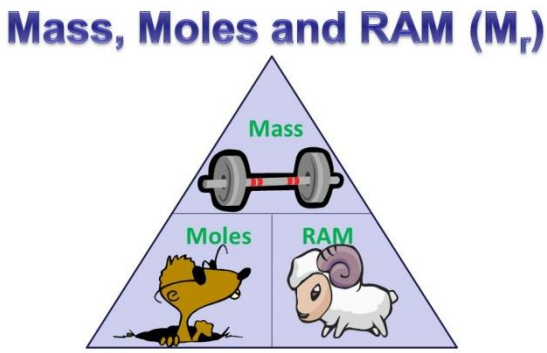
mass of CO_2 = 44.0 × 0.25

= 11.0 g

It can be deduced or worked out using masses found by experiment

Teacher modelling – using stoichiometry to calculate the mass of a substance reacting or being produced given the mass of another substance

In the equation $FeO_2 + 2CO \rightarrow Fe + 2CO_2$ what mass of iron in grams is produced from 880 g of iron(IV) oxide?

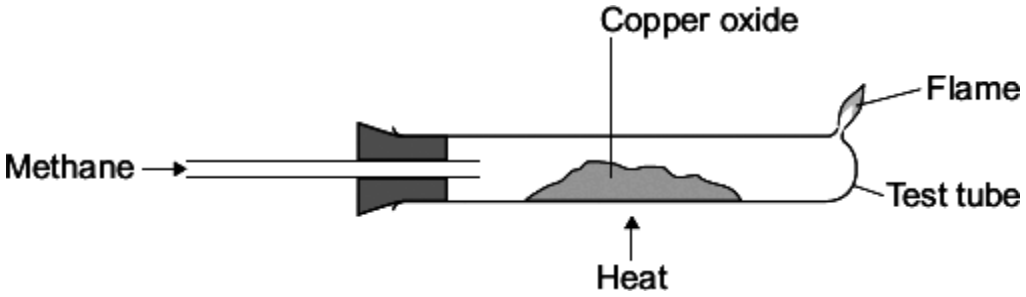


Mr mole lives under the mass!

1. Circle the substances involved in the calculation and write out the formula and coefficient.	1 FeO_2	1 Fe
2. Write in your known mass	880g	
3. Calculate the relative masses for each substance	Fe = 56 O = 2 x 16 = 32 = 88	56
4. Calculate the moles for the known substance	880/88 = 10	
5. Use the equation to determine the mole ratio	10	10
6. Calculate the moles for the unknown substance		10 x 56 =
6. Calculate the unknown mass which is your answers.		56 g

Practice Questions

Q1. An experiment was done on the reaction of copper oxide (CuO) with methane (CH_4).



(a) (i) Calculate the relative formula mass (M_r) of copper oxide (CuO).

Relative atomic masses (A_r): O = 16; Cu = 64.

_____

_____

_____

Relative formula mass (M_r) = _____

(2)

(ii) Calculate the percentage of copper in copper oxide.

_____

_____

_____

Percentage of copper = _____ %

(2)

(iii) Calculate the mass of copper that could be made from 4.0 g of copper oxide.

_____

_____

Mass of copper = _____ g

(1)

Q2. (a) The equation for the reaction between zinc and copper oxide is:

$$Zn + CuO \longrightarrow ZnO + Cu$$

1.59 g of copper oxide reacted.

Calculate the mass of copper produced.

Relative atomic masses (A_r): Cu = 63.5 O = 16 Zn = 65

_____

_____

_____

_____

Mass of copper produced = _____ g

(3)

Q3. (a) Iron(III) chloride can be produced by the reaction shown in the equation:

$$2\ Fe + 3\ Cl_2 \rightarrow 2\ FeCl_3$$

Calculate the maximum mass of iron(III) chloride ($FeCl_3$) that can be produced from 11.20 g of iron.

Relative atomic masses (A_r): Cl = 35.5; Fe = 56.

_____

_____

_____

_____

Maximum mass of iron(III) chloride = _____ g

(3)

Q4. (a) $C_{21}H_{44}$ can be cracked to produce ethene.

$$C_{21}H_{44} \rightarrow 3C_2H_4 + C_{15}H_{32}$$

Relative formula mass (M_r) of $C_{21}H_{44}$ = 296

Calculate the mass of $C_{21}H_{44}$ needed to produce 50.4 kg of ethene.

_____

_____

_____

_____

_____

Mass = _____ kg **(3)**

Q5. Magnesium reacts with iron chloride solution.

$$3\ Mg + 2\ FeCl_3 \longrightarrow 2\ Fe + 3\ MgCl_2$$

(a) 0.120 g of magnesium reacts with excess iron chloride solution.

Relative atomic masses (A_r): Mg = 24 Fe = 56

Calculate the mass of iron produced, in mg

_____

_____

_____

_____

_____

_____

_____

_____

_____

_____

Mass of iron = _____ mg

(5)

TOTAL = _____ / 19

Lesson 8 – Using masses to balance equations (HT only)

Learning Objectives

- Convert masses in grams to amounts in moles.
- Balance an equation given the masses of reactants and products.
- Change the subject of a mathematical equation.

Explain and Understand

Stoichiometry of a reaction

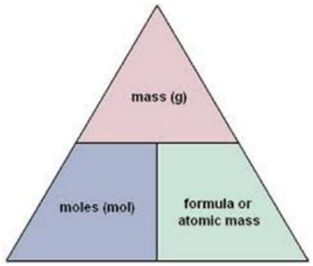
The stoichiometry of a reaction is the ratio of the amounts of each substance in the balanced equation. It can be deduced or worked out using masses found by experiment. If you are given the masses of the reactants and products you can then convert it into moles. Once you have calculated the moles you can treat it as a ratio problem and then convert it into the simplest whole number ratio just like you would do in maths!

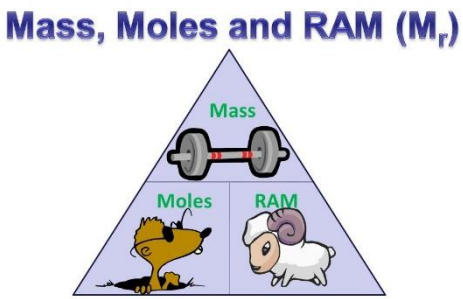
As you did in the previous lesson it is really important to approach this problem in a methodical way. The next few pages give you an opportunity to do this. Instructions will take you through calculating the moles for each substance and then how to convert it into the simplest whole number ratio.

Unfortunately there are not any previous exam style questions on this topic therefore the explain and understand question section is slightly larger than normal in order to give you the opportunity to practise this skill.

Teacher modelling – using masses to construct a balanced symbol equation

4.40 g of propane (C_3H_8) reacts with 16.0 g of oxygen (O_2) to produce 13.2 g of carbon dioxide (CO_2) and 7.20 g of water (H_2O)



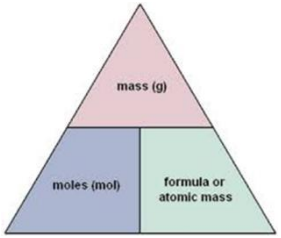


Mr mole lives under the mass!

MASS

Mr Mole

1. Construct the unbalanced symbol equation	C_3H_8	+ O_2	➔ CO_2	+ H_2O
2. Write in the mass (g) for each substance	4.40g	16.0g	13.2g	7.20g
3. Calculate the R.F.M for each substance	44	32	44	18
4. Calculate the moles for each substance	4.4 / 44 = 0.1	16 / 32 = 0.5	13.2 / 44 = 0.3	7.2 / 18 = 0.4
5. Divide each mole by the smallest mole value to get the equation ratio *(0.1)*	0.1 / 0.1 = 1	0.5 / 0.1 = 5	0.3 / 0.1 = 3	0.4 / 0.1 = 4
*6. *You may need to X10, X2 etc to ensure that the equation is the simplest, whole number ratio.*	NA	NA	NA	NA
7. Write out the full balanced equation.	**1**C_3H_8	+ **5**O_2	➔ **3**CO_2	+ **4**H_2O

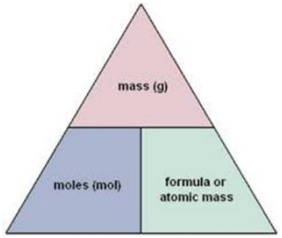


Teacher modelling – using masses to construct a balanced symbol equation

46.0 g of ethanol (C_2H_5OH) reacts with 112.0 g of oxygen (O_2) to produce 88.0 g of carbon dioxide (CO_2) and 54.0 g of water (H_2O).

1. Construct the unbalanced symbol equation	C_2H_5OH +	O_2 →	CO_2 +	H_2O
2. Write in the mass (g) for each substance				
3. Calculate the R.F.M for each substance				
4. Calculate the moles for each substance				
5. Divide each mole by the smallest mole value to get the equation ratio				
*6. *You may need to X10, X2 etc to ensure that the equation is the simplest, whole number ratio.*				
7. Write out the full balanced equation.				

Teacher modelling – using masses to construct a balanced symbol equation



79.2 g of $C_{14}H_{30}$ is cracked to produce 40.0 g of C_7H_{16}, 17.6 g of C_3H_8 and 22.4 g of C_2H_4.

1. Construct the unbalanced symbol equation	$C_{14}H_{30}$ →	C_7H_{16} +	C_3H_8 +	C_2H_4
2. Write in the mass (g) for each substance				
3. Calculate the R.F.M for each substance				
4. Calculate the moles for each substance				
5. Divide each mole by the smallest mole value to get the equation ratio				
*6. *You may need to X10, X2 etc to ensure that the equation is the simplest, whole number ratio.*				
7. Write out the full balanced equation.				

Use the information from above to help you to complete the following tasks:

Masses to moles

1. Use number of moles = $\frac{\text{mass of chemical (g)}}{\text{molar mass}}$ to convert the following masses into moles:

   a) 15.2 g of iron(II) sulfate _____

   b) 6 g of ethane _____

   c) 87 g of magnesium hydroxide _____

   d) 0.63 g of nitric acid _____

   e) 0.098 g of sulfuric acid _____

2. Use mass (g) = molar mass × number of moles to:

   a) rearrange the formula to write an expression for the molar mass

   _____

   b) rearrange the formula to write an expression for the number of moles.

   _____

3. Use the following masses of reactants and products to write balanced symbol equations.

   a) 237.0 g of sodium thiosulfate ($Na_2S_2O_3$) reacts with 109.5 g of hydrochloric acid to produce 175.5 g of sodium chloride, 48.0 g of sulfur, 96.0 g of sulfur dioxide and 27.0 g of water.

   _____

   _____

b) 1.62 g of hydrogen bromide (HBr) reacts with 0.98 g of sulfuric acid to produce 0.36 g of water, 0.64 g of sulfur dioxide and 1.60 g of bromine (Br_2).

_____

_____

4. 21.30 g of chlorine reacts with 24.00 g of sodium hydroxide to produce 29.25 g of sodium chloride. There are two possible chemical reactions that can happen:

$$Cl_2 + 2NaOH \rightarrow NaClO + NaCl + H_2O$$

$$3Cl_2 + 6NaOH \rightarrow NaClO_3 + 5NaCl + 3H_2O$$

Which is the correct reaction?

_____

_____

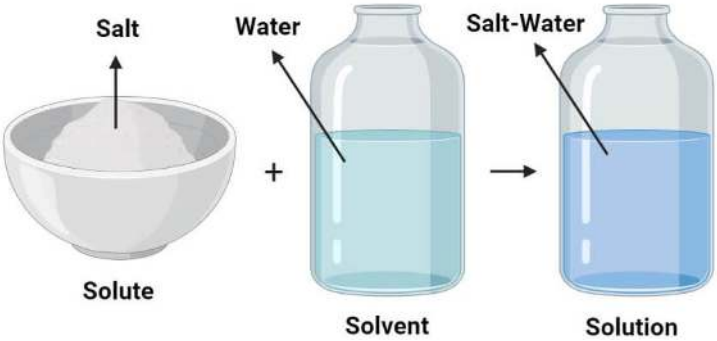
Lesson 9 – Concentration of solutions

Learning Objectives

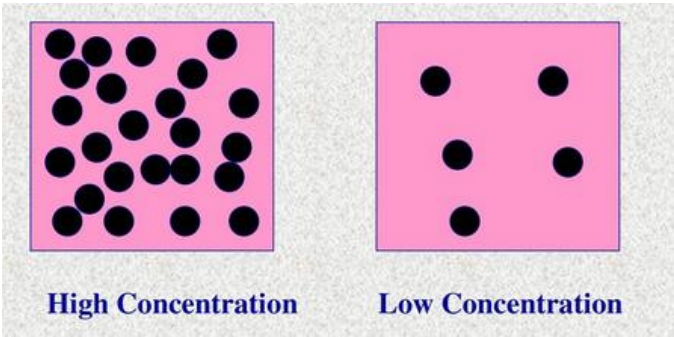
- Define concentration
- Calculate the mass of solute in a given volume of solution of known concentration
- (HT only) Explain how the mass of a solute and the volume of a solution is related to the concentration of the solution.

Explain and Understand

A **solution** forms when a **solute** dissolves in a **solvent**. For example salt in water forms a salt (saline) solution, salt is the **solute** and water is the **solvent**.



Concentration is the amount of substance in a given area. Another way to think of it is how crowded the solute particles are. If there are a lot of solute particles in the solution, we describe it as a **high concentration** or concentrated. If there are few solute particles in the solution we describe it as **low concentration** or dilute.

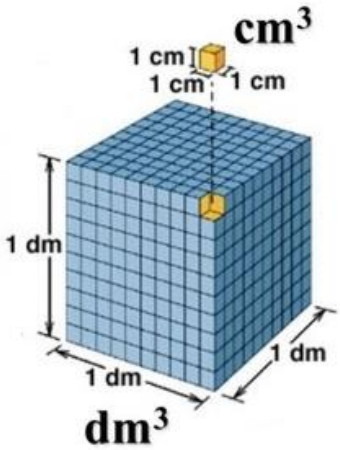


Concentration is measured as **mass** of solute per given **volume**. We use the units grams (g) per decimetre cubed (dm^3). For example a saline solution with the concentration of $50g/dm^3$ has a greater concentration than a solution of $1g/dm^3$.

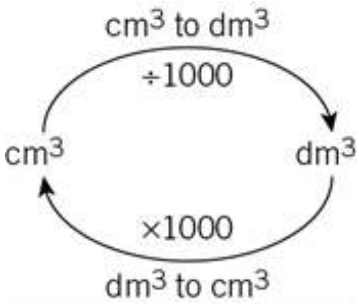
What is a cubic decimetre (dm^3)?

A cubic decimetre (dm^3) is actually the same volume as 1 litre. In science the SI unit for volume is the cubic decimetre so we have to use this and not litres.

Like the name suggests a decimetre (dm) is 1/10th of a meter. A centimetre (cm) is 1/100th of a metre. Therefore a cm^3 is a much smaller volume than a dm^3.



Sometimes you are given volumes in cm^3 you always must convert them into dm^3. To convert cm^3 into dm^3 you divide by 1000. For example 50cm^3 is equal to 0.05dm^3.



STOP here and practise:

Turn the following volumes into dm^3 (remember to show your workings):

1. 450 cm^3

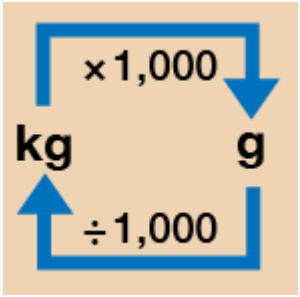
2. 5,000 cm^3

3. 989 cm^3

4. 3 cm^3

5. 100,000 cm^3

Another useful conversion is kilogram (g) to gram (g). To turn kilograms into grams x 1000.



Calculating concentration

To calculate the concentration you need the mass of solutes dissolved in grams (g) and the volume of solvent in cubic decimetres (dm^3).

Concentration (g/dm^3) = mass (g) ÷ volume (dm^3)

Worked example: Calculate the concentration of the solution formed when 10.0 g of sodium chloride is dissolved in 2.00 dm^3 of water.

10.0 g ÷ 2.00 dm^3 = **<u>5 g/dm^3</u>**

Sometimes you may have to convert the units to g and dm^3 before you can divide.

Worked example: Calculate the concentration of the solution formed when 4 g of copper sulfate is dissolved in 250 cm^3 of water

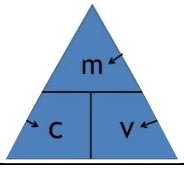
First convert the volume to cubic decimetres:

250 cm^3 ÷ 1000 = 0.25dm^3

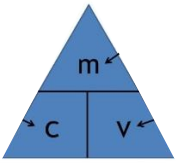
Then divide the mass by the volume:

4g ÷ 0.25dm^3 = **<u>16 g/dm^3</u>**

Calculating concentration – modelling. Complete the calculations.

	I do	We do	You do
	Calculate the concentration of the solution formed 10.0 g of sodium chloride is dissolved in 20cm^3 dm^3 of water.	Calculate the concentration of the solution formed when 4 g of copper sulfate is dissolved in 250 cm^3 of water	A student dissolved 38.5g of potassium manganate i n 38.5 cm^3 of water, what is the concentration?
1. Write out the concentration equation	C = m / v 		
2. Convert cm^3 into dm^3 by dividing by 1000	20 / 1000 = 0.02dm^3		
3. Substitute the values into the equation	10 x 0.02 =		
4. Write the answer with the units	0.2 g/dm^3		

Calculating mass – modelling. Complete the calculations.

	I do	We do	You do
	Calculate the mass of solute is dissolved in 25cm^3 to produce a 2 g /dm^3 concentrated solution?	Calculate the mass of solute is dissolved in 200cm^3 to produce a 5 g /dm^3 concentrated solution?	Calculate the mass of solute in 40cm^3 in a 15 g /dm^3 to produce a concentrated solution?
1. Write out the concentration equation	m = c X v 		
2. Convert cm^3 into dm^3 by dividing by 1000	25 / 1000 = 0.025dm^3		
3. Substitute the values into the equation	2 x 0.025 =		
4. Write the answer with the units	0.05 g		

STOP here and practise:

Calculate the concentrations of each of the following solutions in units of g/dm^3 (remember to always show your workings):

6. 2.5 g of glucose dissolved in 0.5 dm^3 of water

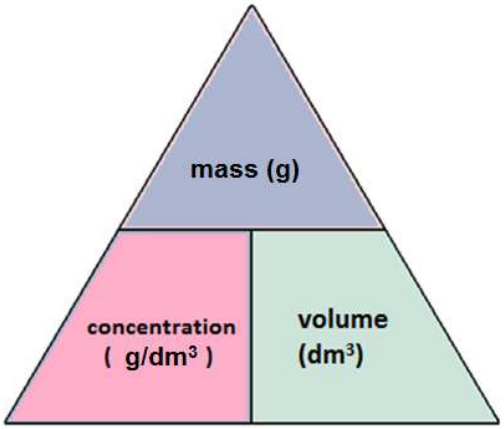
7. 26.25 g of potassium chloride dissolved in 1500 cm^3 of water

8. A student dissolved 38.5g of potassium manganate in 38.5 dm^3 of water, what is the concentration?

9. 7 kg of glucose dissolved in 1,400 cm^3 of water

10. In a factory 2kg of sugar was added to 10dm^3 of water. What is the concentration of the sugar solution produced?

Using a formula triangle is a simple way to remember how to calculate concentration and will help you if you have to calculate mass or volume from a concentration value. You need to learn this formula triangle.



11. 30 g of calcium chloride dissolved in 12.00 dm^3 of water. What is the concentration in g/dm^3

12. If I have 10g of sodium chloride and I want to make a solution of $5g/dm^3$ what volume of solvent should I use to dissolve the solute?

13. A student made a solution with the concentration of $50g/dm^3$ the volume of the solution was $0.75dm^3$. How much solute did the student use when making this solution?

14. What mass of copper carbonate should be added to $500cm^3$ to make a solution with the concentration of $20g/dm^3$?

15. A student made two solutions – solution **x** and solution **y**. Solution **x** contained 15 g of magnesium sulfate in 50 cm^3 of water. Solution **y** contains 100 g of magnesium sulfate in 100 cm^3 of water. Which solution has a higher concentration? Show all your workings.

16. A student was given 500cm^3 of a sodium sulfate solution which had the concentration of 5g/dm^3. The student removed 50cm^3 of this solution and placed it in an evaporating basin and heated until all the liquid had evaporated. What should be the mass of sodium sulfate in the evaporating basin?

Lesson 10 – Limiting reactants (HT only)

Learning Objectives

- Define the term limiting reactant.
- Be able to explain the effect of a limiting quantity of a reactant on the amount of products

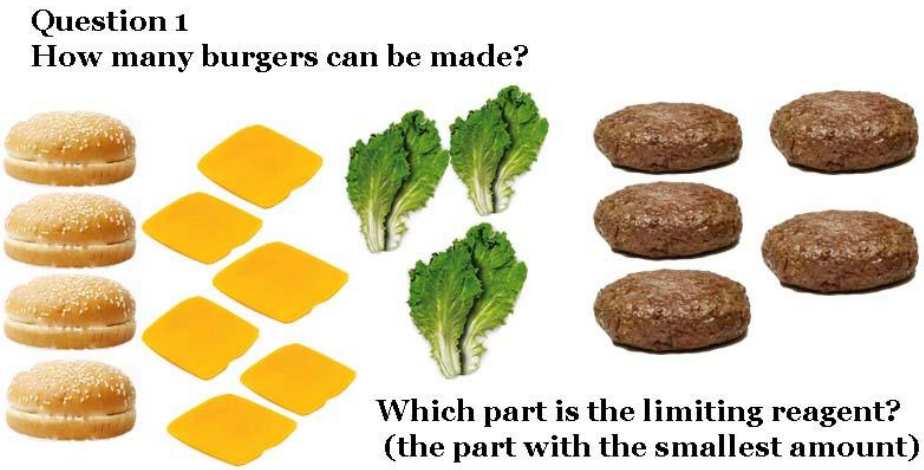
Explain and Understand

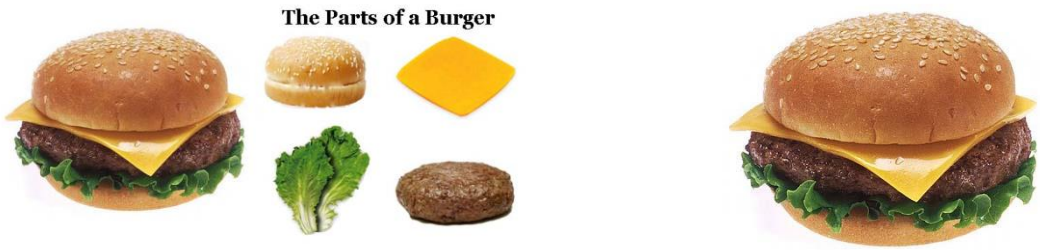
Limiting reactants

A reaction finishes when one of the reactants is all used up. The other reactant has nothing left to react with, so some of it is left over:

- the reactant that is all used up is called the limiting reactant
- the reactant that is left over is described as being in excess

The mass of product formed in a reaction depends upon the mass of the limiting reactant. This is because no more product can form when the limiting reactant is all used up. Let's use making burgers as an analogy.





Each burger requires one bap, a slice of cheese, a pair of lettuce leaves and a beef burger. There are four buns, five slices of cheese, three pairs of lettuce leaves and five meat burgers therefore only three burgers can be made. The lettuce is the limiting factor of the burger as there are only three lots of lettuce. This can be applied to chemical reactions.

Example

Which reactant is limited and which is in excess in the following example? **2.7 g of aluminium are reacted with 9.125 g of HCl.**

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

Convert mass into moles for both reactants:

$$moles\ of\ Al = \frac{mass\ (g)}{M_r} = \frac{2.7}{27} = 0.1$$

$$moles\ of\ HCl = \frac{mass\ (g)}{M_r} = \frac{9.125}{36.5} = 0.25$$

Now look at the ratio of the reactants:

2 mol Al : 6 mol HCl (1 mol Al : 3 mol HCl)

0.1 mol Al : 0.3 mole HCl (simplified)

There are only 0.25 moles of HCl (instead of 0.3 moles), so the HCl will run out first. It is the limiting reactant.

Now you can use the moles of the limiting reactant to calculate the mass of the product (Lesson 8 content). Remember to use the molar ratio between the limiting reactant and the product.

1. Given the following reaction:

$$C_3H_8 + 5O_2 \text{ -------> } 3CO_2 + 4H_2O$$

a) If you start with 14.8 g of C_3H_8 and 3.44 g of O_2, determine the limiting reagent

b) Determine the number of moles of carbon dioxide produced

c) Determine the number of grams of H_2O produced

d) Determine the number of grams of excess reagent left

2. Given the following equation:

$$Al_2(SO_3)_3 + 6\ NaOH \text{ ------> } 3\ Na_2SO_3 + 2\ Al(OH)_3$$

a) If 10.0 g of $Al_2(SO_3)_3$ is reacted with 10.0 g of NaOH, determine the limiting reagent

b) Determine the number of moles of $Al(OH)_3$ produced

c) Determine the number of grams of Na_2SO_3 produced

d) Determine the number of grams of excess reagent left over in the reaction

3. Given the following equation:

$$Al_2O_3 + Fe \text{ ------> } Fe_3O_4 + Al$$

a) If 25.4 g of Al_2O_3 is reacted with 10.2 g of Fe, determine the limiting reagent

b) Determine the number of moles of Al produced

c) Determine the number of grams of Fe_3O_4 produced

d) Determine the number of grams of excess reagent left over in the reaction

Practice Questions

Q1. Titanium is a transition metal.

(a) 40 kg of titanium chloride was added to 20 kg of sodium.

The equation for the reaction is:

$$TiCl_4 + 4\ Na \longrightarrow Ti + 4\ NaCl$$

Relative atomic masses (A_r): Na = 23 Cl = 35.5 Ti = 48
Explain why titanium chloride is the limiting reactant.
You **must** show your working.

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(4)

Q2. A scientist produces zinc iodide (ZnI_2). This is the method used:

1. Weigh 0.500 g of iodine.
2. Dissolve the iodine in ethanol.
3. Add an excess of zinc.
4. Stir the mixture until there is no further change.
5. Filter off the excess zinc.
6. Evaporate off the ethanol.

(a) Explain why the scientist adds excess zinc rather than excess iodine.

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(b) Calculate the minimum mass of zinc that needs to be added to 0.500 g of iodine so that the iodine fully reacts.
The equation for the reaction is:

$$Zn + I_2 \longrightarrow ZnI_2$$

Relative atomic masses (M_r): Zn = 65 I = 127

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Minimum mass of zinc = _____ g

(3)

Q3. The equation for the complete combustion of butane is:

$$2\ C_4H_{10} + 13\ O_2 \longrightarrow 8\ CO_2 + 10\ H_2O$$

14.5 g of butane was burned in 72.0 g of oxygen. Determine the limiting reactant.
You must include calculations in your answer.
Relative atomic masses (A_r): C = 12 H = 1 O = 16

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Q4. (a) Fluorine reacts with sulfur to produce sulfur hexafluoride (SF_6).

$$S + 3F_2 \rightarrow SF_6$$

Relative formula masses, M_r: F_2 = 38 SF_6 = 146
Calculate the mass of sulfur hexafluoride produced when 0.950 g of fluorine is reacted with an excess of sulfur. Give your answer to **3 significant figures**.

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