Chemistry Units C1, C2 &C3

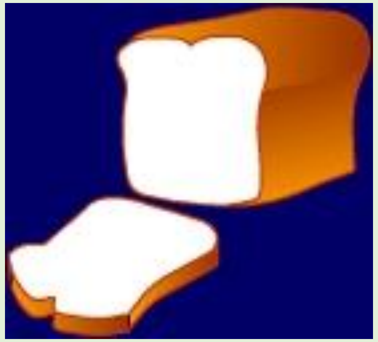
EXAM MASTERCLASS

Unit 1

The physical states of matter

At normal temperature almost all substances exist in one of three physical states:

liquids



solids

gases

Kinetic Theory

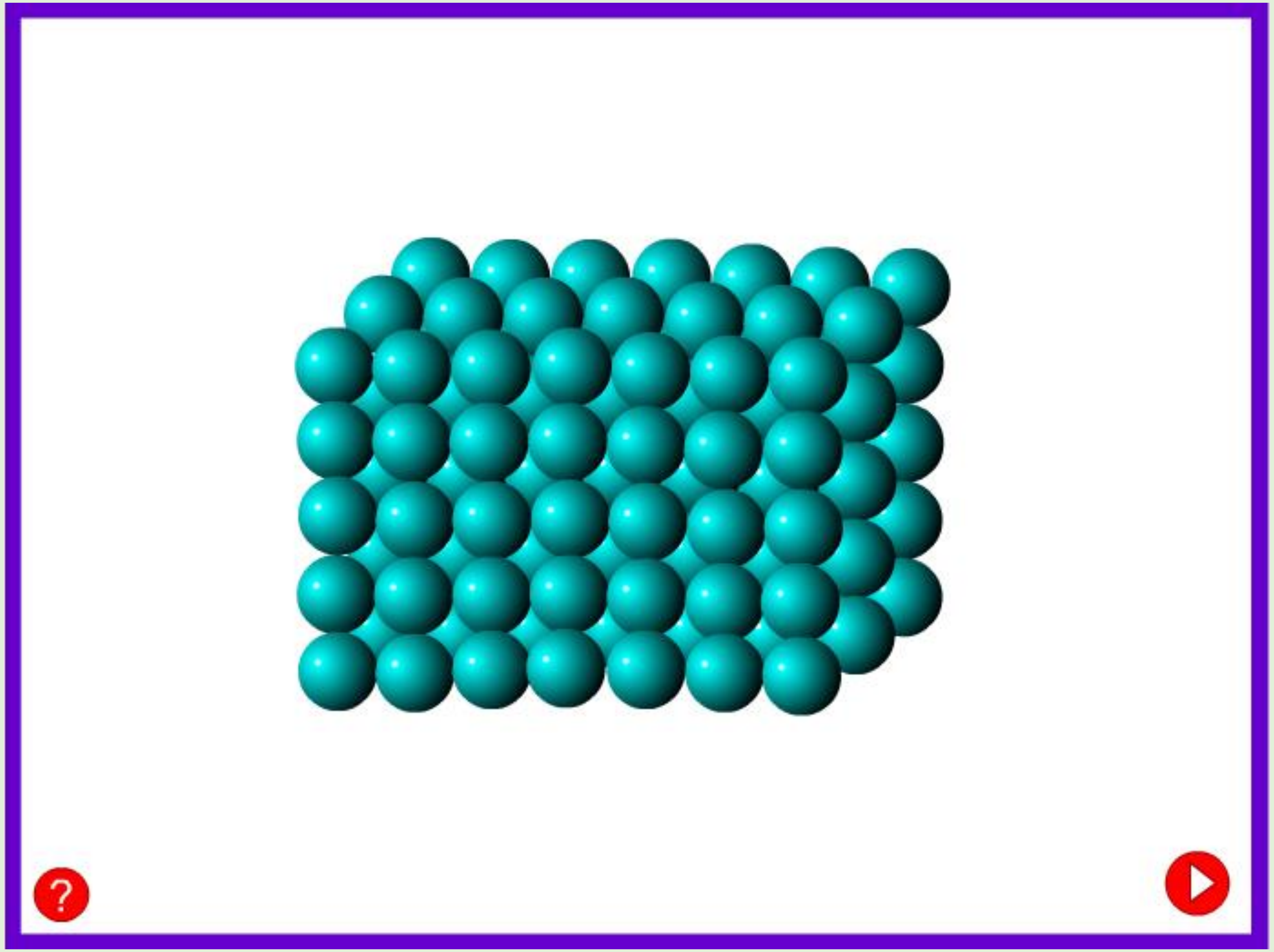
The existence of solids, liquids and gases is explained by the kinetic theory.

This is based on the following ideas:

1. All substances are made up of particles (atoms, ions and molecules).
2. That these particles are attracted to each other, some strongly and others weakly.
3. That these particles move around (i.e. have **kinetic** energy).
4. That their kinetic energy increases with temperature.

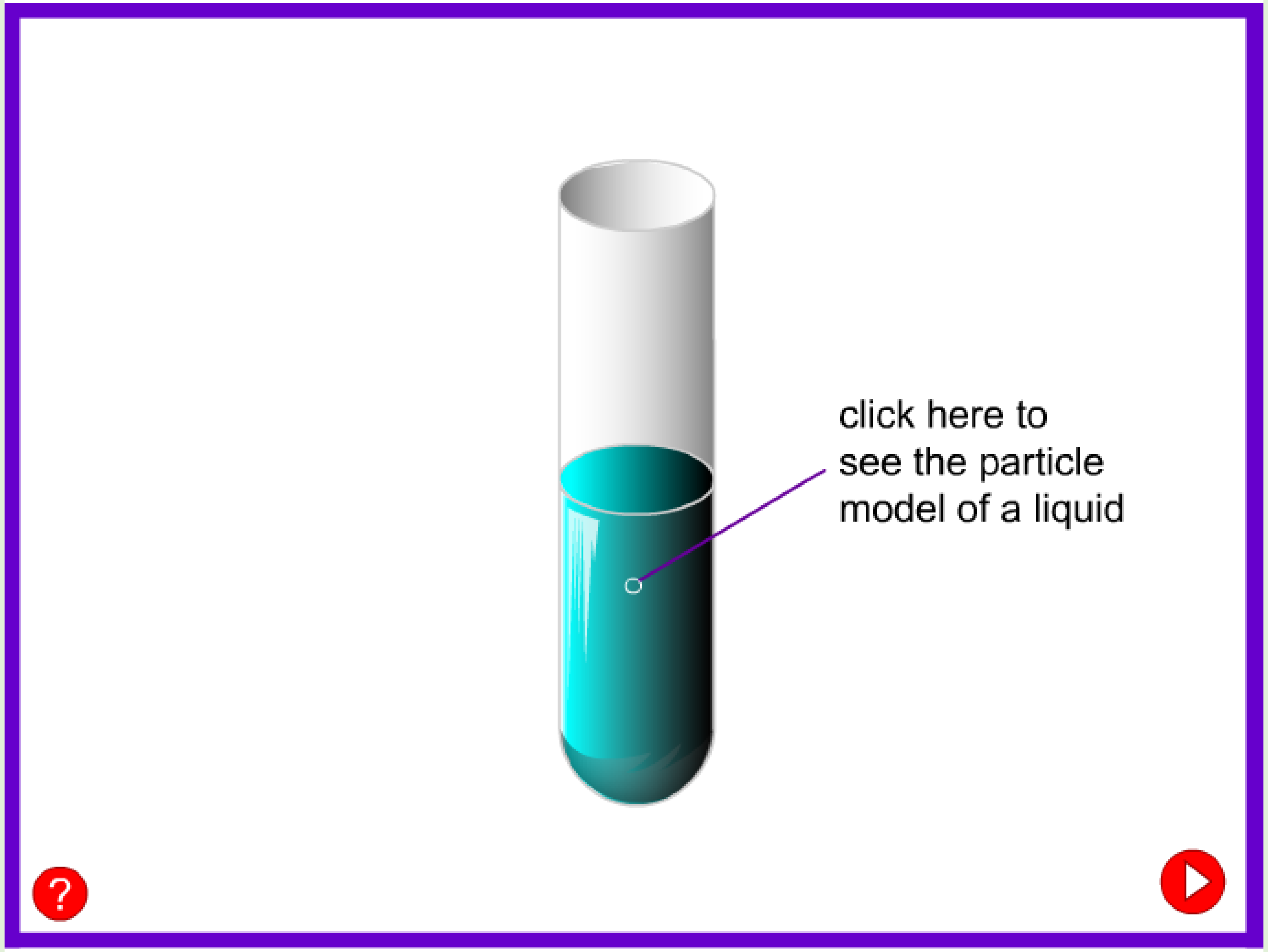
Particle arrangements: solids

This animation shows a 2-D view of the motion of the atoms in a 3-D solid.

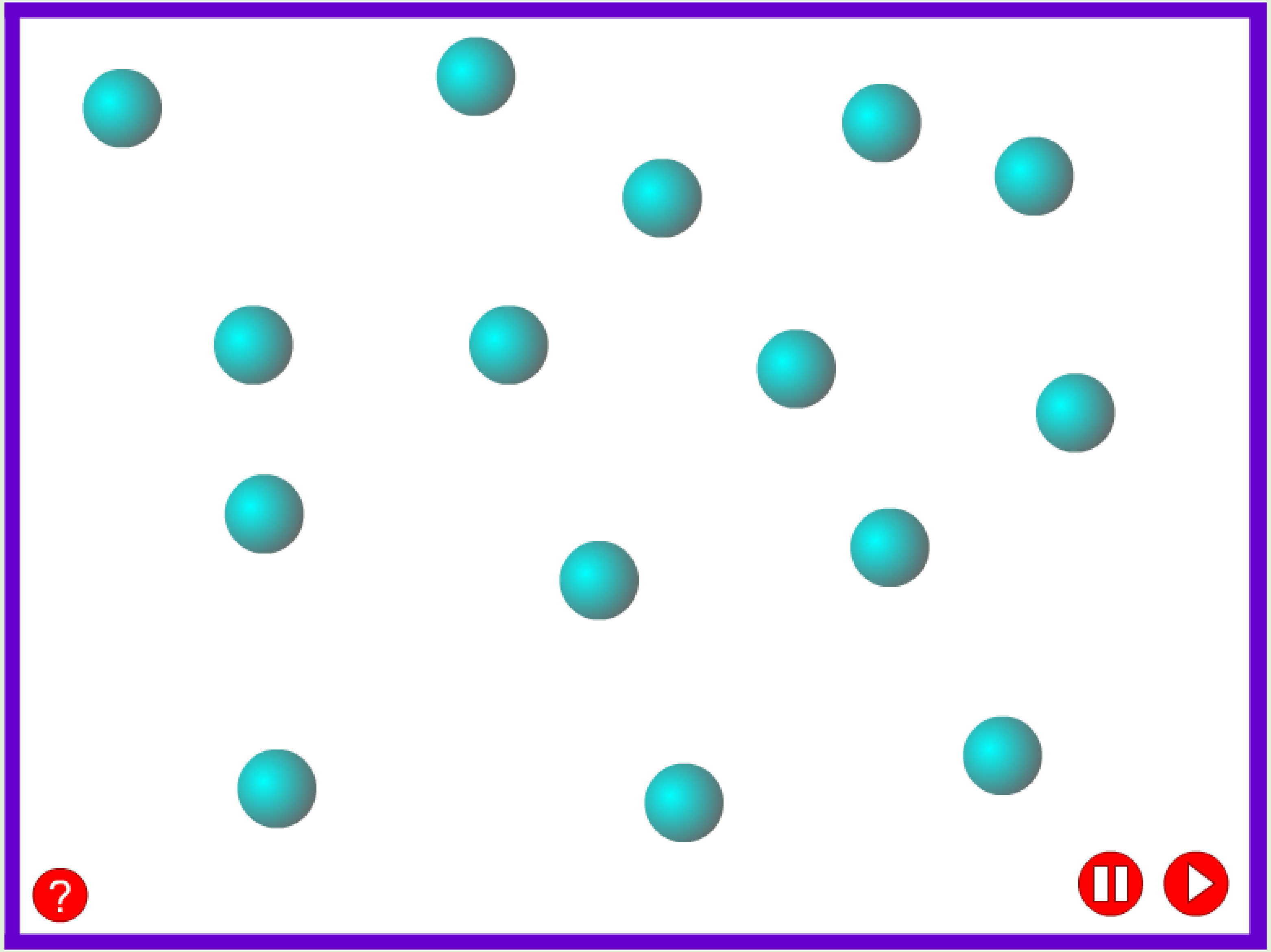


Particle arrangements: liquids

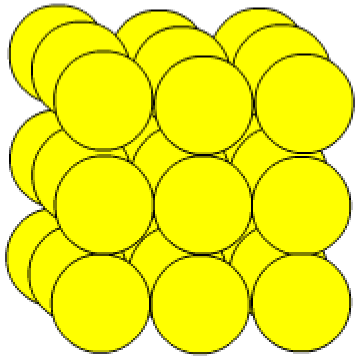
This animation shows a 2-D view of the motion of the atoms in a liquid. There is no order.

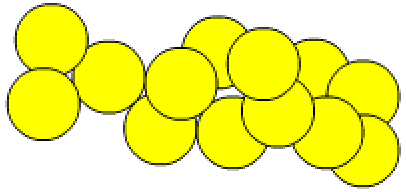


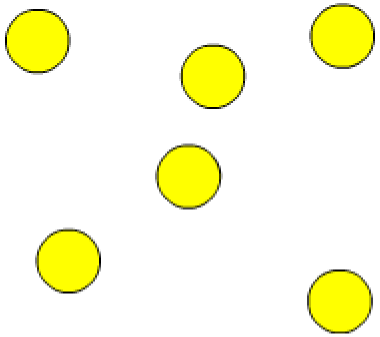
Particle arrangements: gases



Properties - Density







Solid

Particles close together.
High density.

Liquid

Particles close together.
High density.

Gas

Particles well spread out.
Low density.

Below are some properties of solids, liquids and gases. Drag them to the right place in the table

Flow

Cannot be compressed

Can be compressed

Do not flow

Takes shape of a container bottom

Shape stays the same

Take shape of the whole container

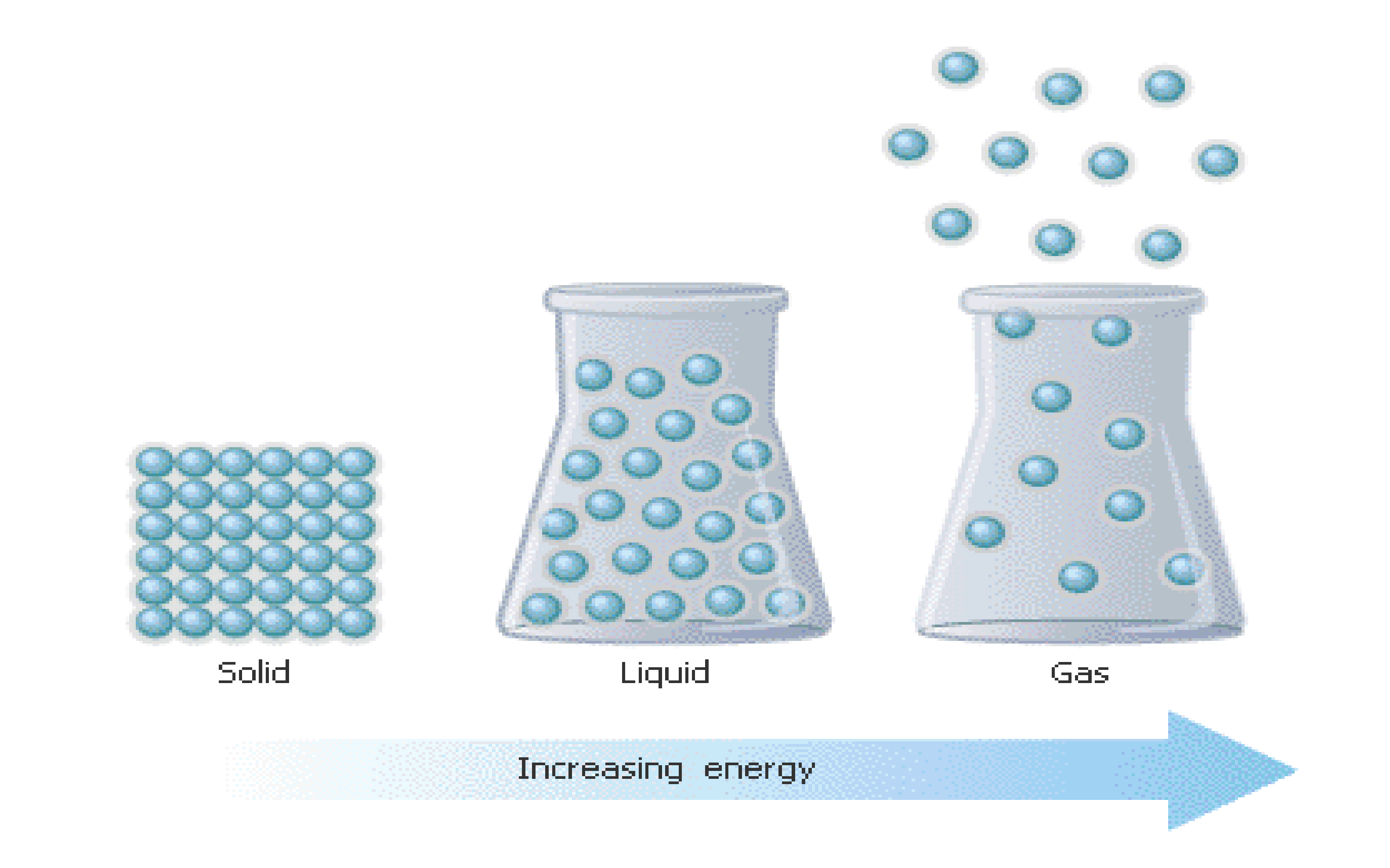
Fixed volume

No definite volume

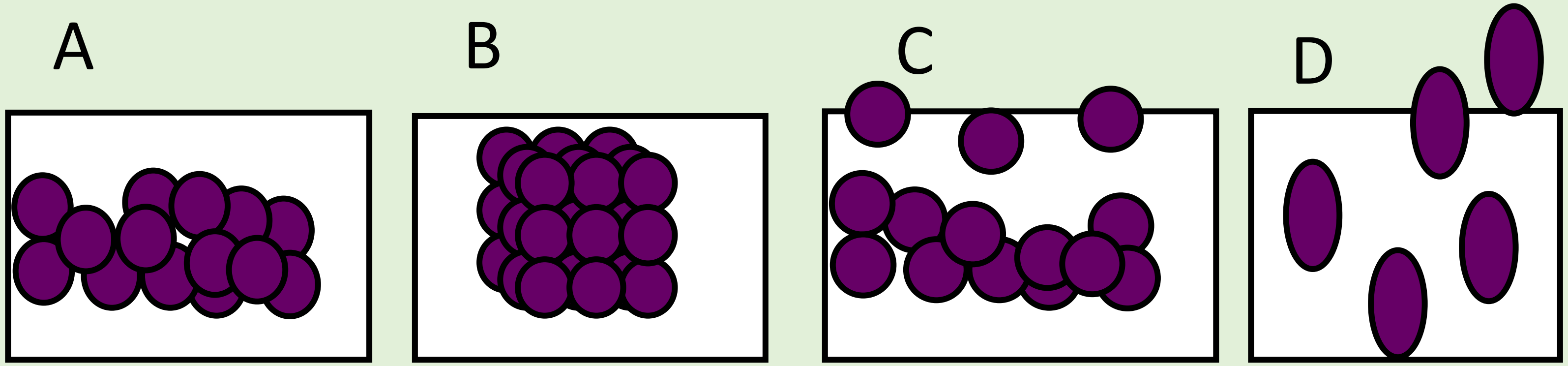
Solid	Liquid	Gas

?

Changing States of Water -The Particles



1. Which of the diagrams represents a liquid evaporating?



Particle Model Limitations (HT only)

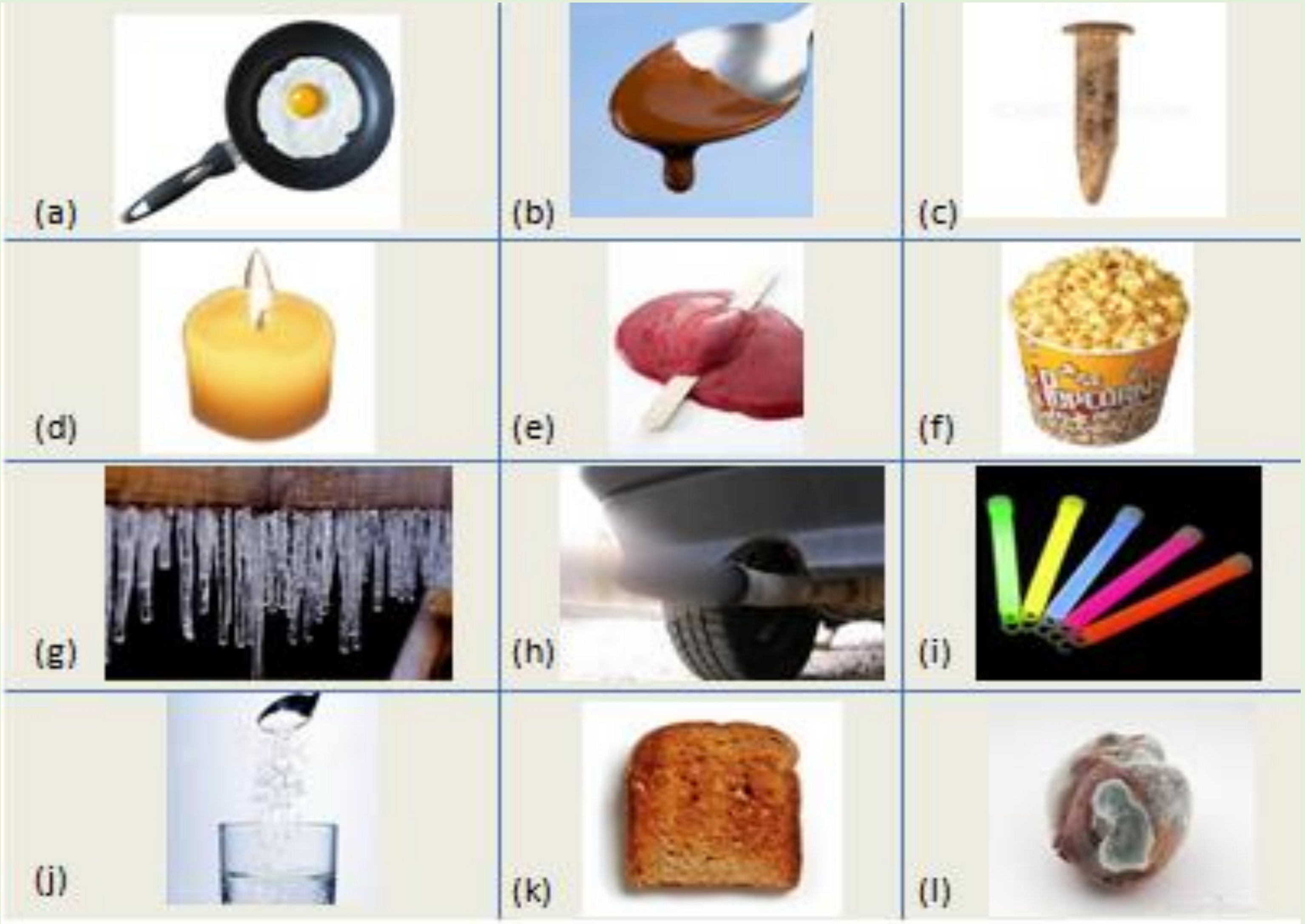
- **The model deals with particles that are inelastic spheres (e.g. like bowling balls)**
- **That doesn't take into account the forces of attraction between particles, the size of the particles and the space between them**

Simple Chemical reactions

What is a chemical reaction?

What is a physical change?

Are these chemical or physical changes?



- A chemical reaction is the formation of a new substance from two or more other substances.
- Eg – cooking an egg, burning coal or lighting a firework.
- A physical change is something that although changing states still is the same chemically.
- Eg – ice melting to water and boiling into steam.

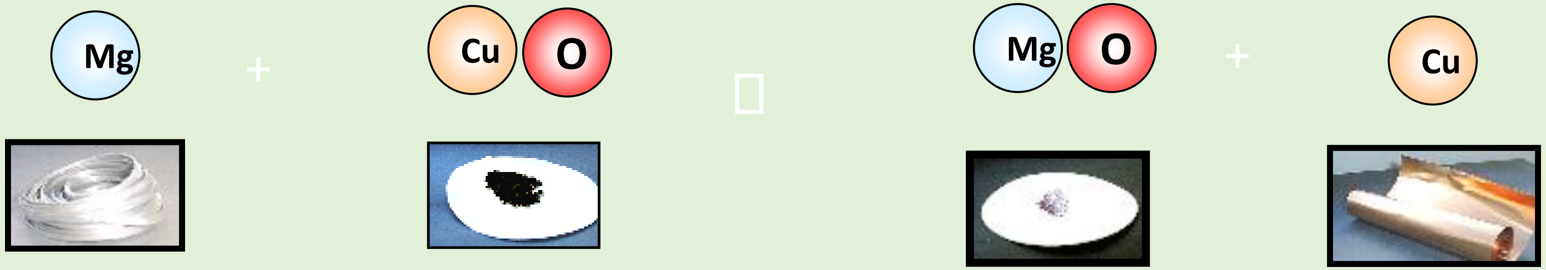
- Chemical equations are good indicators if a chemical or physical change has taken place

Reactants and Products

In chemical reactions new substances are formed.

We call the starting materials ***reactants*** and the substances that are formed: ***products***.

For example -



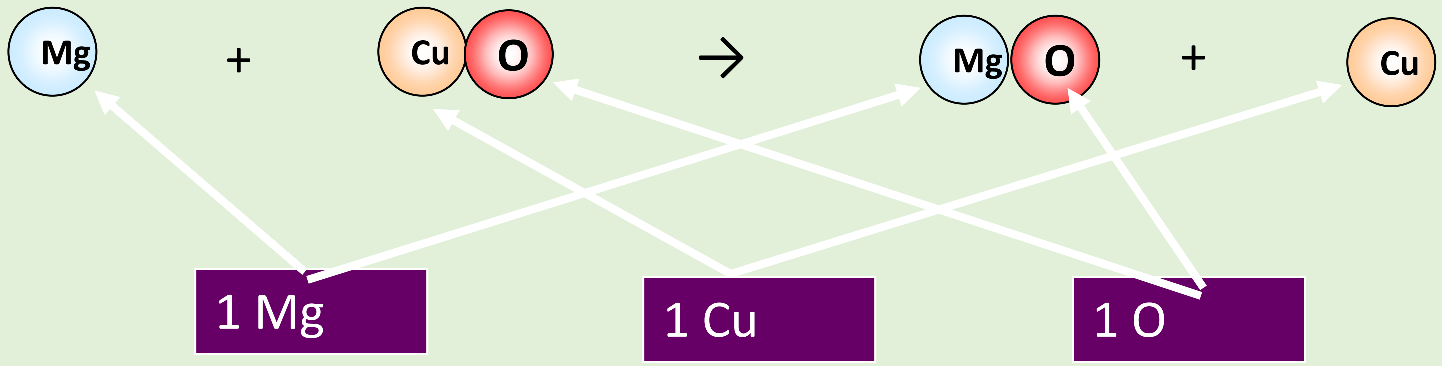
Substance	Reactant or Product?
magnesium oxide	product
magnesium	reactant
copper oxide	reactant
copper	product

No new atoms

Note that chemical reactions can produce very different looking substances.

This is because atoms have bonded (joined) together in new ways.

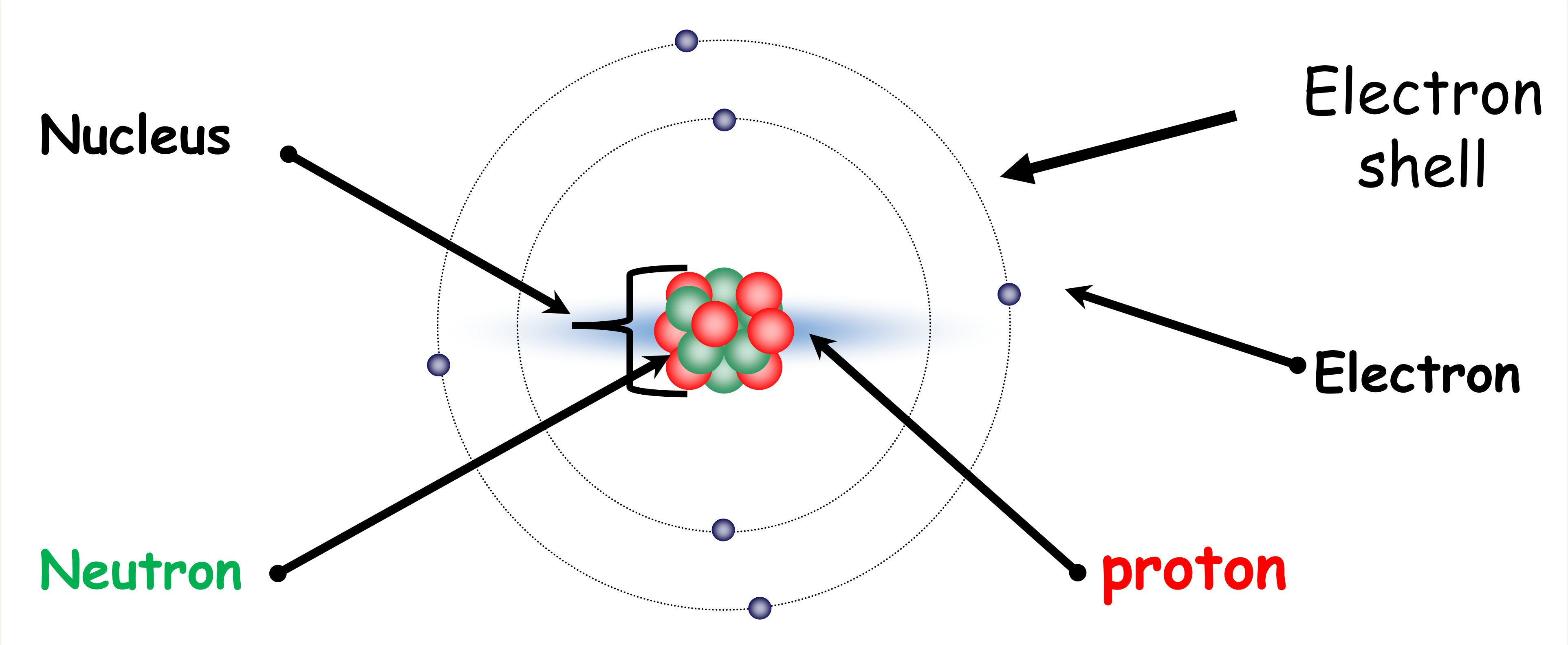
It is *not* because any new atoms have been formed.



The same number of each type of atom are present before and after a chemical reaction- **this is called THE CONSERVATION OF MASS**

Developing the Atomic Model

But how did scientists discover the structure of this?



Keywords: atom, atomic model, electrons, spheres, plum pudding, Rutherford, Marsden, alpha scattering, nuclear model, Bohr, Chadwick.

New experimental evidence may lead to a scientific model being changed or replaced.

Keywords: atom, atomic model, electrons, spheres, plum pudding, Rutherford, Marsden, alpha scattering, nuclear model, Bohr, Chadwick.

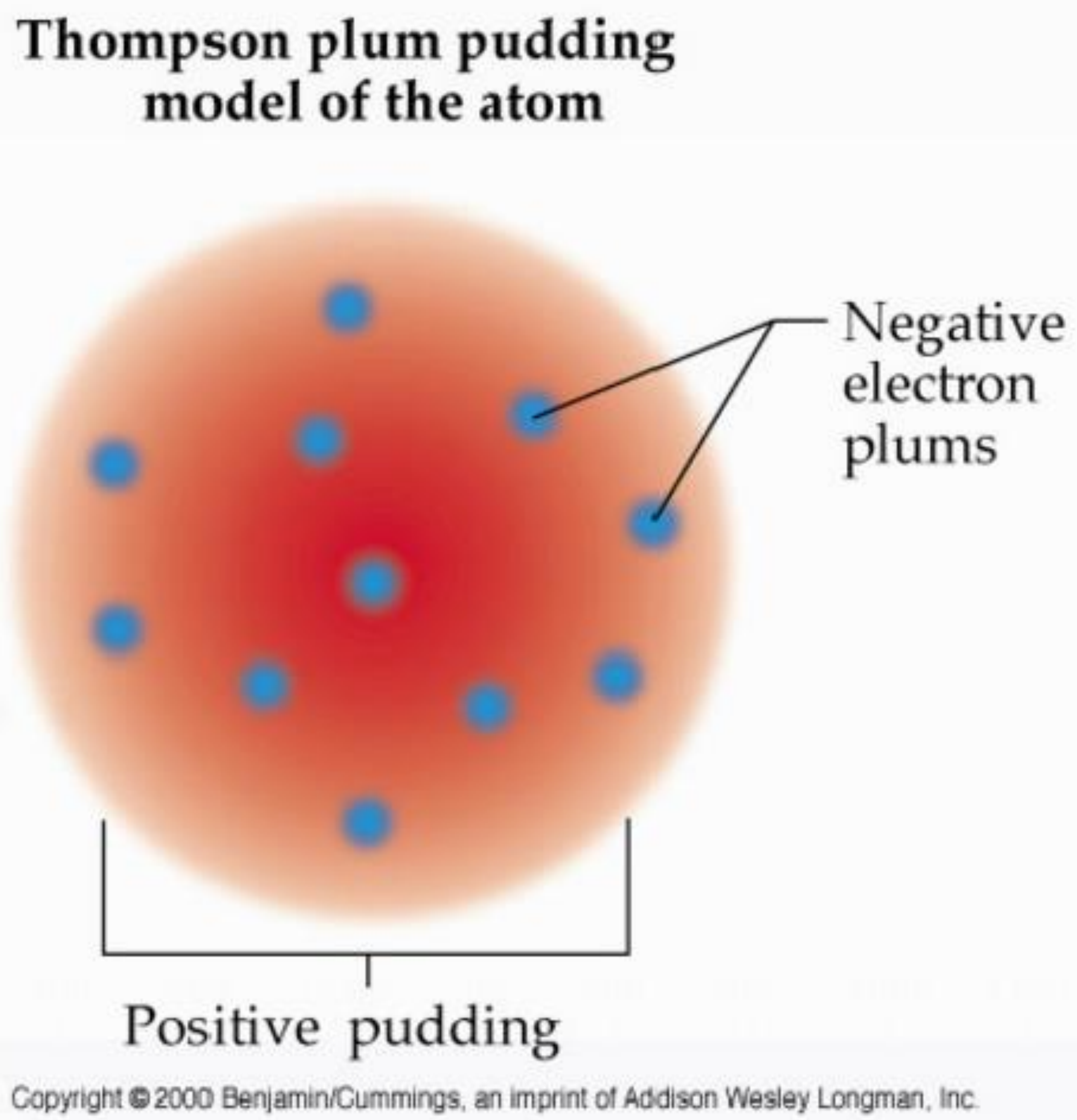
John Dalton

1 Jan 1803

John Dalton was an English chemist that created the Atomic Theory of Matter, a composition of previous findings by Democritus and his own findings. He included in this theory that all matter is made of atoms, that atoms cannot be created nor destroyed and also, atoms of different elements combine in whole ratios to form chemical compunds. His theory would later contribute to an advance in the atomic model.

J.J. Thomson

1 Jan 1897



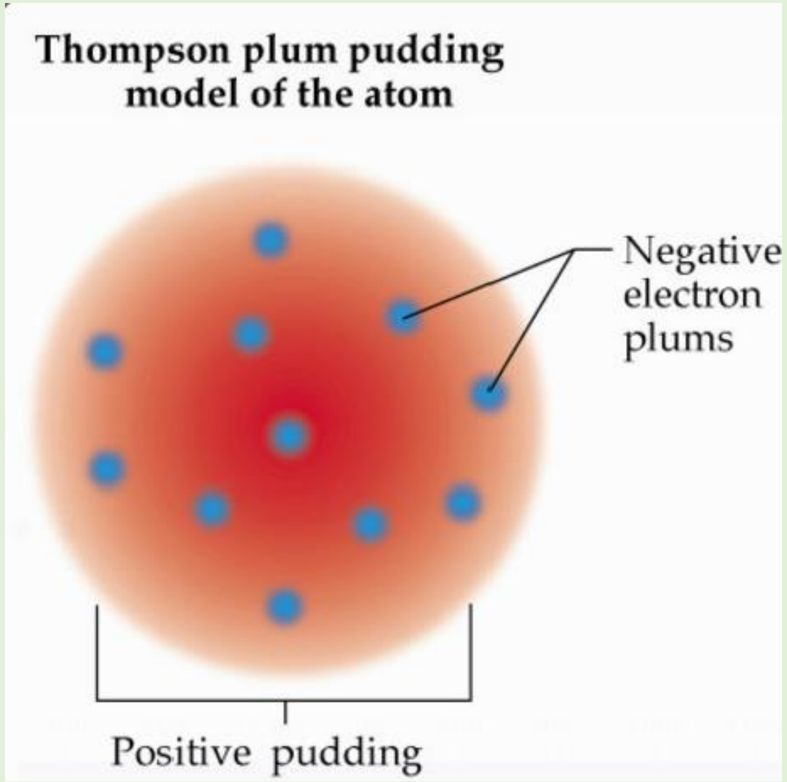
J.J. Thomson was a very important scientist when it came to the atomic model. Up until his time, all models of the atom looked like a big solild ball. J.J. Thomson discovered the electron, which led him to create the "plum pudding" atomic model. In this model, he thought that the atom was mostly positive, and negative electrons wandered around the atom. The "plum pudding" model influenced other scientists to make better atomic models.

Plum-pudding model in more detail

The earliest model of the atom.

In this model, the atom was imagined to be a sphere of positive charge with negatively charged electrons dotted around inside it - *like plums in a pudding.*

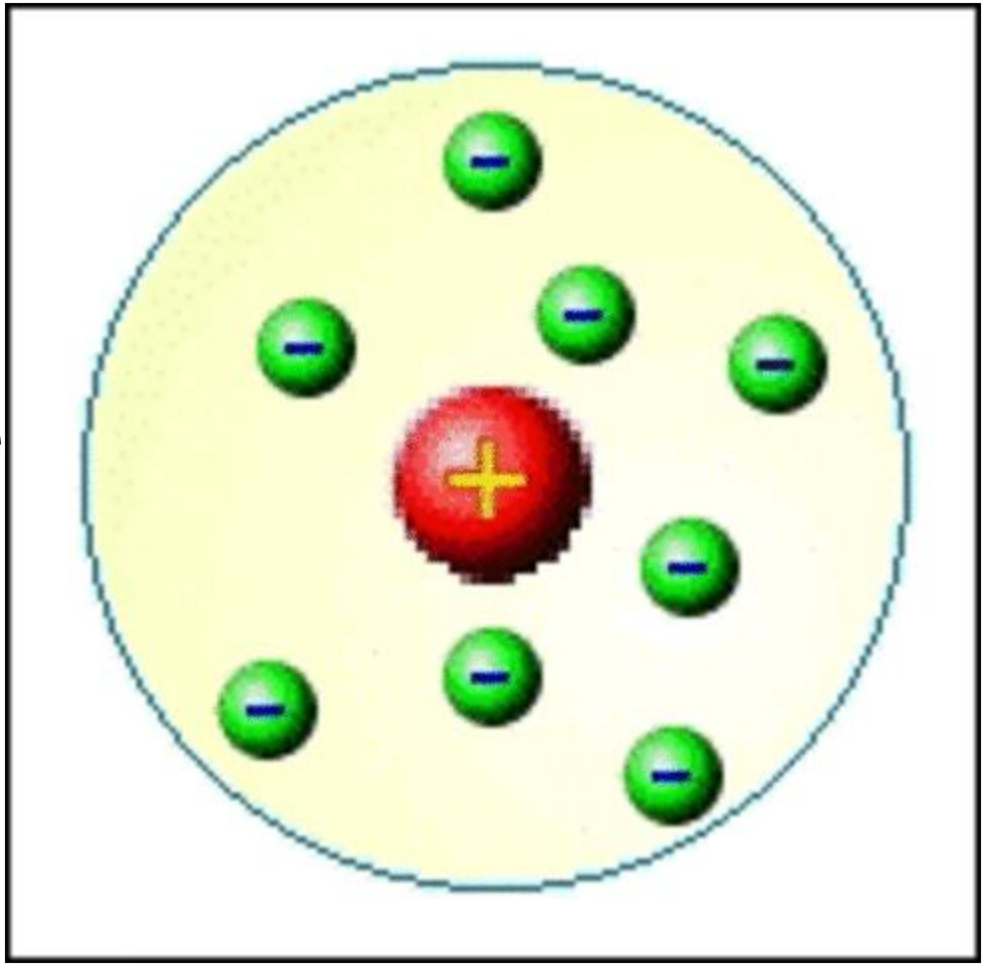
An experiment carried out in 1905 showed that the plum pudding model could not be correct.

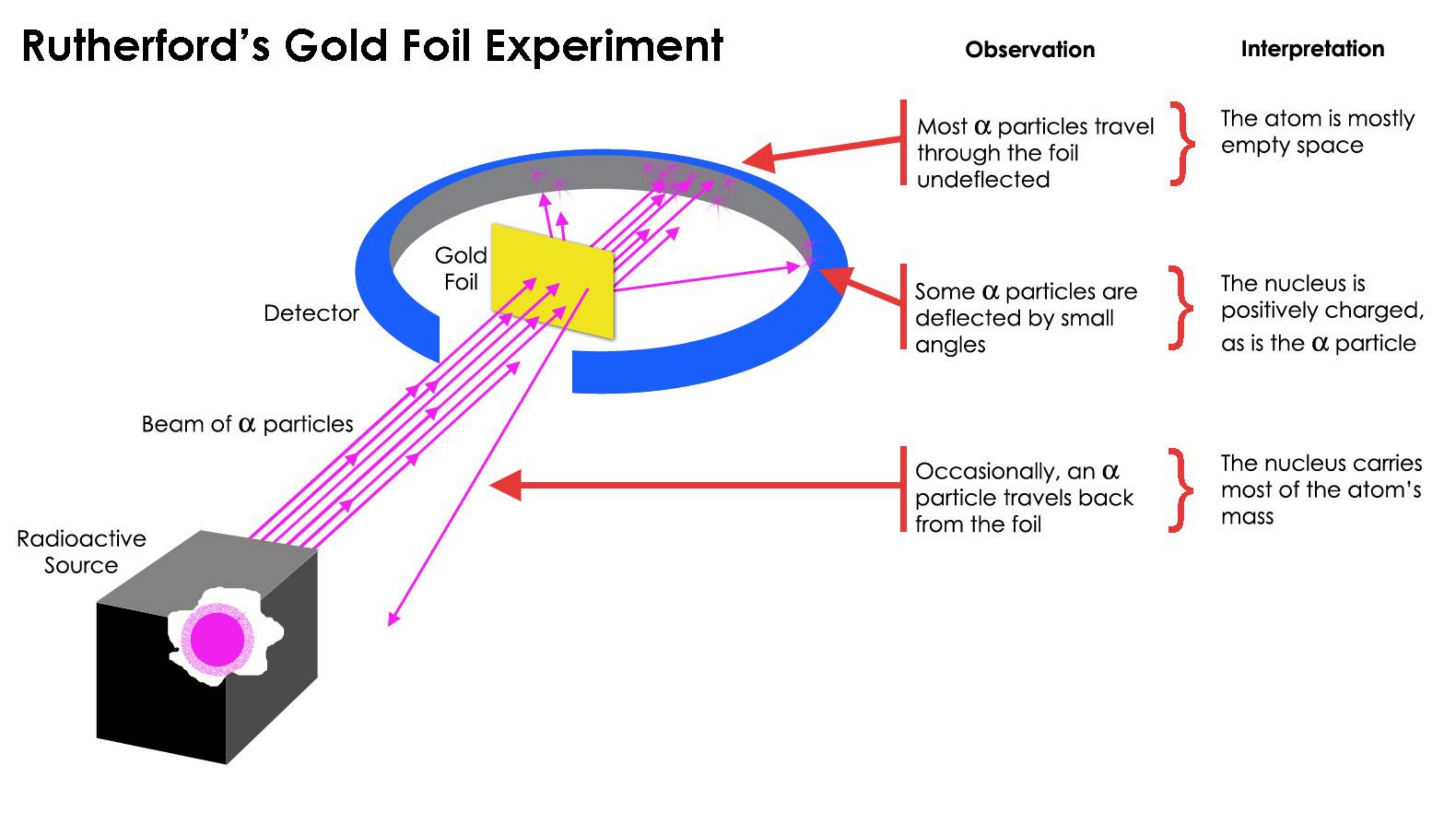


Keywords: atom, atomic model, electrons, spheres, plum pudding, Rutherford, Marsden, alpha scattering, nuclear model, Bohr, Chadwick.

Ernest Rutherford 1909

Ernest Rutherford was another scientist that changed the atomic model. He felt that J.J. Thomson's model was incorrect, so he created a new one. He created the nucleus, and said that instead of the positive matter being the whole atom, it was just in the middle. He said the atom was mostly empty space and that the electrons surrounded the positive nucleus. This model influenced one of his own students to perfect the atomic model later on.





Keywords: atom, atomic model, electrons, spheres, plum pudding, Rutherford, Marsden, alpha scattering, nuclear model, Bohr, Chadwick.

Rutherford's experiment

Rutherford designed an experiment to test the plum pudding model.

- It was carried out by his assistants Hans Geiger and Ernest Marsden.

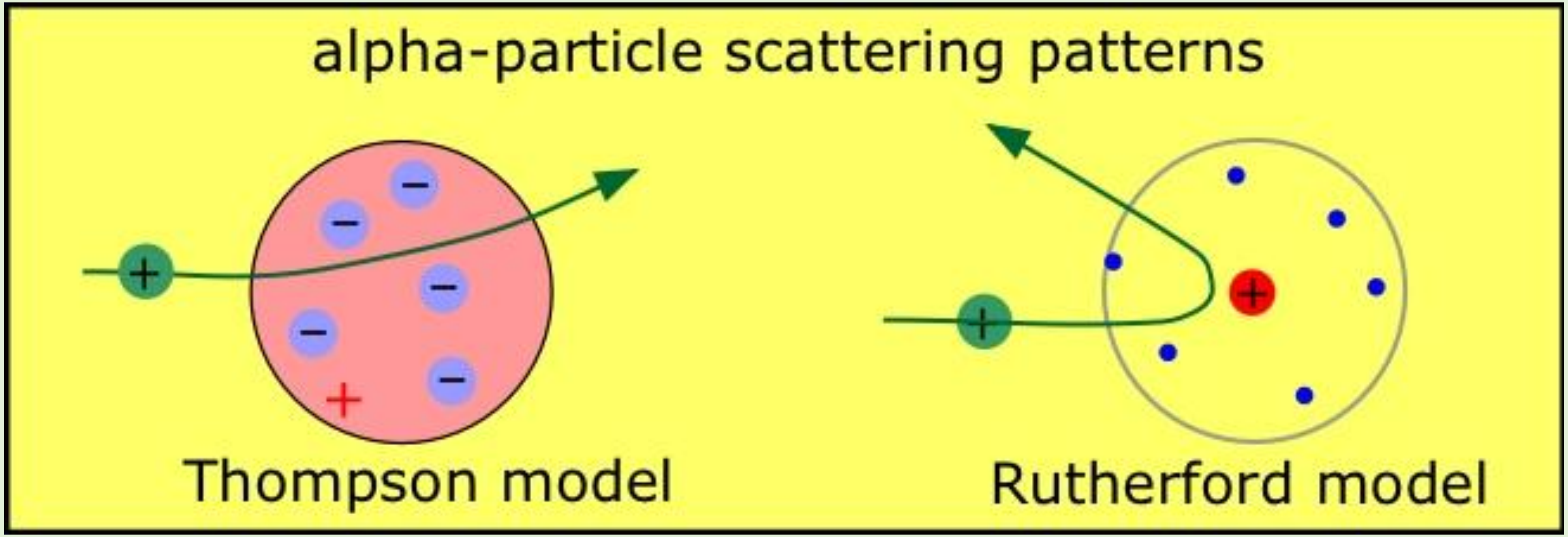
A beam of *alpha particles* was aimed at very *thin gold foil* and their passage through the foil detected.

The scientists expected the alpha particles to pass straight through the foil, but this did not happen.

The scientists realised that the positively charged alpha particles were being *repelled* and *deflected* by a tiny concentration of positive charge in the atom.

As a result of this experiment, the plum pudding model was replaced by the nuclear model of the atom.

Keywords: atom, atomic model, electrons, spheres, plum pudding, Rutherford, Marsden, alpha scattering, nuclear model, Bohr, Chadwick.



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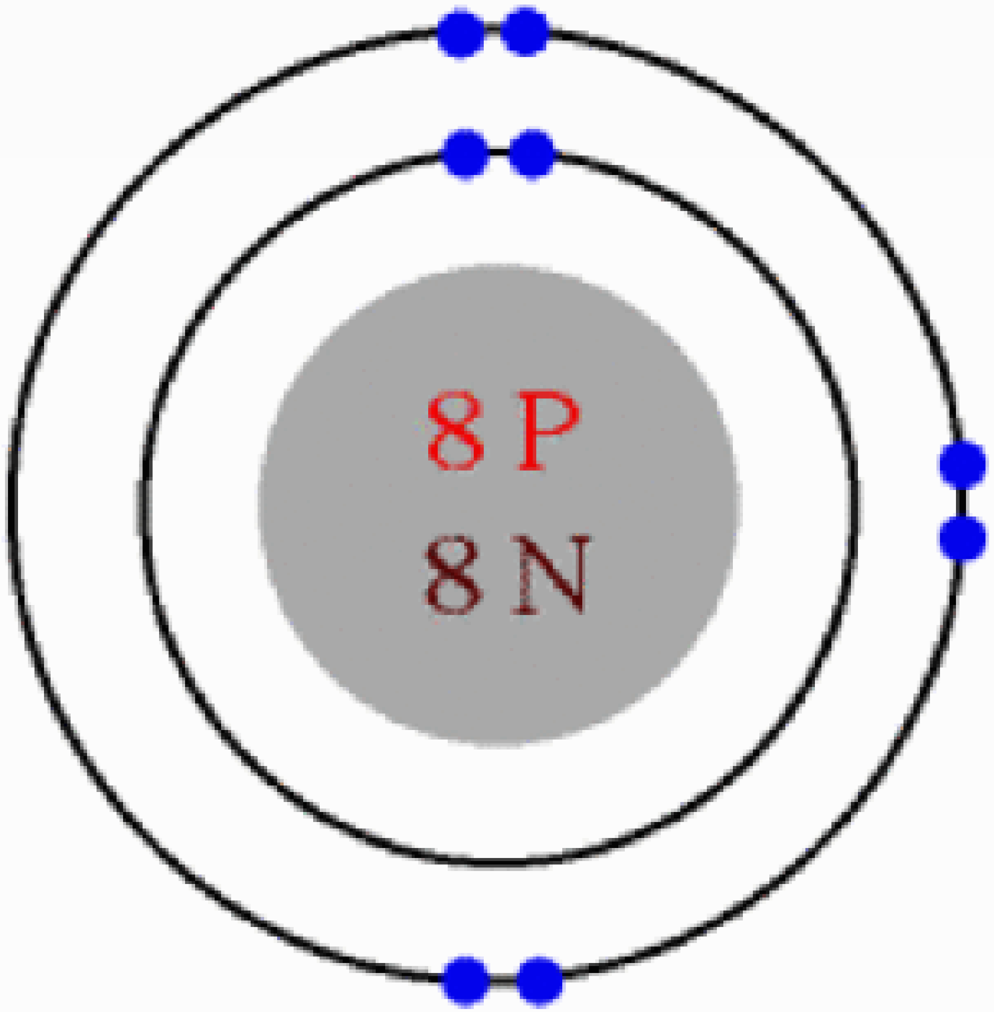


Niels Bohr

1 Jan 1913

Niels Bohr was a Danish scientist that was a student of Rutherford. He decided to make a new model based off of Rutherford's model, but changed the orbit of the electron. Also, he created energy levels in the atom, where only a certain amount of electrons could fit on one energy level of the atom. Bohr also used Planck's ideas in order to create quantum mechanics, his new concept regarding energy. This model is still used to this day.

Unit 2

Relative Formula Mass

Relative Atomic Mass (Ar)

- The atoms of each element have a different mass.
- **Carbon is given a relative atomic mass of 12.**
- The RAM of other atoms compares them with carbon.

Element	Symbol	Times as heavy as carbon	R.A.M
Helium	He	one third	**4**
Beryllium	Be	three quarters	**12**
Molybdenum	Mo	Eight	**96**
Krypton	Kr	Seven	**84**
Oxygen	O	One and one third	**16**
Silver	Ag	Nine	**108**
Calcium	Ca	Three and one third	**40**

Relative Formula Mass (Mr)

- To calculate formula mass we add together the atomic masses of all the atoms shown in the formula. (N=14; H=1; Na=23; O=16; Mg=24; Ca=40)

Substance	Formula	Formula Mass
Ammonia	$\mathbf{NH_3}$	14 + (3x1)=17
Sodium oxide	$\mathbf{Na_2O}$	(2x23) + 16 =62
Magnesium hydroxide	$\mathbf{Mg(OH)_2}$	24+ 2(16+1)=58
Calcium nitrate	$\mathbf{Ca(NO_3)_2}$	40+ 2(14+(3x16))=164

Definitions to Learn:

- **Relative Atomic mass (Ar) is the mean mass of an element compared to 1/12 of the mass of a C^{12} atom.**

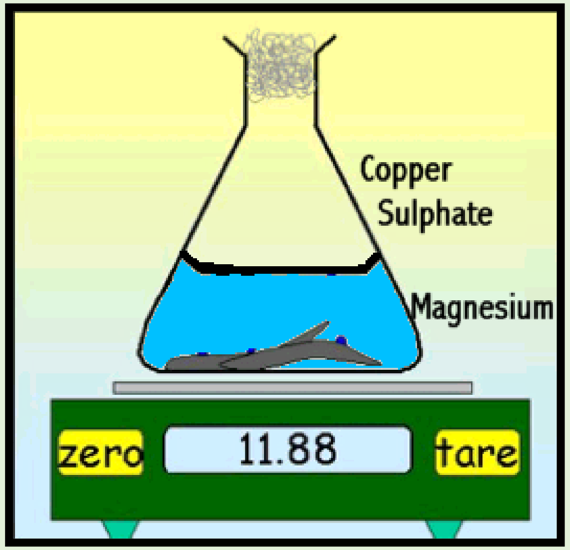
- **Relative Formula Mass (Mr) is the mean mass of a unit of a substance compared to 1/12 of the mass of a C^{12} atom.**

- **Relative Molecular mass refers to the Relative Formula mass of molecules**

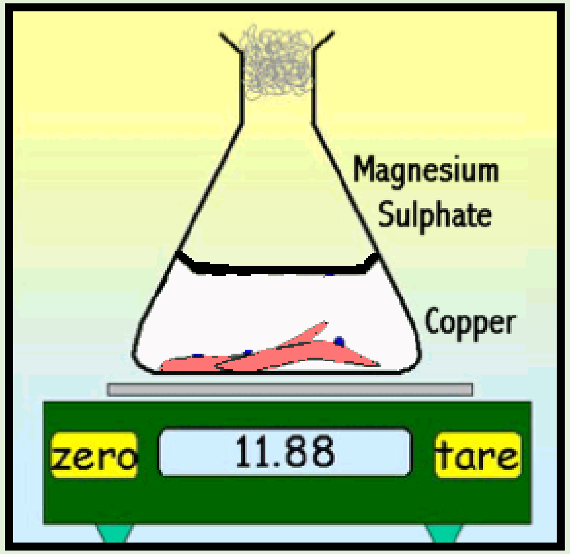
Reacting Masses

Conservation of Mass

- New substances are made during chemical reactions
- The same atoms are present before and after reaction. They have just joined up in different ways.
- Because of this **the total mass of reactants is always equal to the total mass of products**.
- This idea is known as the **Law of Conservation of Mass.**

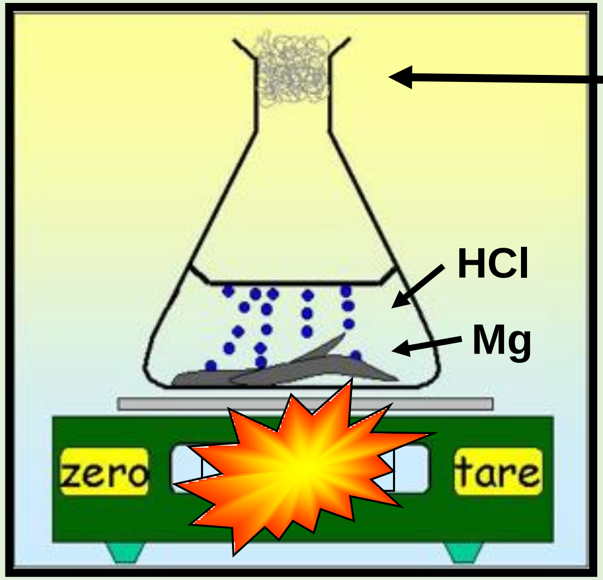


Reaction but no mass change



Conservation of Mass

- There are examples where the mass may ***seem*** to change during a reaction.
- **Eg. where a gas is given off** the mass of the chemicals in the flask will decrease because gas atoms will leave the flask. If we carry the same reaction in a strong sealed container the mass is unchanged.

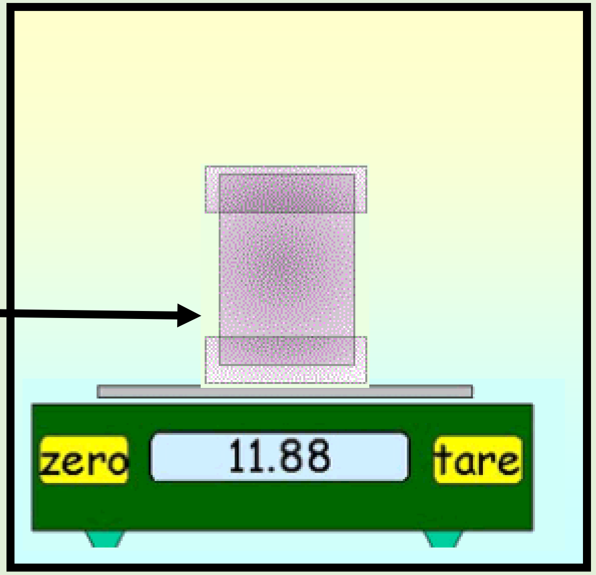


Gas given off.

Mass of chemicals in flask decreases

Same reaction in sealed container:

No change in mass



Reacting Mass and Equations

Atomic masses: C=12; O=16

- By using the formula masses in grams we can deduce what masses of reactants to use and what mass of products will be formed.

carbon + oxygen → carbon dioxide

$$C + O_2 \rightarrow CO_2$$

$$12 + 2 \times 16 \rightarrow 12+(2\times16)$$

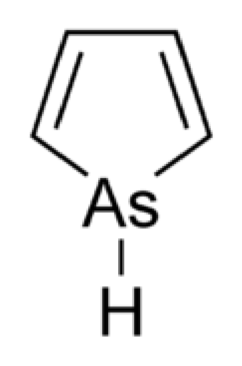
12g 32g 44g

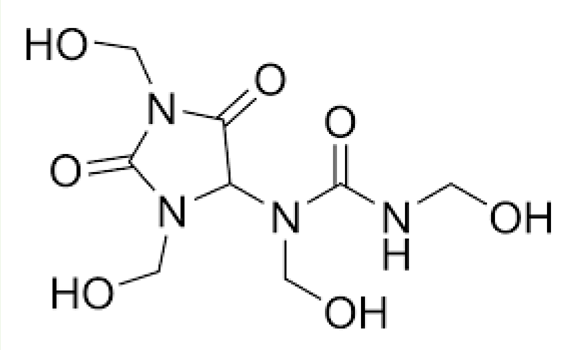
So we need 32g of oxygen to react with 12g of carbon and 44g of carbon dioxide is formed in the reaction.

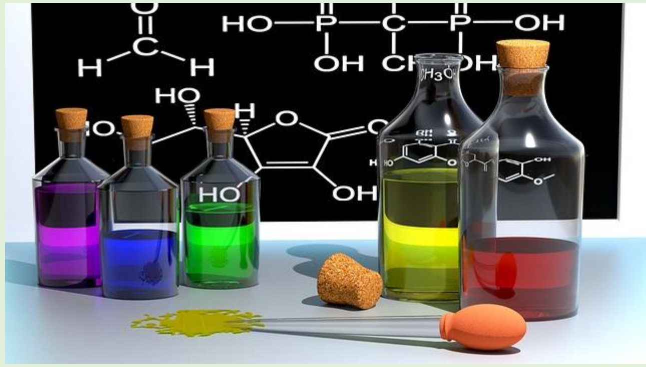
Note: total mass of the reactants = the total mass of the products

04/01/2022

Empirical formula

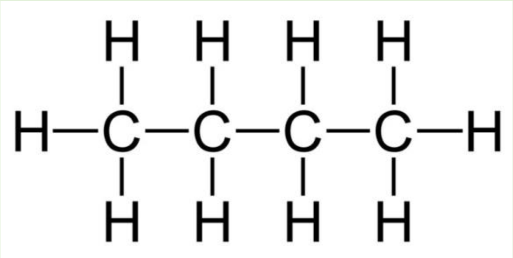






Molecular formula

- Tells us the number of elements in a compound.
- Butane for example



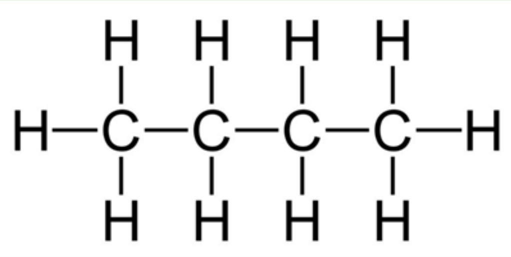
- Has the molecular formula, C_4H_{10}

Empirical formula

- Is the simplest whole number ratio of atoms in a compound.
- SIMPLEST
- WHOLE NUMBER
- RATIO

For example

- What is the ratio of carbon to hydrogen in butane?

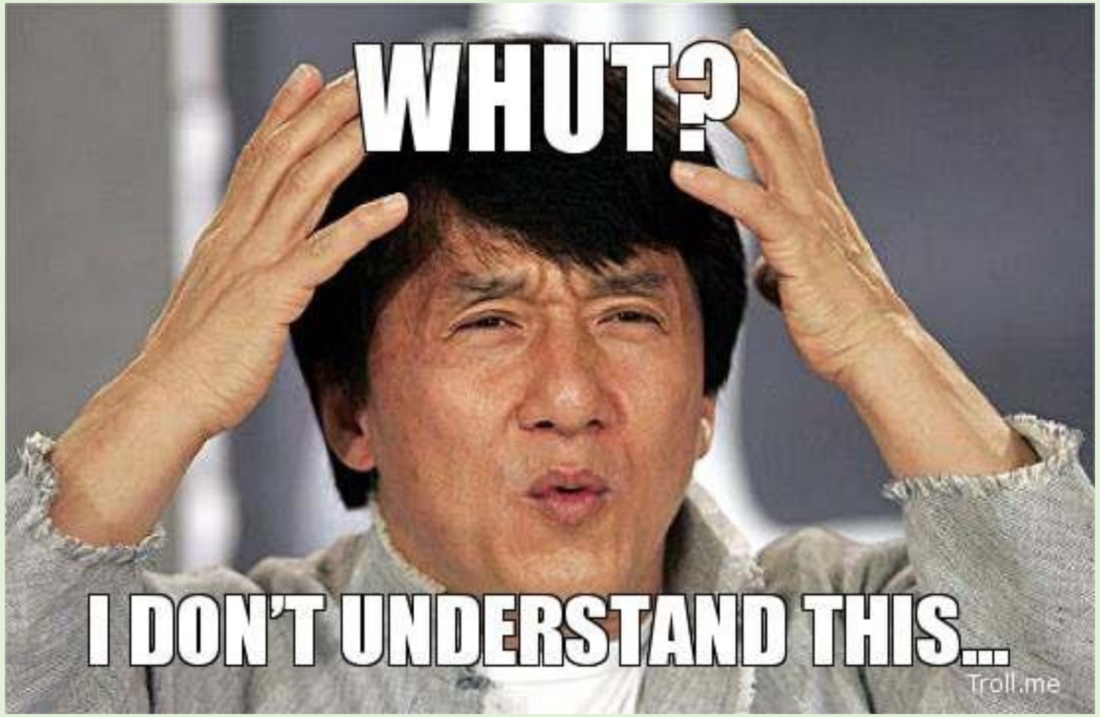


- **C_4H_{10}**

C_2H_5

Worked example

- A sample of an oxide of nitrogen is found to contain **30.4% nitrogen** and **69.6% oxygen**. What is the empirical formula?



Nitrogen	Oxygen
Mass of nitrogen = 14 % of nitrogen = 30.4%	Mass of oxygen = 16 % of oxygen = 69.6%
30.4 / 14 = 2.171	69.6 / 16 = 4.35
2.171 / 2.171 = 1	4.35 / 2.171 2

NO_2

Pure Substances

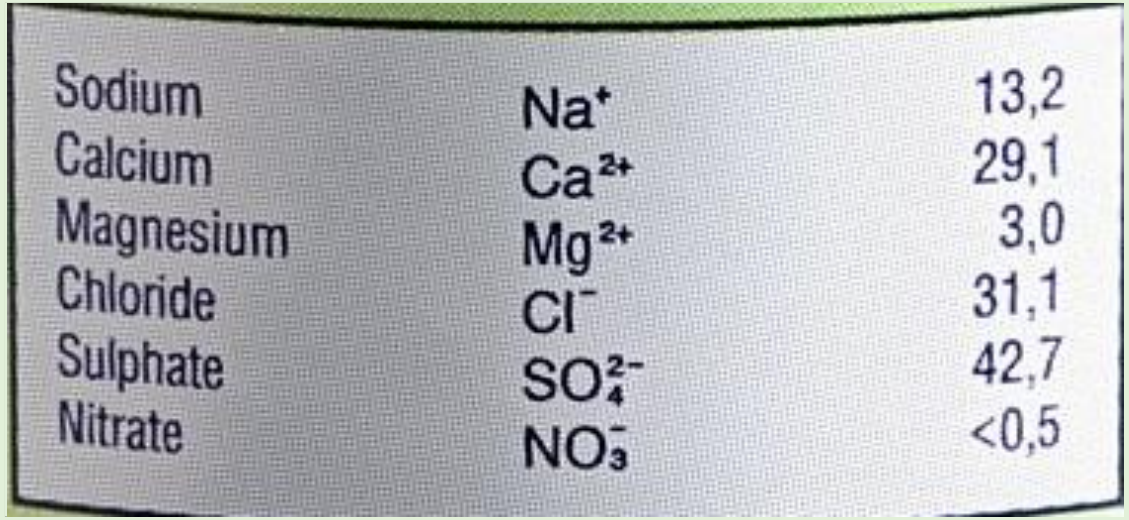
why you think ‘100% Pure Orange Juice’ is not ‘pure’ in the scientific sense?

Pure substances are made of only one type of atom or molecule.

Orange juice, even fresh squeezed from the orange is made of multiple different substances, including sugar, citric acid, vitamin C and water.

1. 'Pure' substance

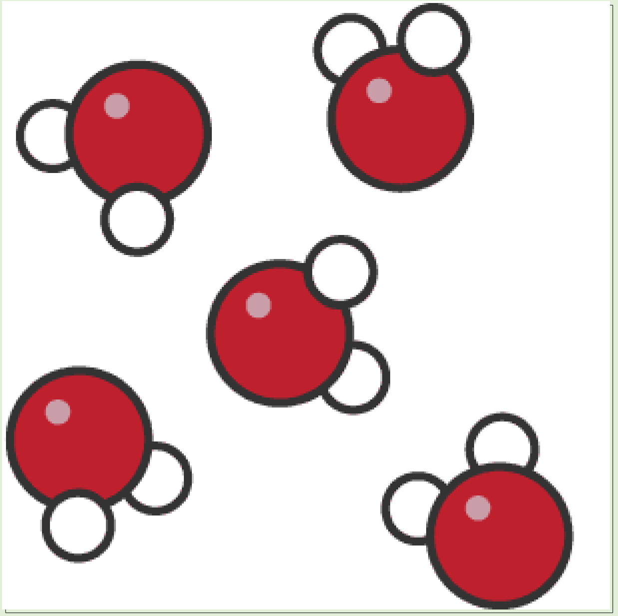
Substances like mineral water are not pure to a scientist. In science, a pure substance contains only one element or compound.

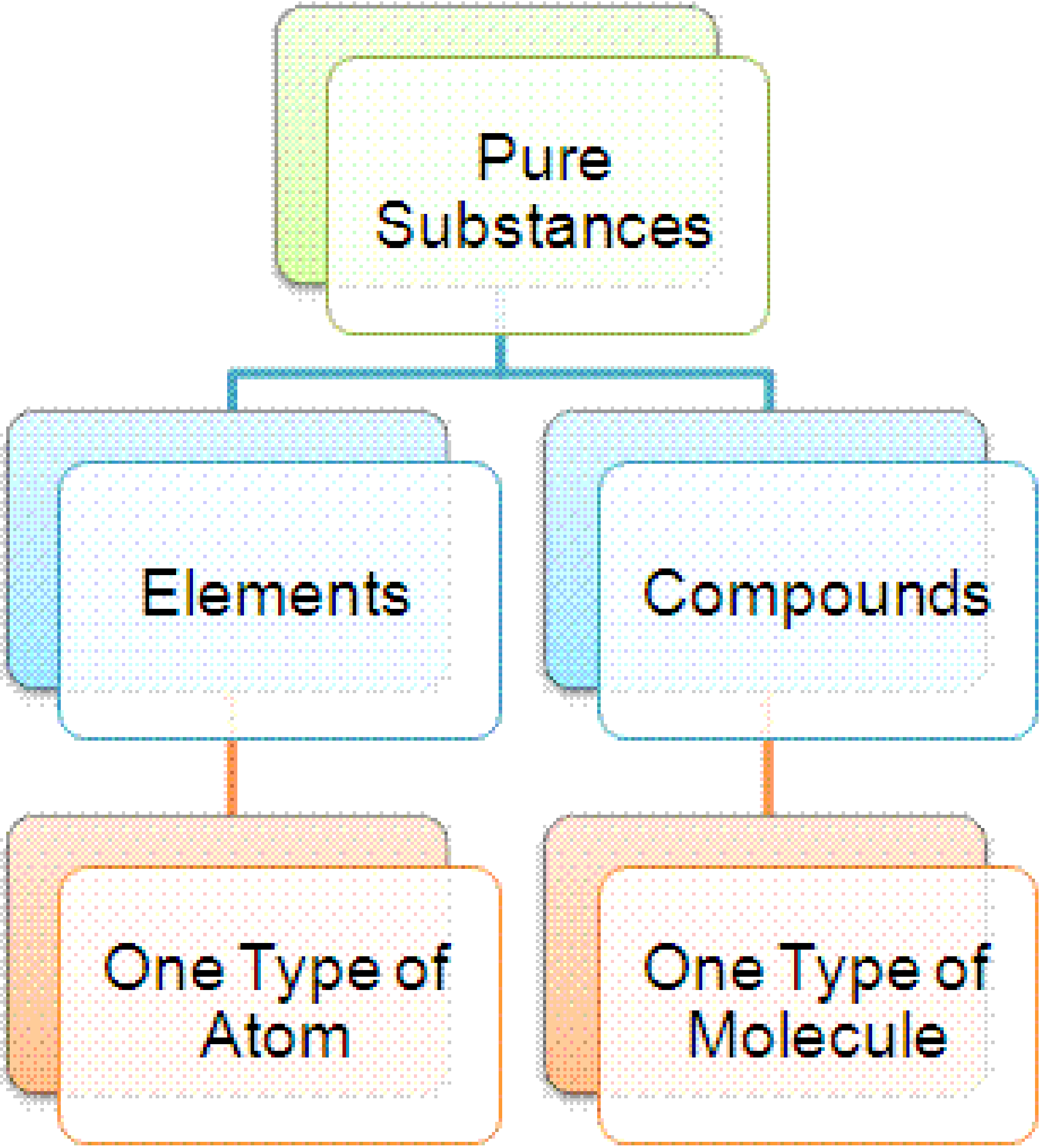


Sodium	Na^+	13,2
Calcium	Ca^{2+}	29,1
Magnesium	Mg^{2+}	3,0
Chloride	Cl^-	31,1
Sulphate	SO_4^{2-}	42,7
Nitrate	NO_3^-	<0,5

2. 'Pure' substance

Completely pure Water consists of only WATER molecules.




Pure Substances
Elements
Compounds
One Type of Atom
One Type of Molecule

3. What is an impure substance?

White paint is not pure and is a mixture.

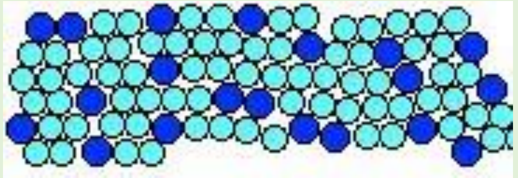
It is a mixture of many substances, like water and white pigment.

Hence, mixtures are impure substances. They contain > one element or compound.

Useful Mixtures

- Air is a mixture of Gases

- Alloys are mixtures of metals with one or more other elements added to them.



- These mixtures are designed to change the property of the metal. For example, pure gold is too soft to use for making jewellery but alloying (mixing it) with copper retains the lustre but makes it harder.

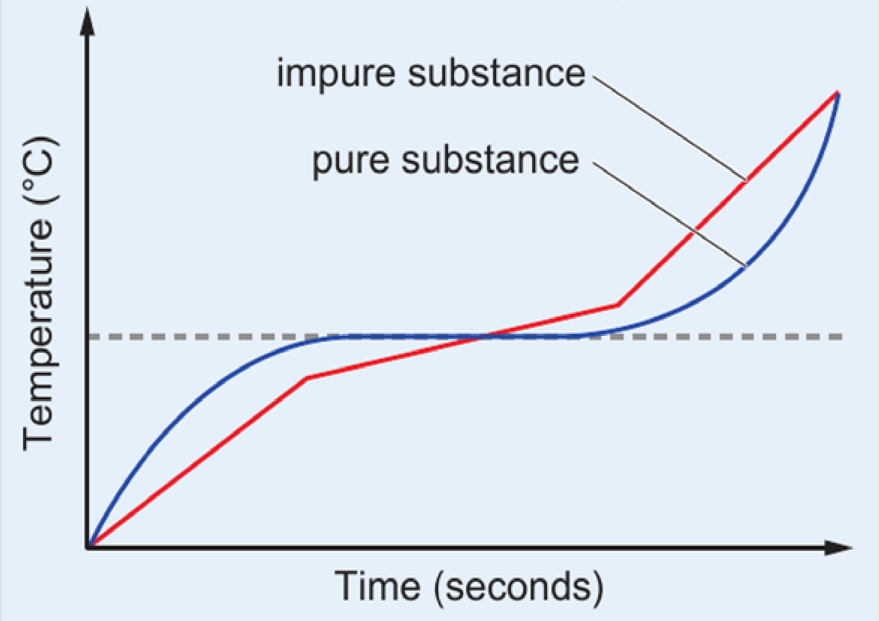
How do we use melting points to determine purity?

Use the temperature at which substances changes from a solid to a liquid to test for pure substances. Pure substances have a sharply defined melting point.

Impure substances have a temperature range over which they melt.

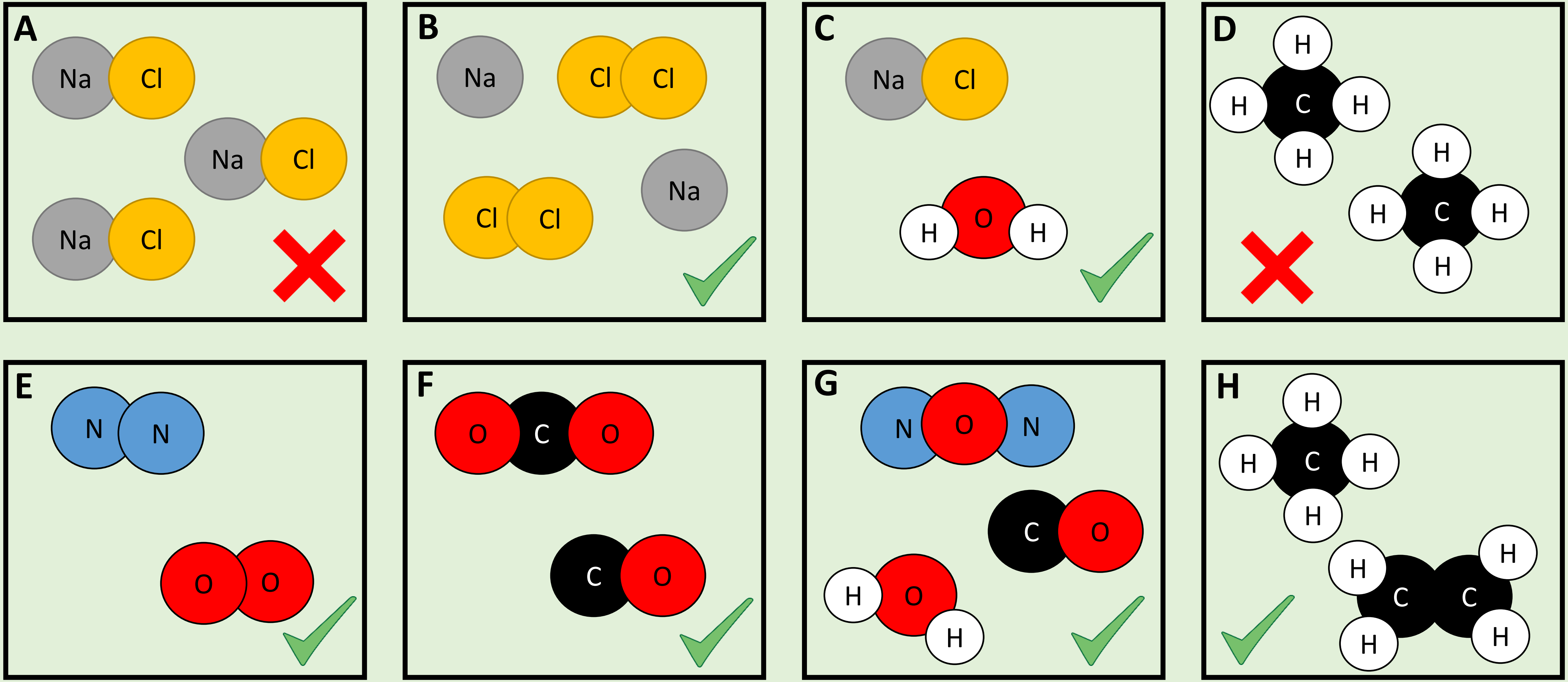
Impurities change melting points. E.g. impurities in water can cause it to boil above 100 °C

This means we can test the purity of a substance by melting or boiling it.



Filtration and crystallisation

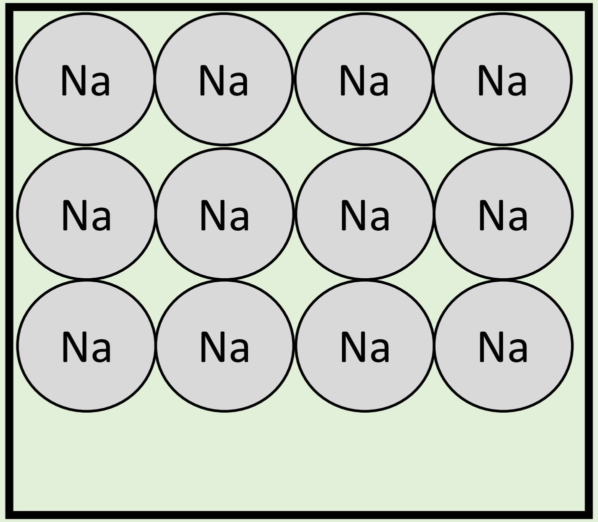
Which of these boxes contain mixtures?

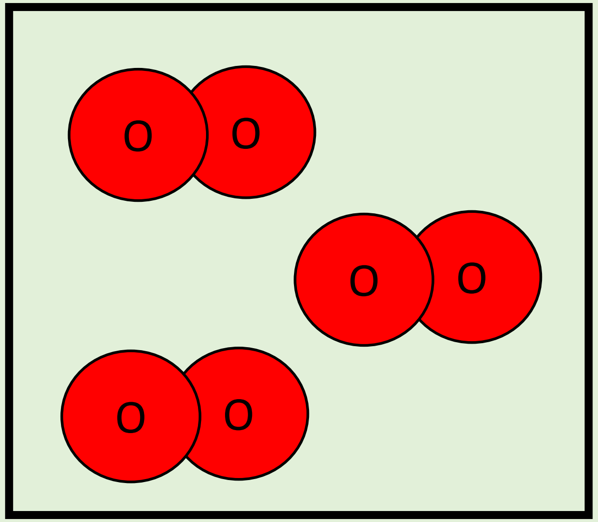


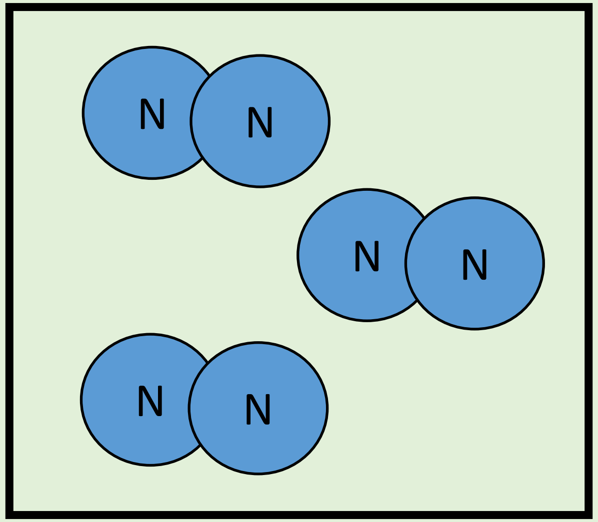
Recap

Element

An element is a pure substance made from only one type of atom.

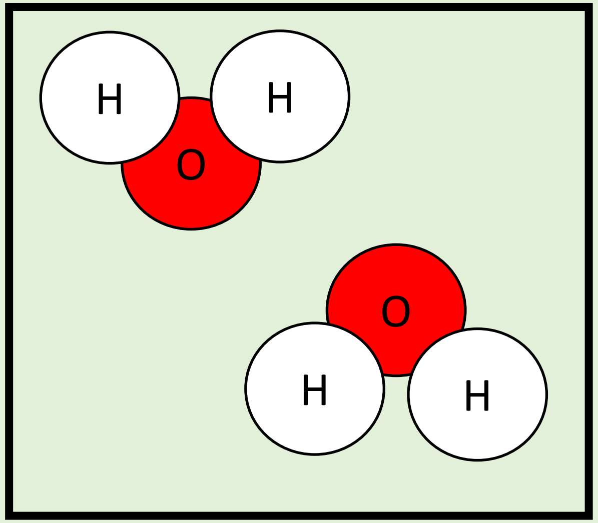


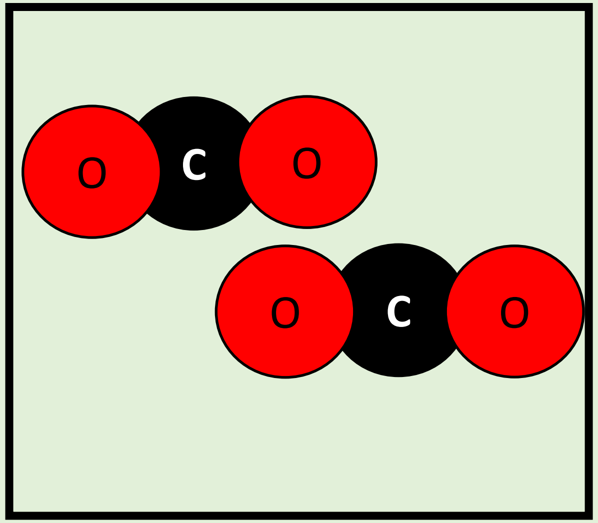


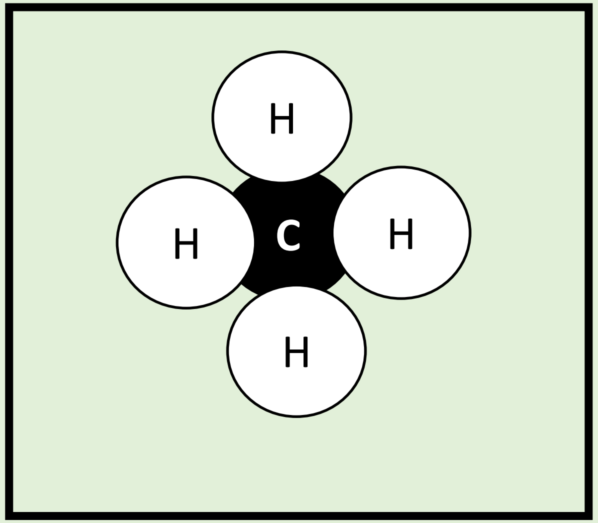


Compound

A compound is made from two or more different elements, which are *chemically bonded* together.

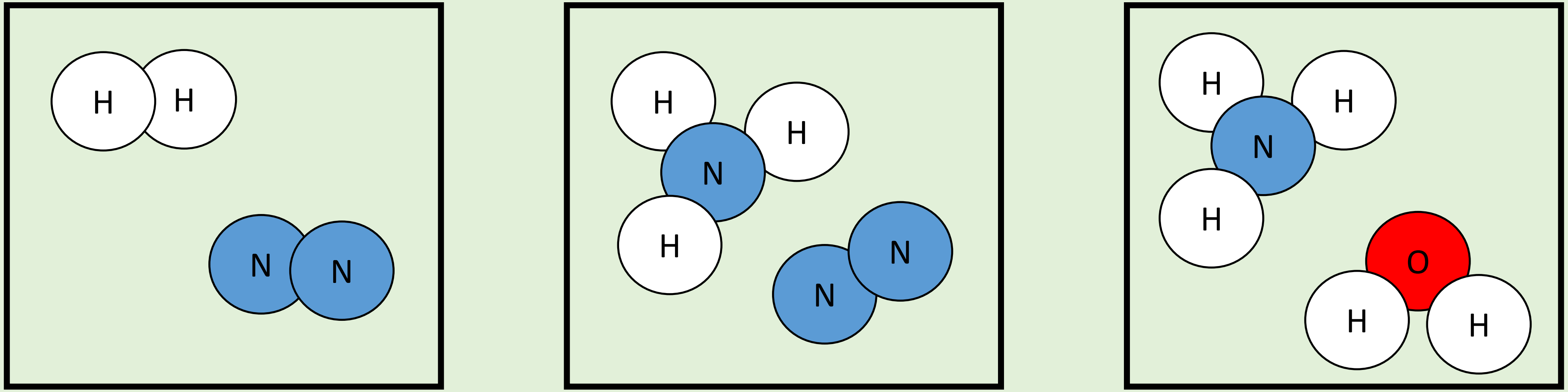






Mixture

A mixture contains two or more *different* elements or compounds that are *not* chemically bonded.



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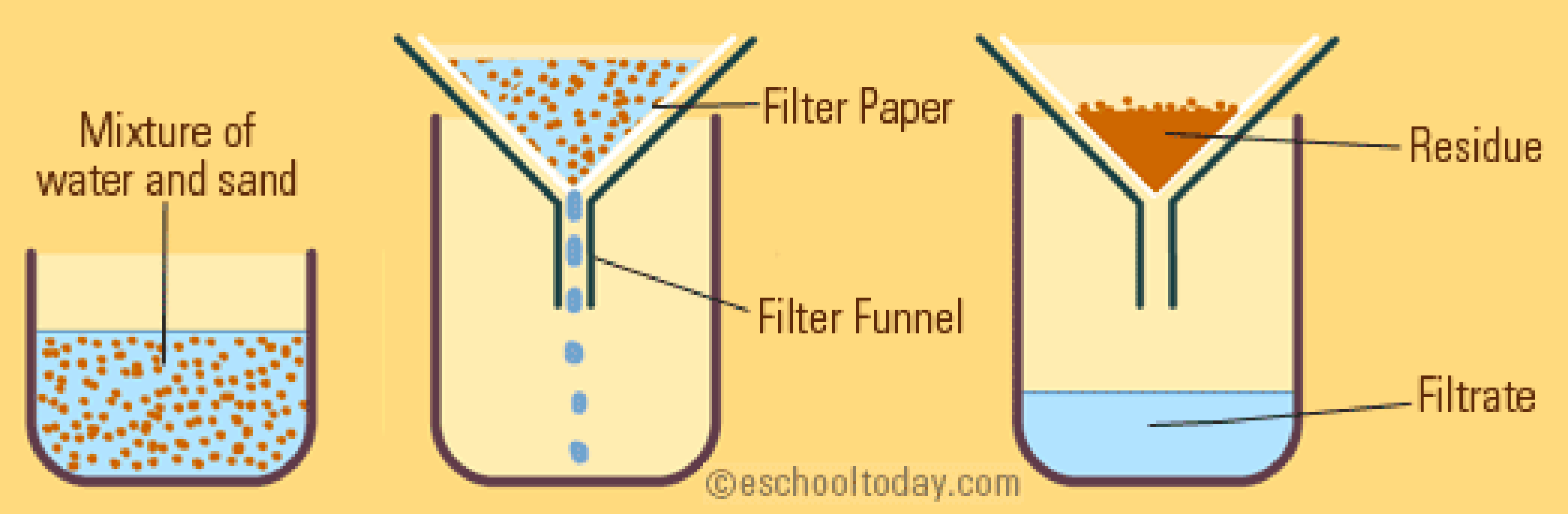
How would you separate this mixture?

water

Copper sulphate crystals

Hint- equipment you will have:
Funnel
Tripod
Evaporating basin
Filter paper
Gauze
Bunsen beaker
Heat proof mat
Beaker

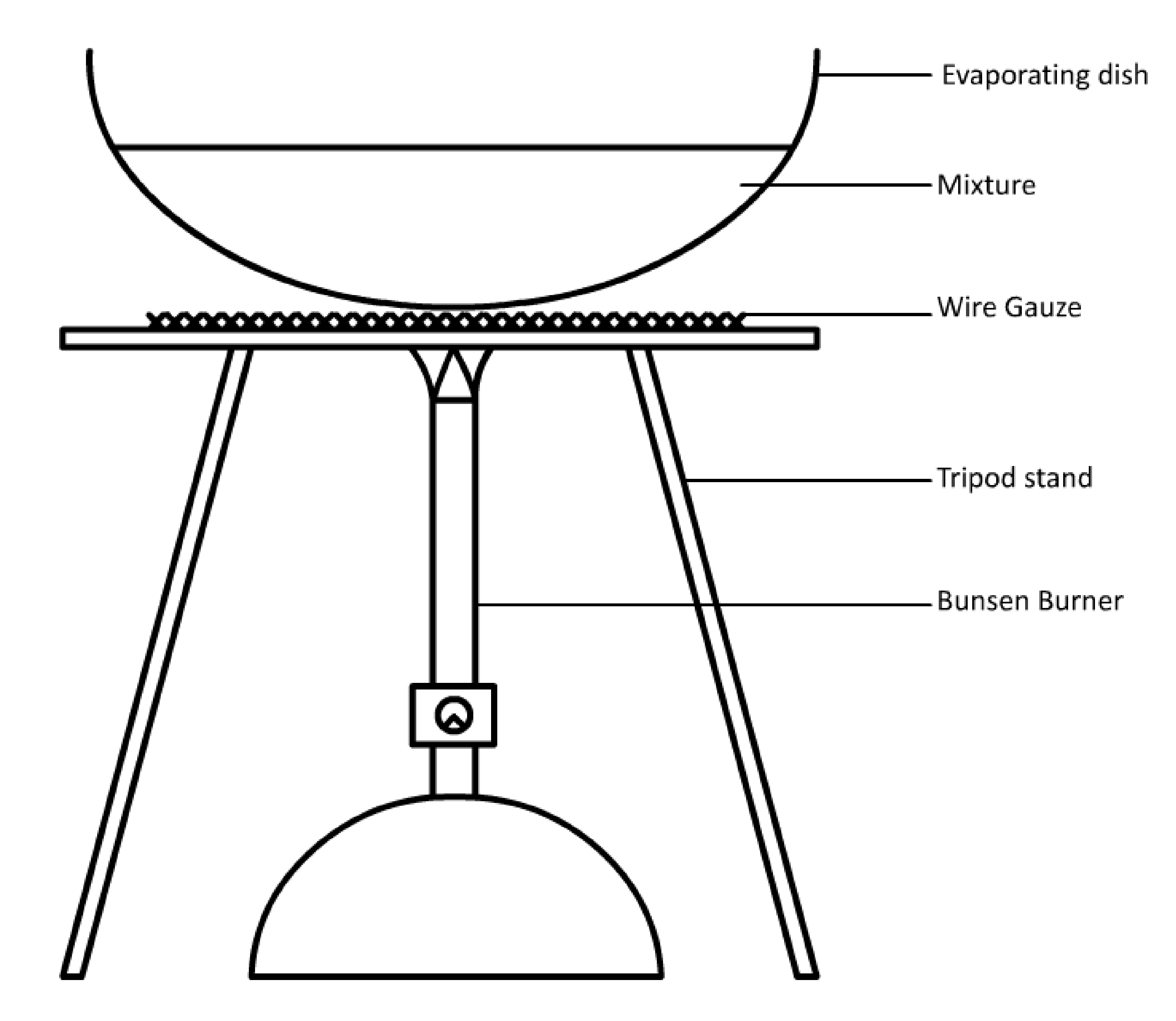
Filtration



Questions:
For our experiment which part is soluble?
For our experiment what will the residue be?
What physical property are we exploiting?

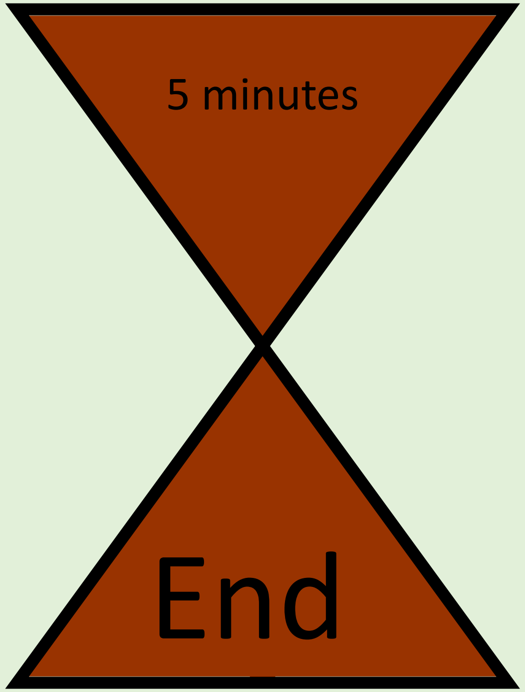
Crystallisation

What physical property are we exploiting?



Exam Question:

How would you separate a mixture of salt and wood chips? (6 marks)



Exam Question:

How would you separate a mixture of salt, water and wood chips? (6 marks)

- Salt is soluble in water.
- Therefore it can be separated by filtration.
- The water and salt mixture will be the filtrate and the wood chips will be the residue.
- Salt an water have different boiling points.
- Salt and water can be separated using crystallisation.
- The water will evaporate and the salt will remain as crystals.

Distillation

- Mixtures of liquids and solids can be separated by separation techniques.

How would you separate pasta from the cooking water?

How are tea leaves kept out of a cup of tea?

How are coffee grains separated from a pot of fresh coffee?

- Insoluble solids can be separated from a liquid by **filtering**.

How can you separate the **solvent** from the solids dissolved in a solution?

- The best method to separate and collect the solvent from a solution is **distillation**.

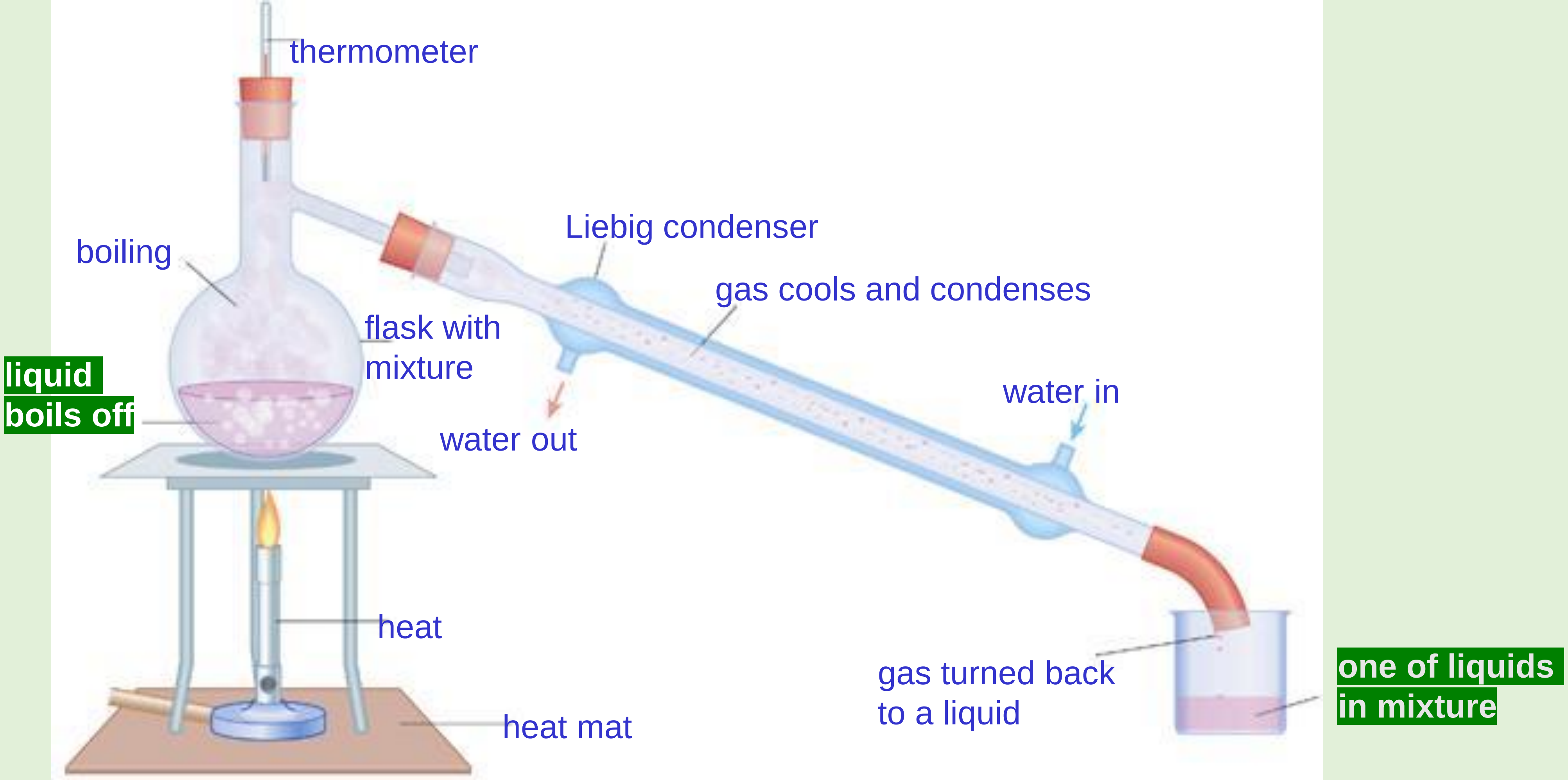
- Distillation involves three stages:

boiling

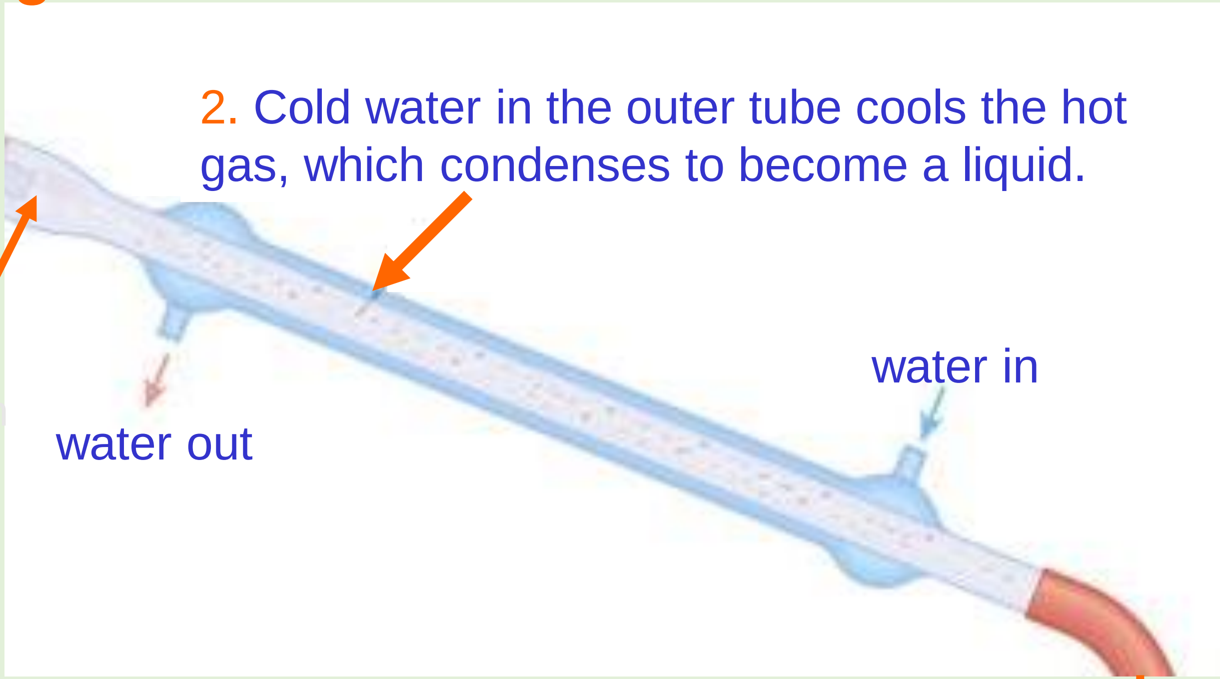
condensing

collecting

Where in the distillation equipment do boiling, condensing and collecting take place?



The most important piece of kit for distillation is the **Liebig condenser**.



Cold water enters the outer tube at the bottom and warm water leaves the outer tube at the top. Why does this happen?

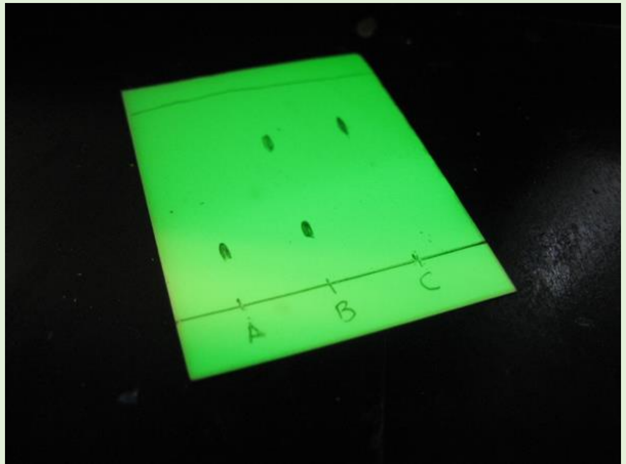
Chromatography

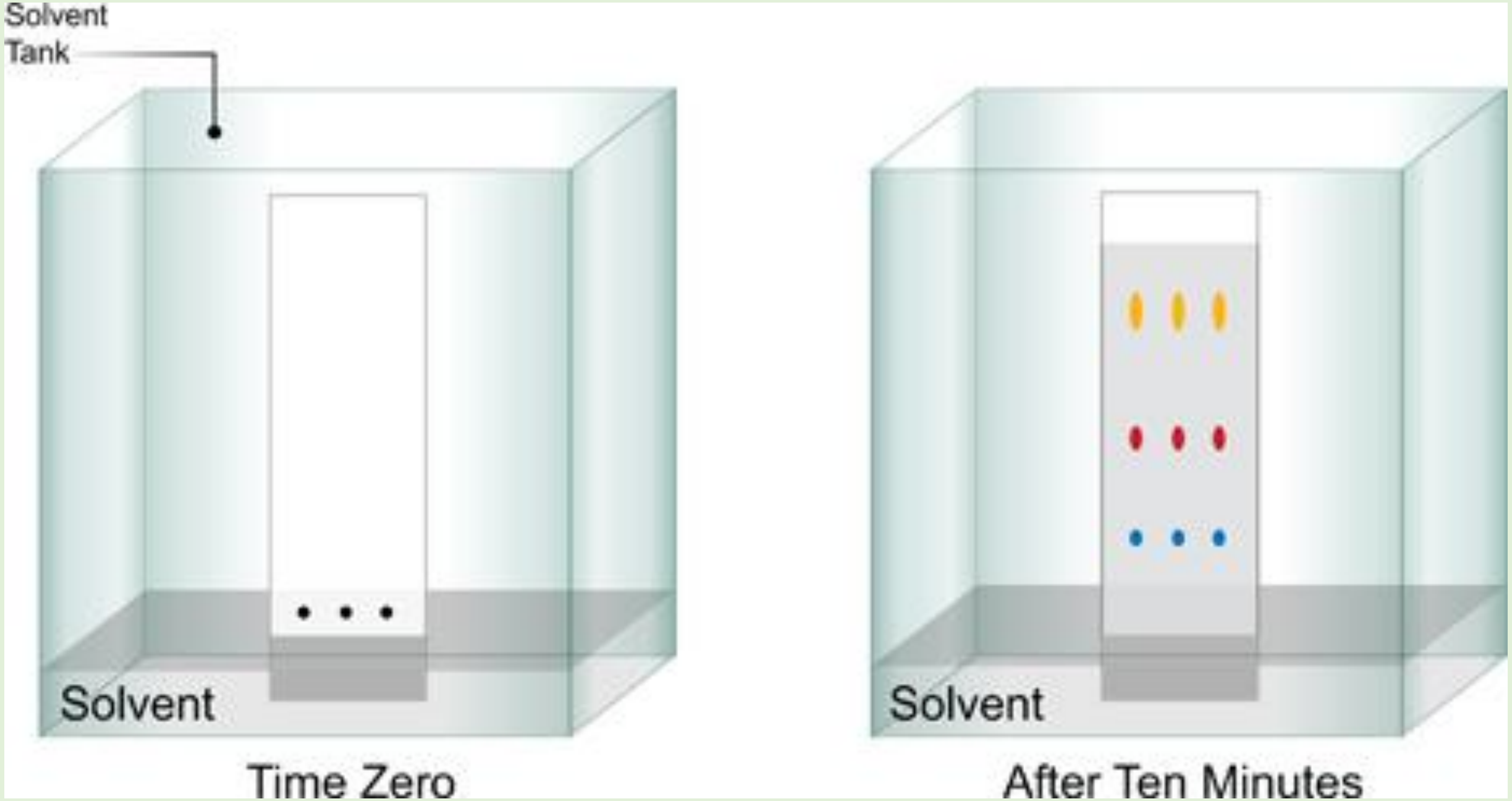
Chromatography

- Separation technique that relies on the affinity of components of a mixture to:-
  - **A Mobile Phase**
  - **A Stationary Phase**
- All types of chromatography rely on this
  - Paper chromatography
  - thin layer chromatography,
  - gas chromatography

Chromatography

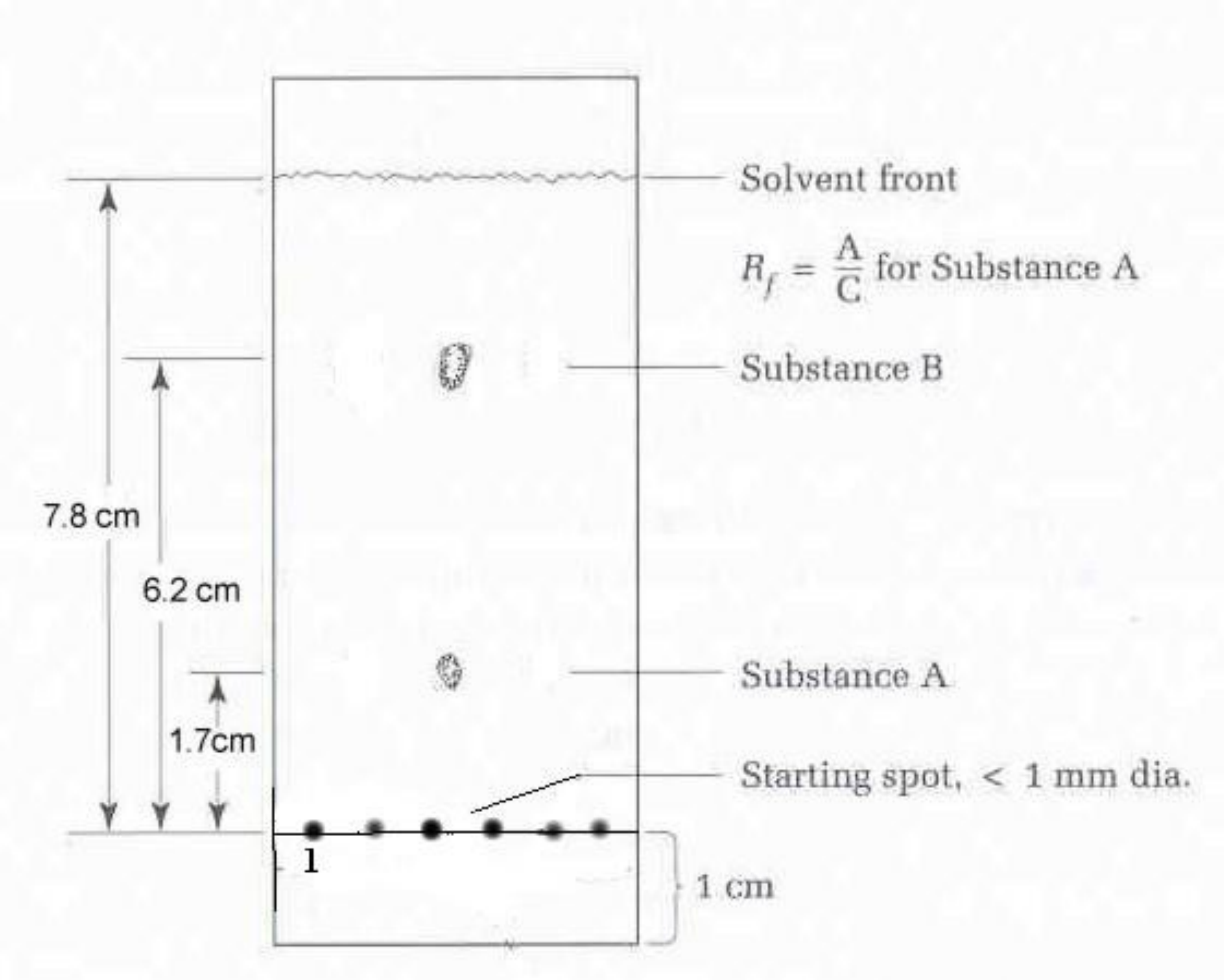
- **Mobile Phase**
  - **This is the phase that is a carrier for the components as they travel through the stationary phase**
- **Stationary Phase**
  - **This is the phase that is stationary!!**




Solvent
Tank
Solvent
Solvent
Time Zero
After Ten Minutes

Rf Values (retention factor)

- Distance travelled by component
- Distance travelled by mobile phase
- Specific to each plate
- Comparable
- Each component on plate has unique Rf value



Gas Chromatography

- Mobile phase – unreactive gas (helium, nitrogen)
- Stationary phases – polymer or liquid on an inert solid support
- Widely used in analytical laboratories
- Quantify components – volatile mixtures
- Retention times measured and used by a computer to identify components

Metals & non-metals

- a comparison of properties

metals are ...

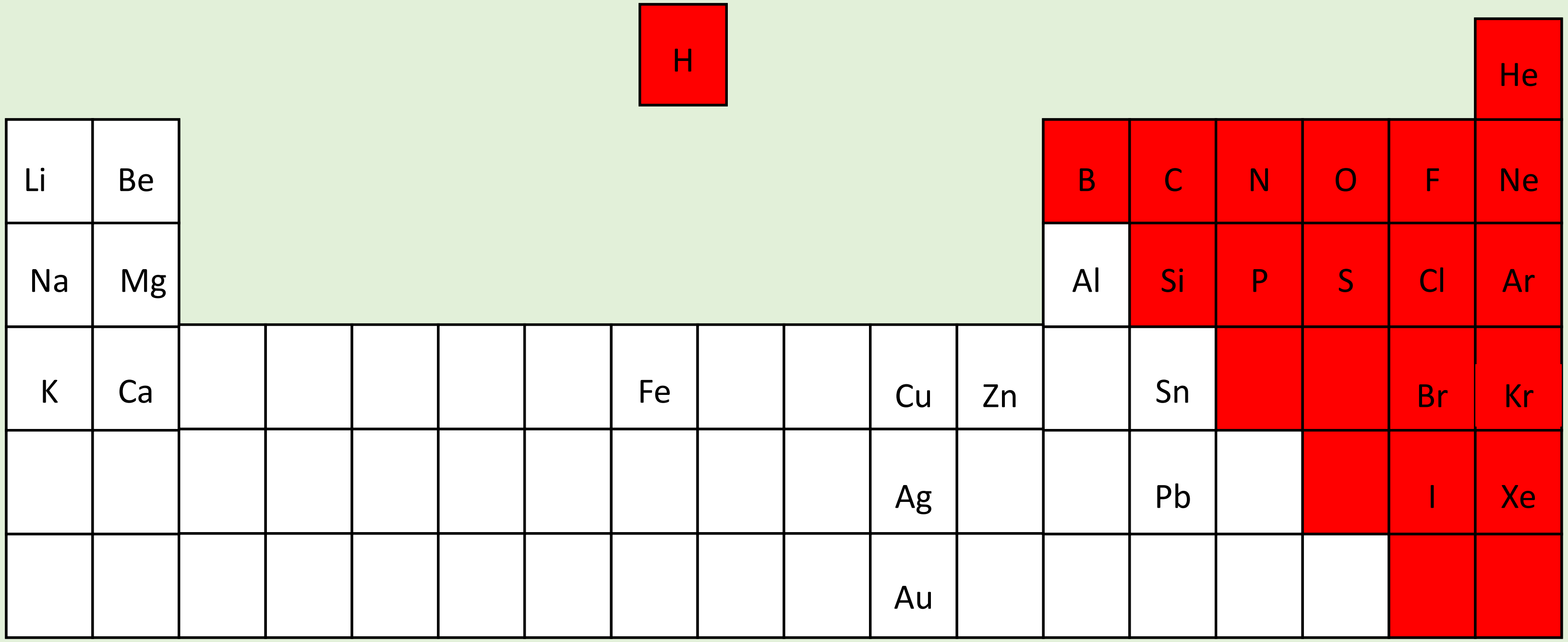
- SHINY
- MALLEABLE & DUCTILE
- GOOD conductors electricity in the SOLID STATE
- usually have HIGH melting points
- form OXIDES known as BASES
- Reactivity INCREASES <u>down</u> a group

non-metals are ...

- usually DULL
- usually BRITTLE
- POOR conductors electricity in the SOLID STATE
- usually have LOW melting points
- form OXIDES that are ACIDIC
- Reactivity DECREASES <u>down</u> a group

Metals & non-metals can be located in the Periodic Table

Metals are found on the left of the periodic table and non-metals are found on the right of the periodic table. You can see that there are more metals than Non-metals.



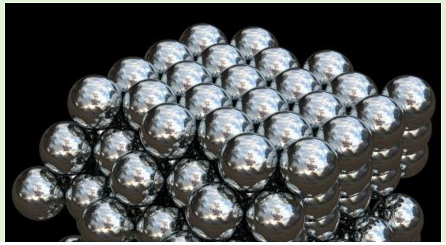
shows the non-metals

The group number lets us know how many electrons are in the outer shell of the atom, whereas, the period number relates to the total number of shells surrounding the nucleus of an atom

Why do metals usually have HIGH melting points?

Any substance that has a HIGH melting point must

- have a GIANT structure
- have a STRONG BONDS



Electronic structures

Rules for Filling Shells

- The inner shell must be full before electrons occupy the next shell
- The first shell contains a maximum of **TWO** electrons
- Subsequent shells have a maximum of **EIGHT** electrons

Electronic Structure

To which Groups and Periods do the following belong?

Electron arrangement	Group	Period	Name of element
2,3	3	2	boron
2,8,8,1	1	4	potassium
2,8,7	7	3	chlorine
2,8	8 or 0	2	neon

RECAP: Types of bonding

Atoms can be joined together in 3 possible ways

Metal and non-metal → **Ionic Bonding**

Non-metals only → **Covalent Bonding**

Metals only → **Metallic Bonding**

All three types involve changes in the electrons in the outermost electron shells of the atoms

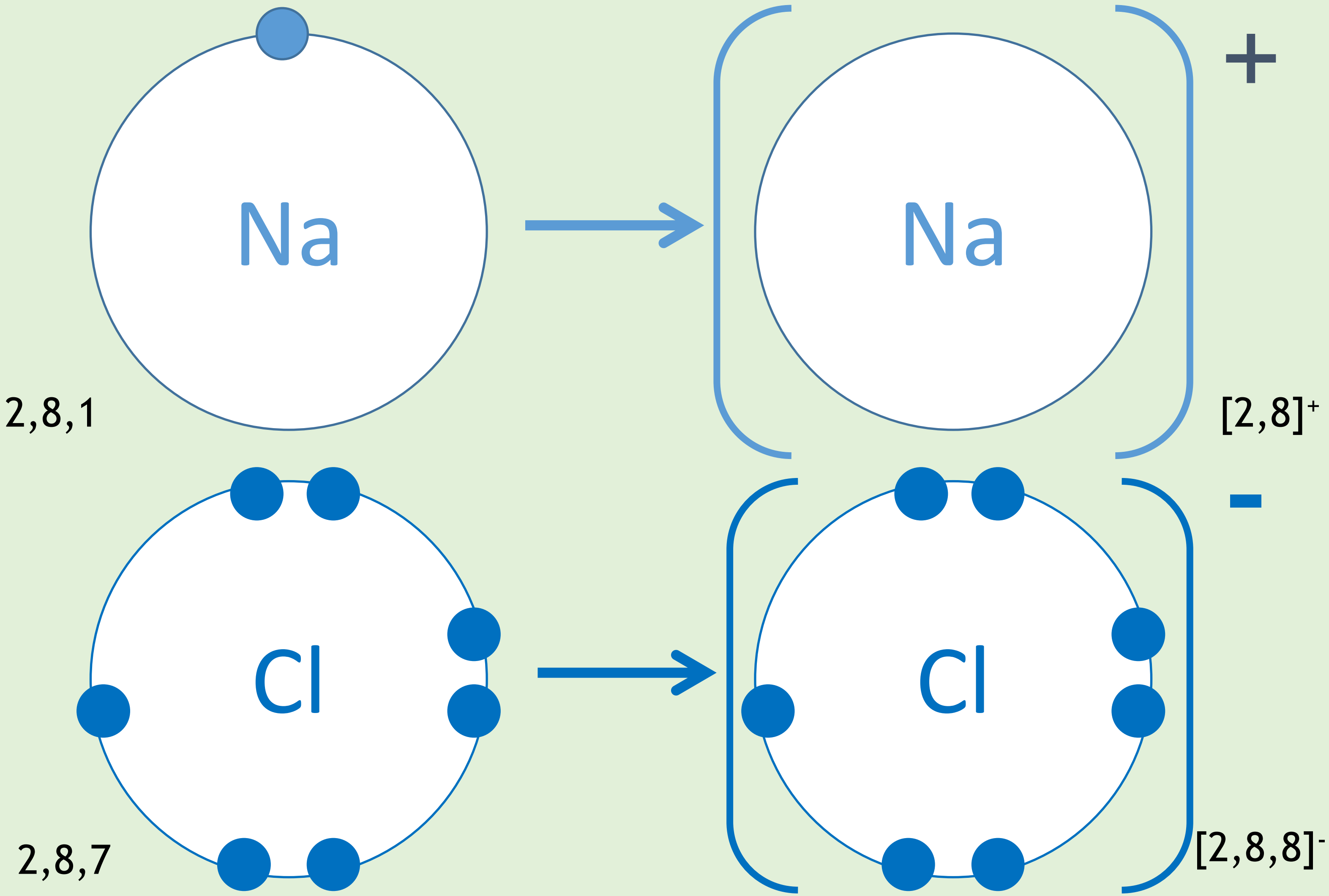
Forming ions

You will need your periodic table

why do atoms bond?

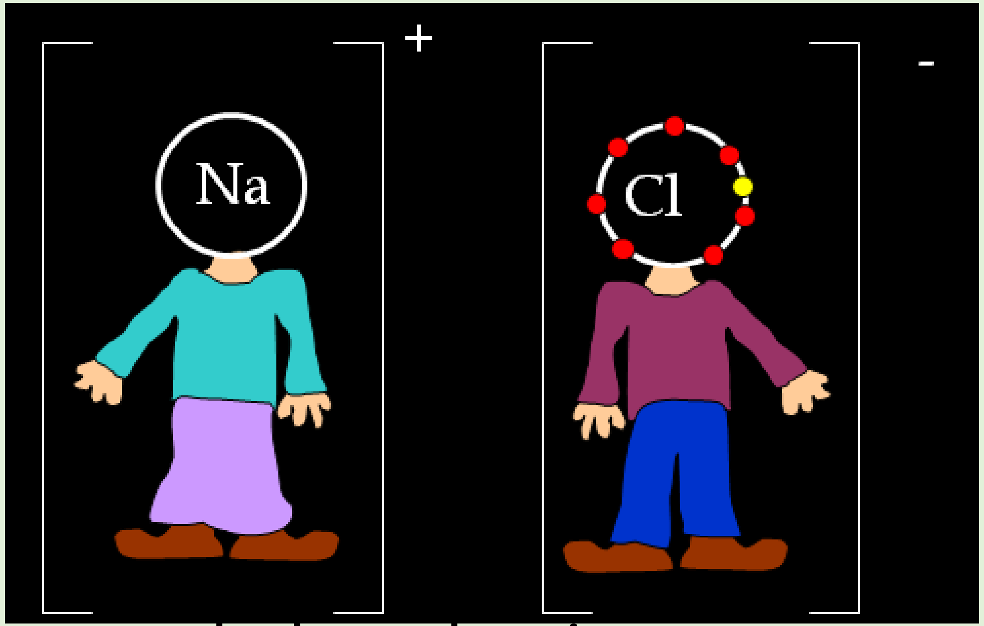
- Atoms bond to get full outer shells.

how do atoms form ions?



WHAT IS AN IONIC BOND?

- The force of attraction between the positively charge ion and negatively charged ion.

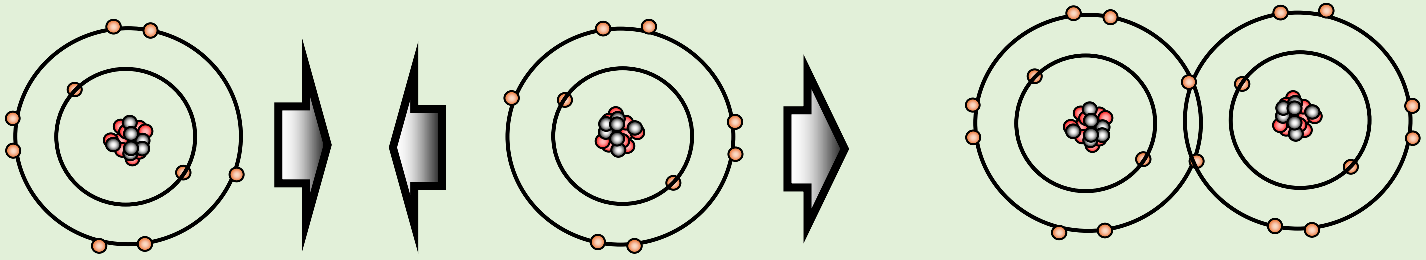


- One element is a metal, the other is a non-metal.

COVALENT BONDING

Covalent compounds

- Covalent compounds are formed when non-metal atoms react together.
- As these atoms come near their outer electrons are attracted to the nucleus of both atoms and become shared by the atoms.
- The shared electrons count towards the shells of both atoms and therefore help fill up incomplete electron shells.



Covalent bonds

- Covalent compounds are held together by this sharing of electrons.
- A pair of electrons shared in this way is known as a **covalent bond.**
- It is sometimes represented in full bonding diagrams (see figure 1). Often these bonds are just shown as a pair of electrons (xx) or even just a line (see figure 2).

Figure 1

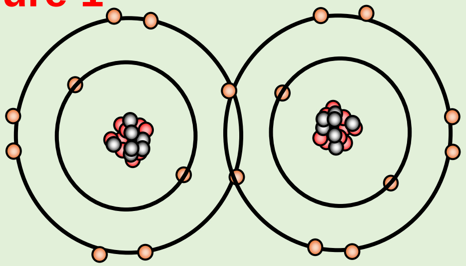


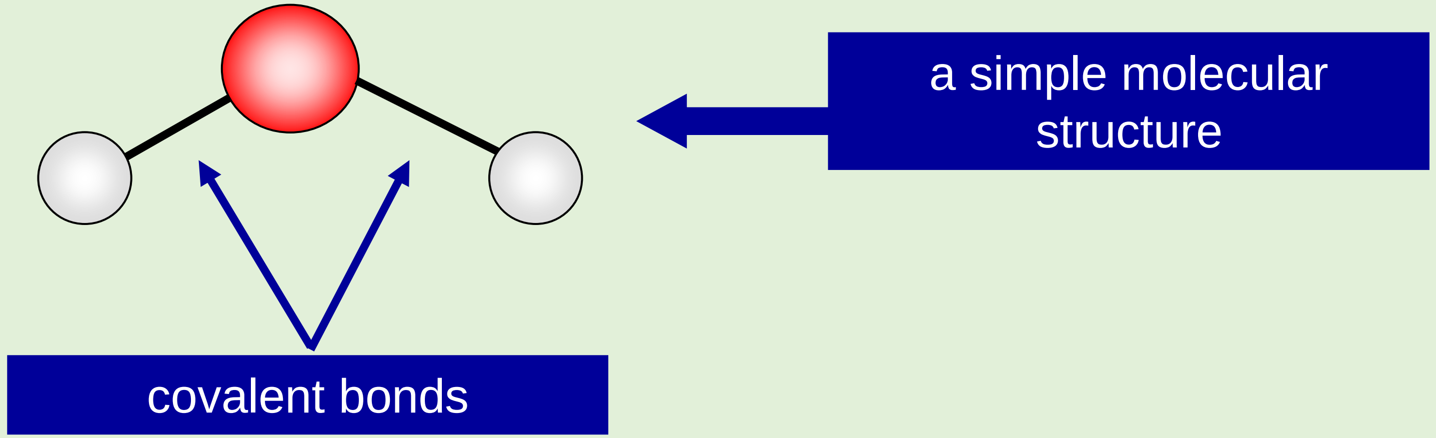
Figure 2

$F \genfrac{}{}{0pt}{}{x}{x} F$

$F - F$

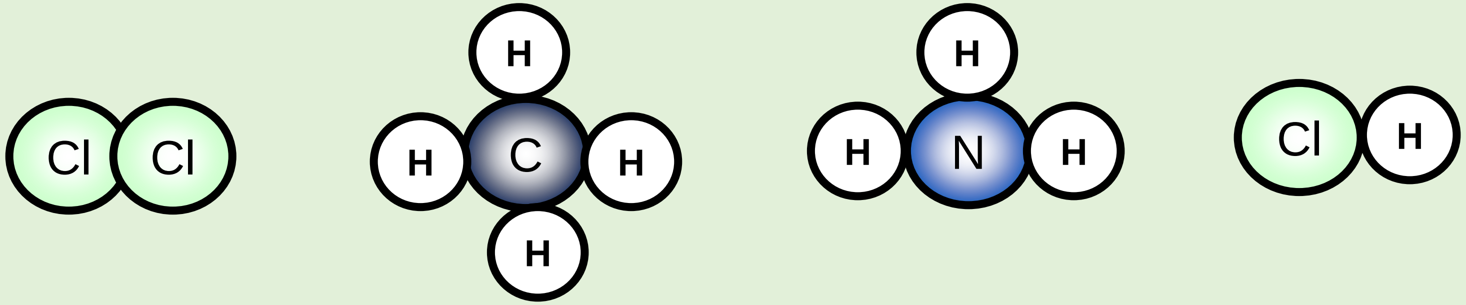
Small covalent structures (Simple Molecules)

- Sometimes just a few atoms join together in this way.
- This produces small covalent molecules – often known as simple molecular structures.



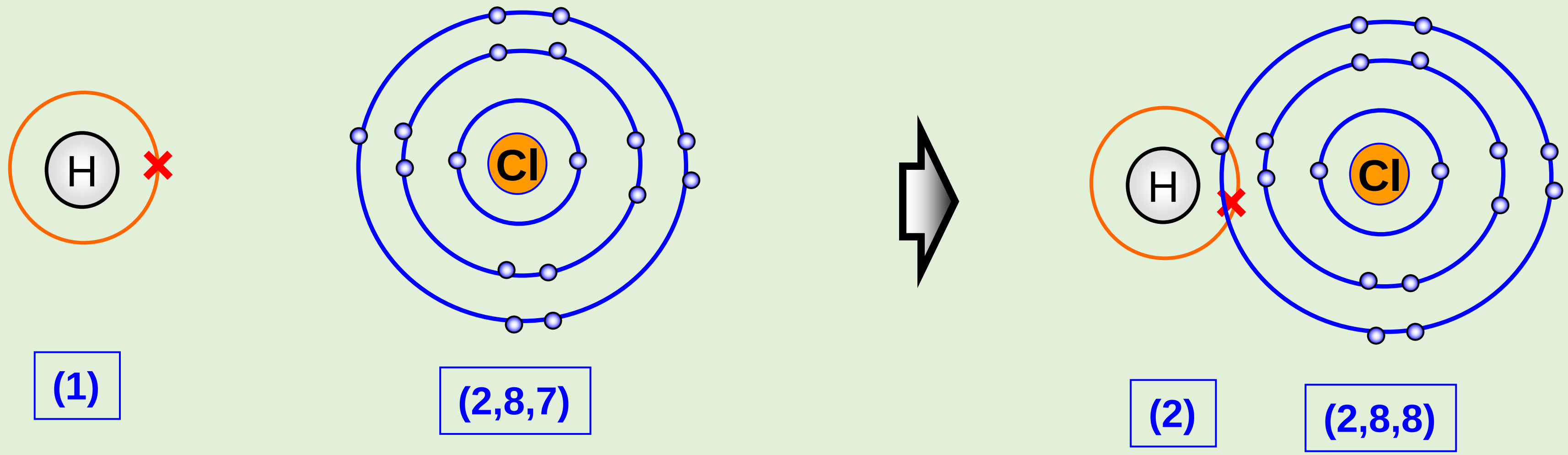
Covalent bonding and electron structures

- The driving force for covalent bonding is again the attainment of outer electron shells that are completely full.
- This is achieved by sharing electrons where the shared electrons count towards the outer shells of both atoms.
- Sometimes this is achieved with equal numbers of each type of atom.
  Sometimes it is not!



Covalent bonding in hydrogen chloride

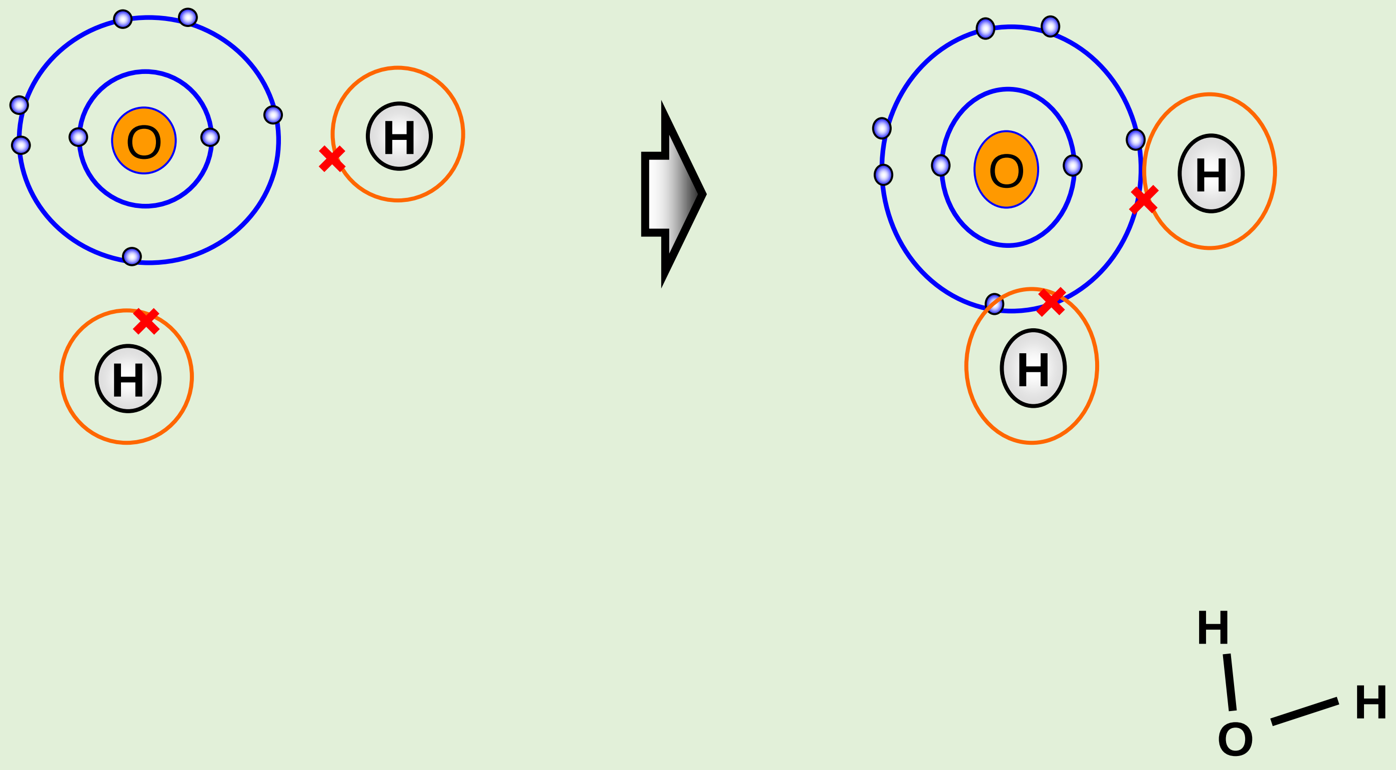
Both hydrogen (1) and chlorine (2.8.7) needs 1 more electron to attain a full outer shell.



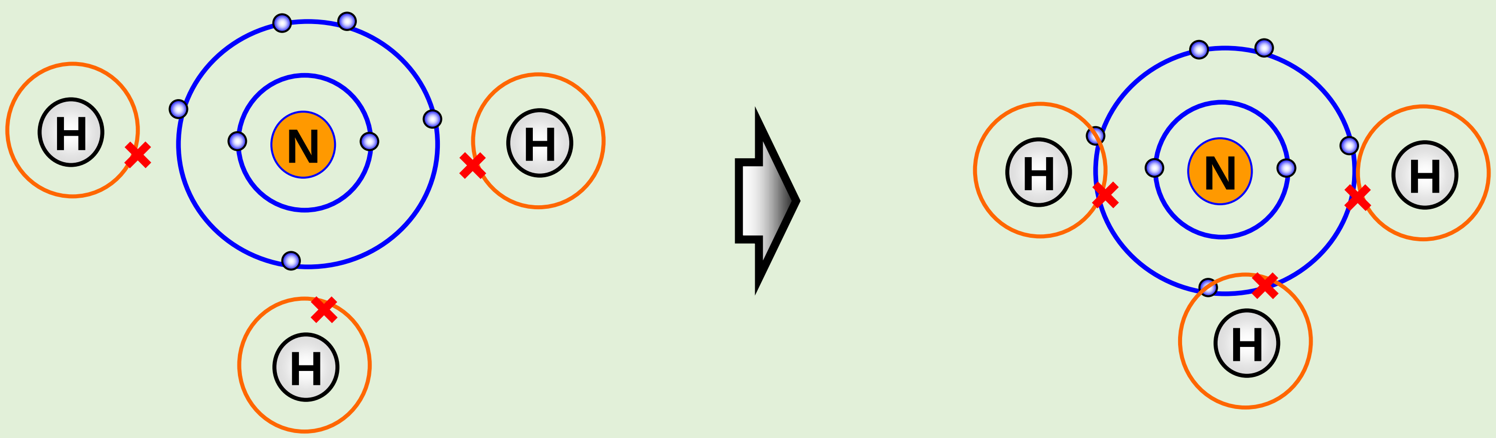
H-Cl

Covalent bonding in water

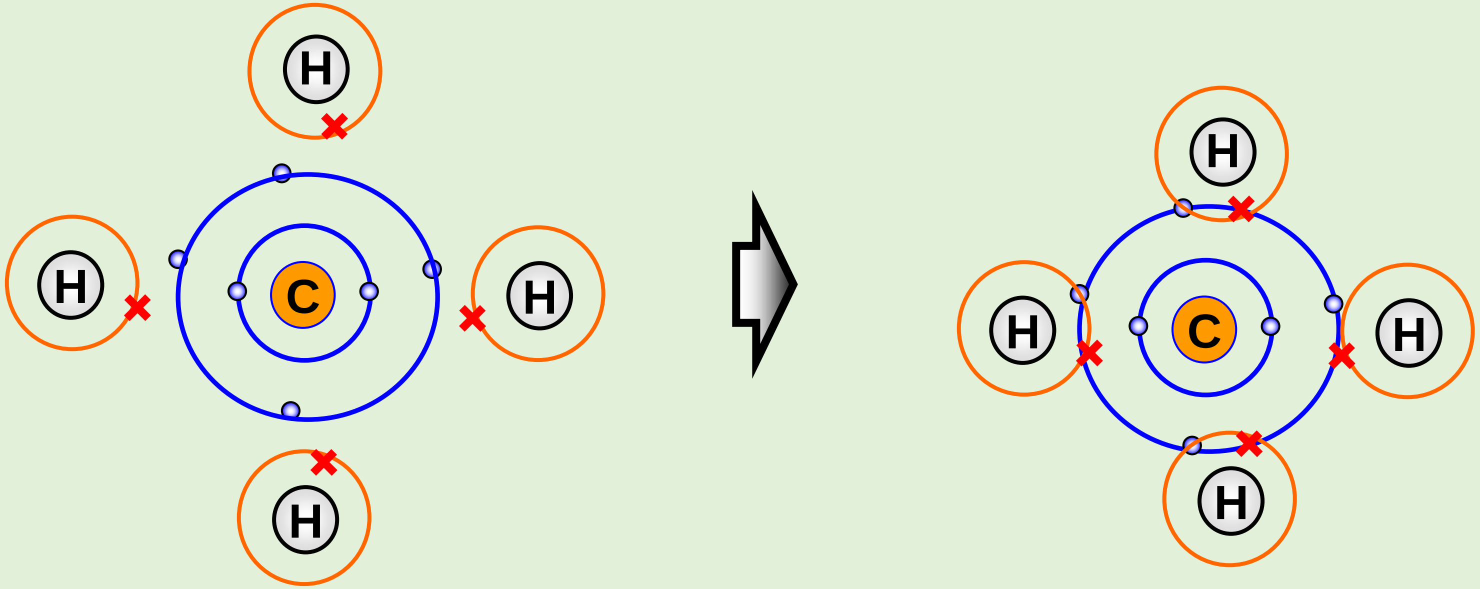
Hydrogen (1) needs 1 more electron but oxygen (2.6) needs 2 more. Therefore, we need 2 hydrogens.



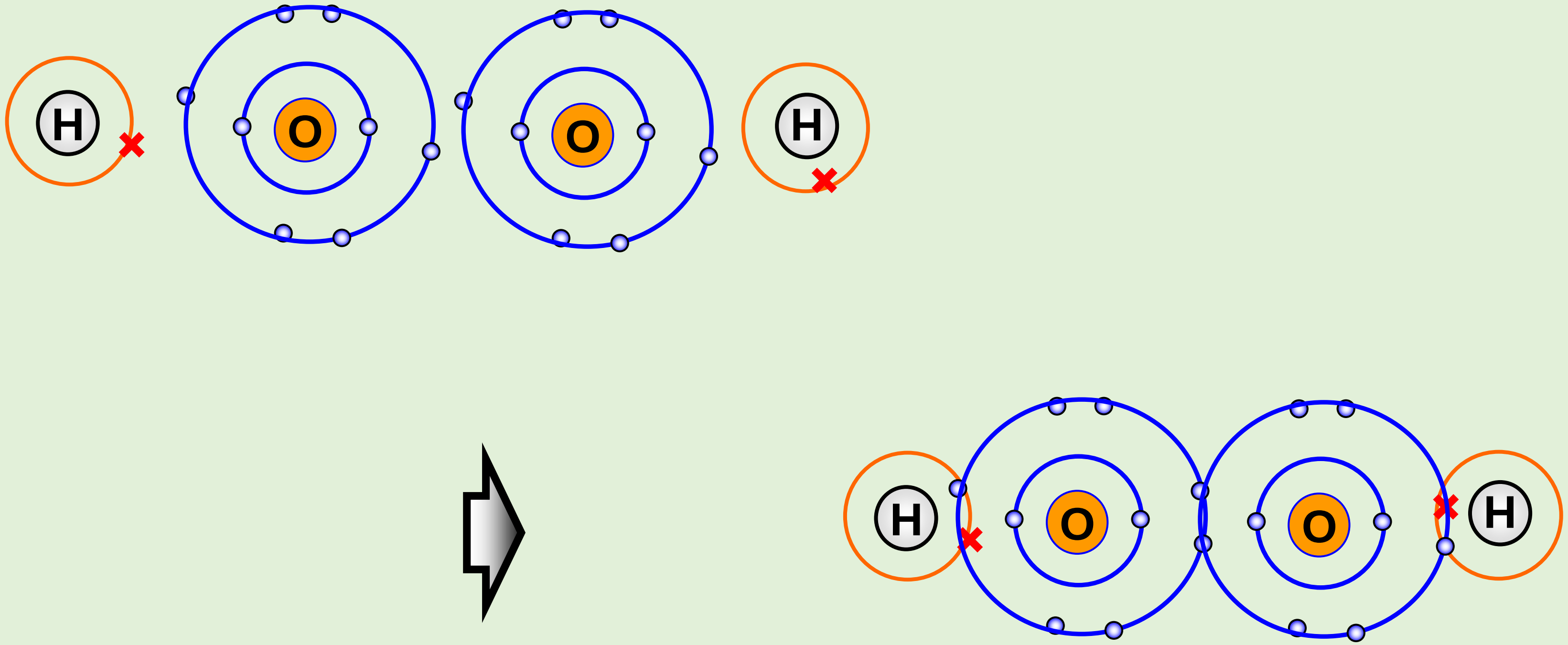
- Hydrogen (1) needs 1 more electron.
- How many does nitrogen (2.5) need? 3
- How many hydrogens per 1 nitrogen? 3
- Draw bonding diagrams for ammonia.



- Hydrogen (1) needs 1 more electron.
- How many does carbon (2.4) need? 4
- How many hydrogens per 1 carbon? 4
- Draw bonding diagrams for methane.



- Copy the atoms below.
- Complete the diagram showing how each atom can achieve full shells.



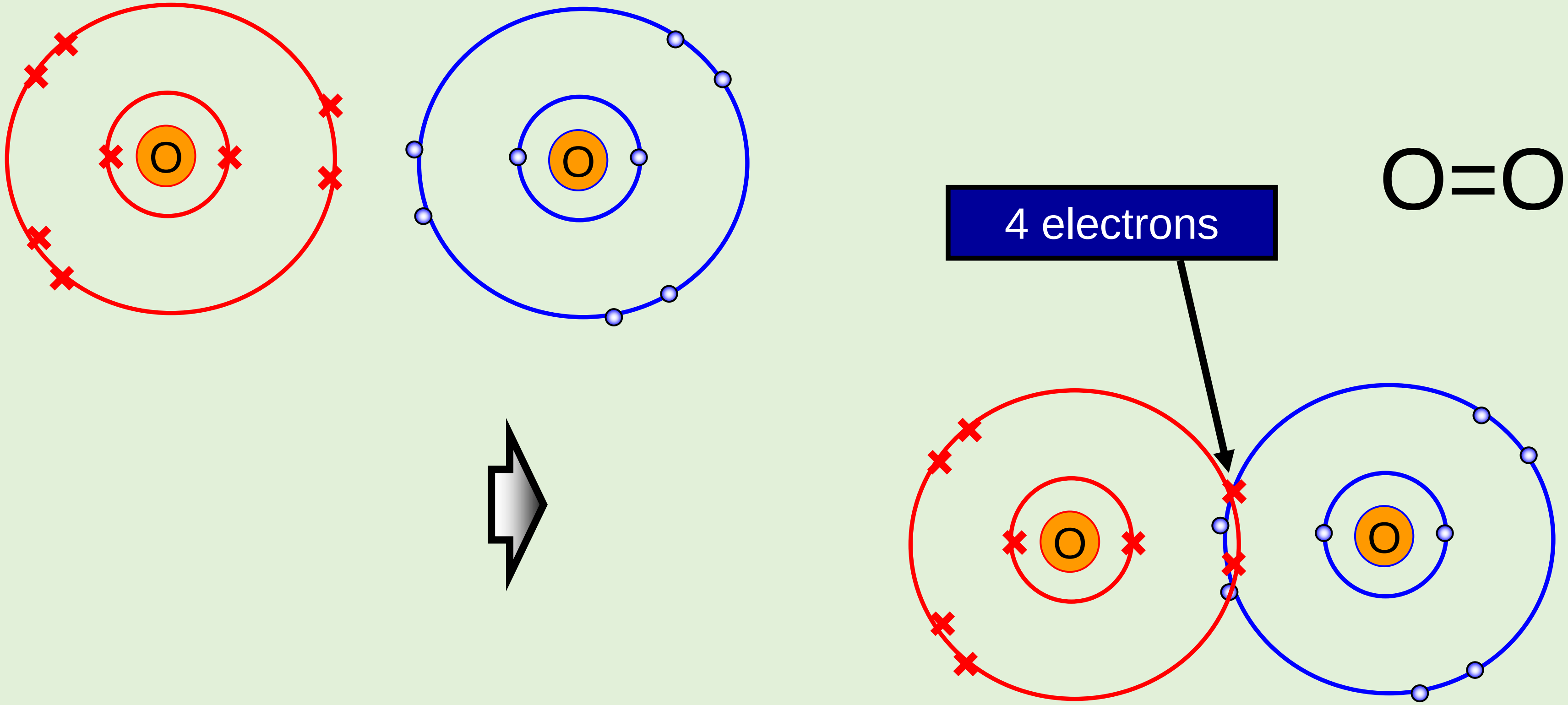
Covalent bonding - multiple bonds

- Mostly electrons are shared as pairs.
- There are some compounds where they are shared in fours or even sixes.
- This gives rise to single, double and triple covalent bonds.
- Again, each pair of electrons is often represented by a single line when doing simple diagrams of molecules.

Cl-Cl	O=O	N≡N
Single bond	Double bond	Triple bond

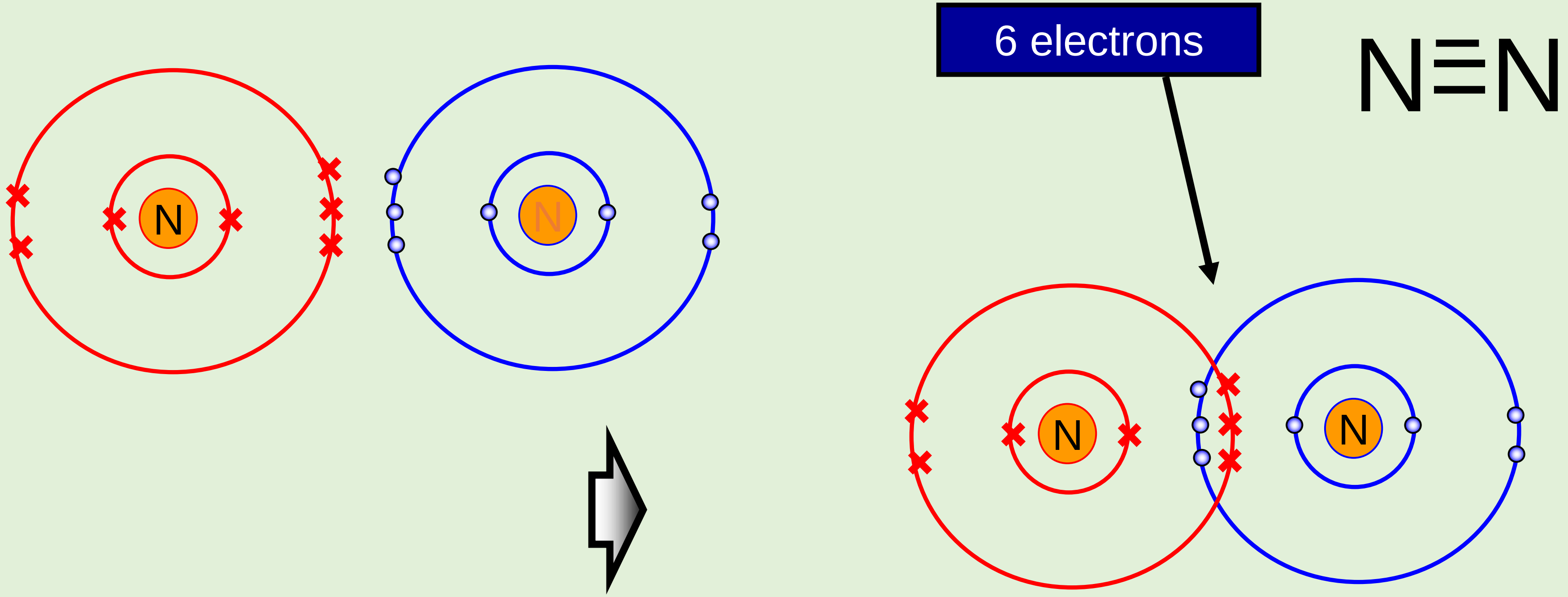
Covalent bonding in oxygen

Oxygen (2.8.6) needs 2 more electrons to attain a full electron shell.



O=O

Nitrogen (2.8.5) needs 3 more electrons to attain a full electron shell and forms a triple bond.
Draw a bonding diagram of nitrogen.

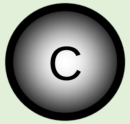


Giant covalent structures

Giant covalent structures

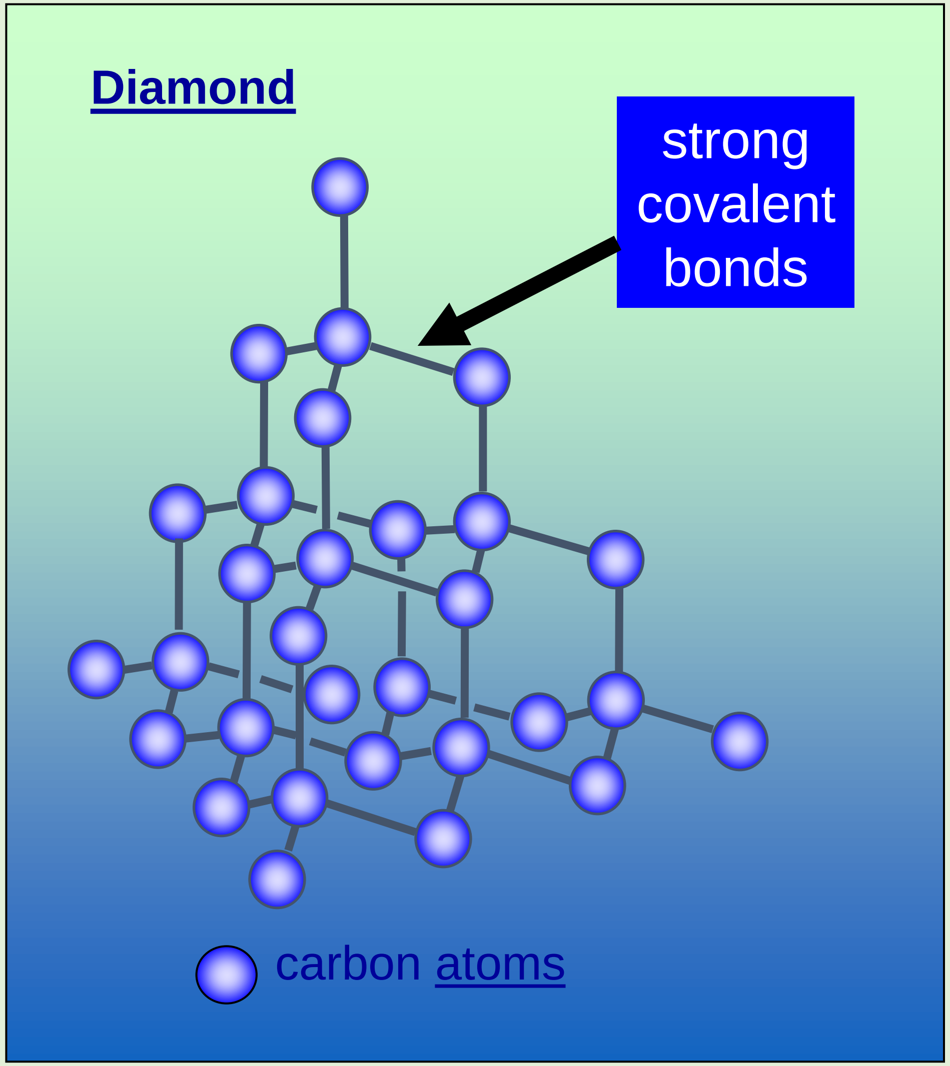
1. Carbon atoms form giant structures.
2. What is interesting is that there is more than one possible arrangement for the atoms.
3. Although this does not affect the chemical properties it can make a huge difference to the physical properties such as hardness, slipperiness, melting point and density.

Different arrangements of the same element are called ***allotropes***.



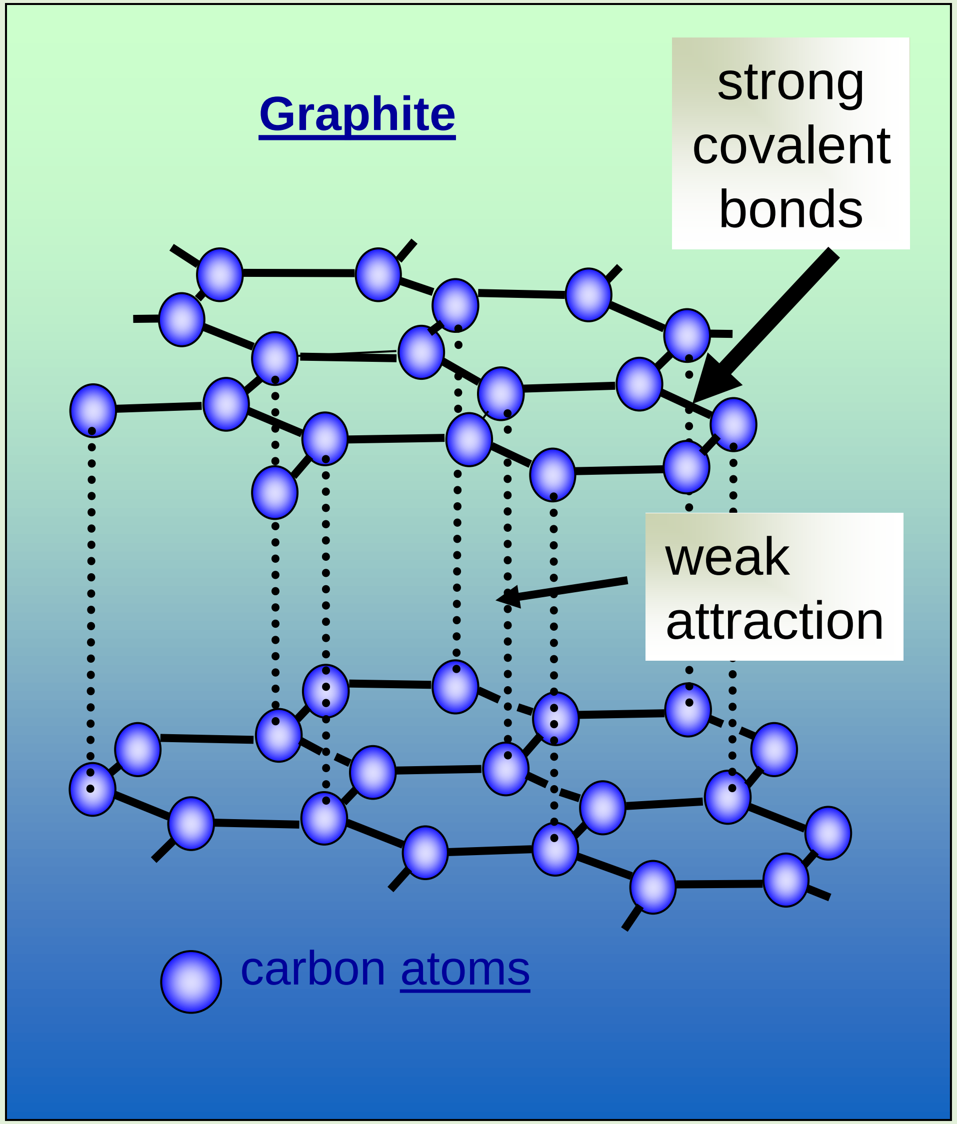
Giant covalent structures: diamond

- One form of carbon is diamond.
- Each diamond consists of millions of carbon atoms bonded into a single giant structure.
- It is **very** hard.



Giant covalent structures: graphite

- A more common form of carbon is graphite.
- Millions of carbon atoms are bonded into a giant structure but within this structure the layers are only weakly joined.

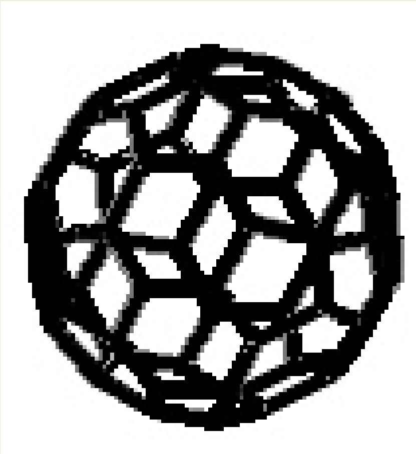


Giant covalent structures: e.g. Buckminsterfullerene; “bucky ball”

- During the few decades new forms of carbon have been discovered some of which have “closed cage” arrangements of the atoms.

- These are large but are not really giant molecules. The are **SIMPLE COVALENT MOLECULES**

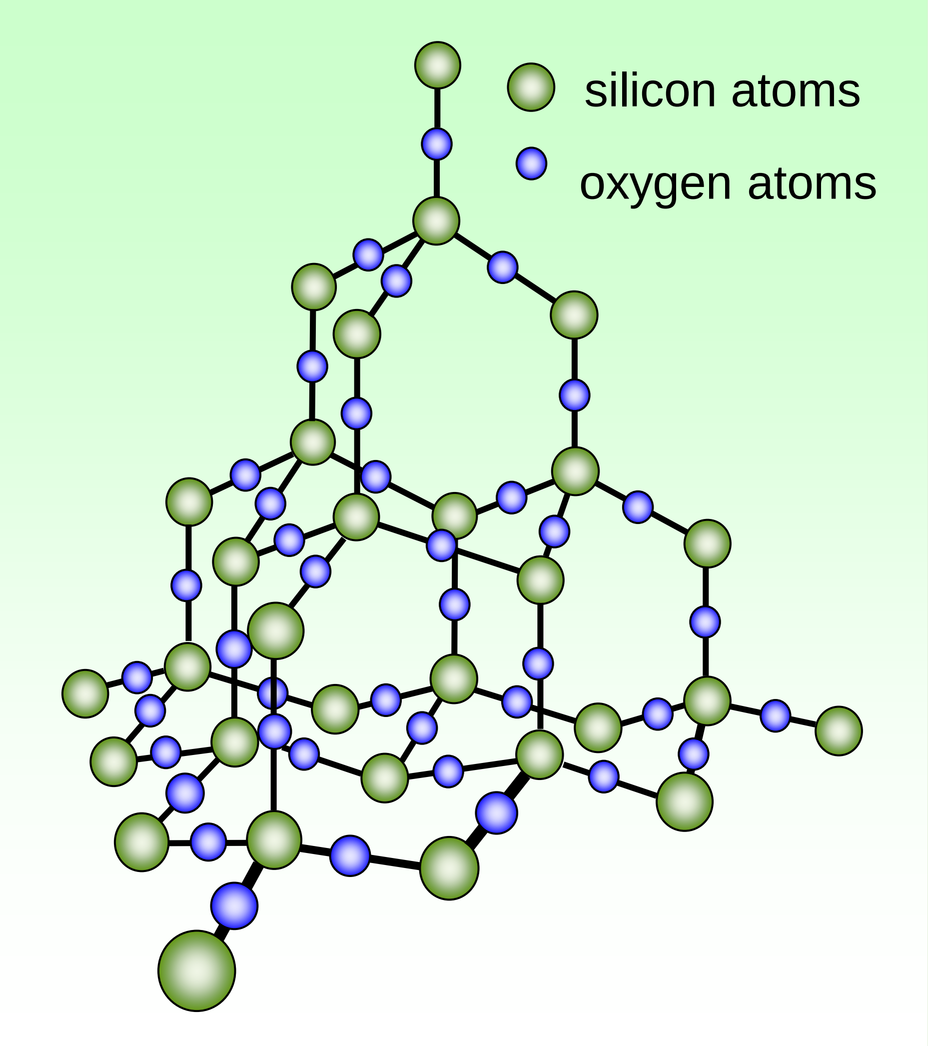
One of them contains 60 carbon atoms and bears remarkable similarities to a football!



Giant covalent structures: sand

- Sand is an impure form of silicon dioxide.

- Although it is a compound, it has a giant covalent structure with certain similarities to diamond.



BONDING AND PHYSICAL PROPERTIES

Bonding and physical properties

The type of structure that substances have has a huge effect upon physical properties.

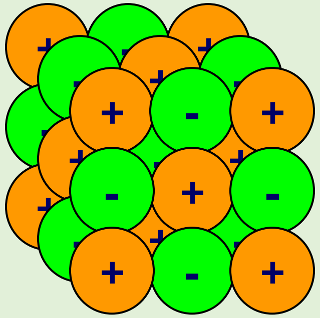
These are things such as:

- Density
- Conductivity
- Malleability/ brittleness
- Melting point

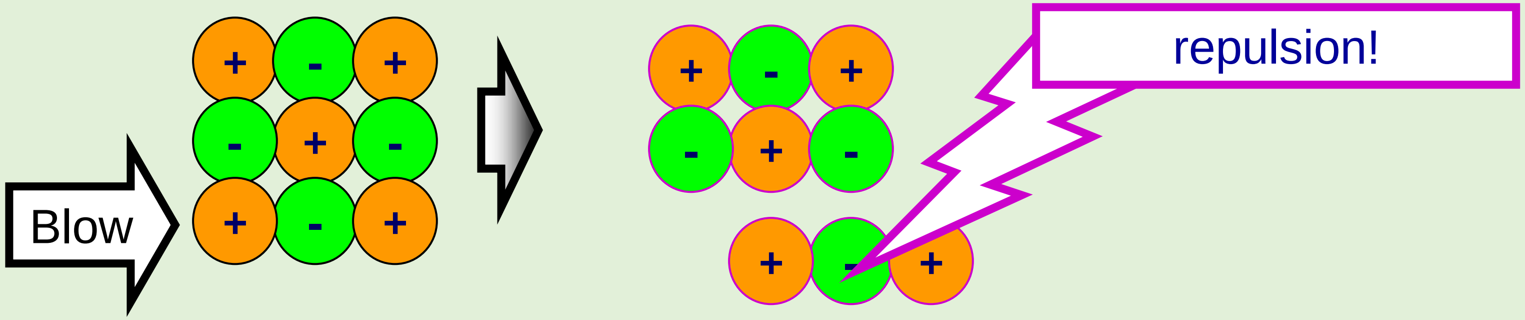
The next few slides illustrate just a few of the general patterns.

Bonding and physical properties

- Ionic compounds are very brittle.
- Opposite charges attract, so neighbouring ions are pulled together.
- When something hits the substance a layer of ions will be pushed so that they are next to ions with the same charge.

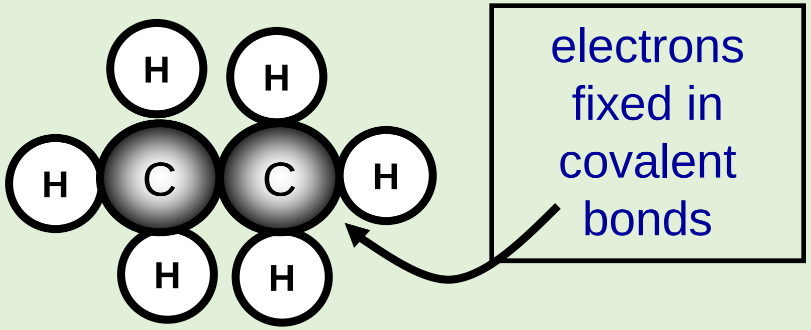


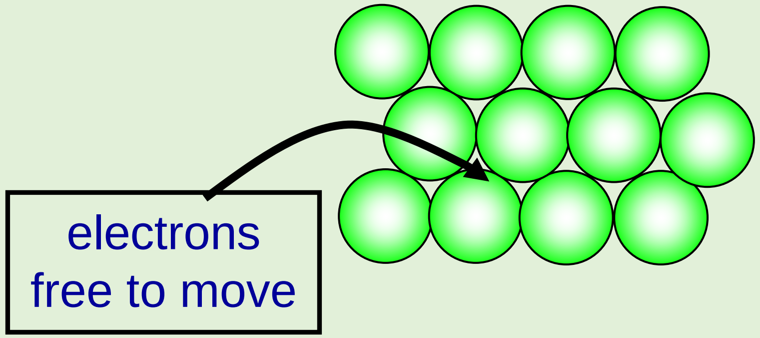
Attraction becomes:



Bonding and physical properties

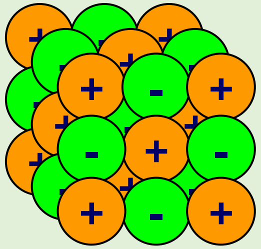
- Covalent substances do not conduct electricity.
- This is because in covalent substances the outer electrons are fixed (localised) between specific atoms.
- Metals conduct electricity.
- In metals the electrons can, given a potential, move anywhere throughout the structure.



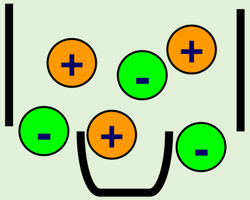


Bonding and physical properties

- Ionic substances do not conduct electricity as solids.
- When molten or dissolved they will conduct (and also undergo electrolysis).
- This is because the electricity is carried through the solution by the ions which are free to move when the ionic compound is molten or in solution.



Solid – not free to move
Doesn't conduct



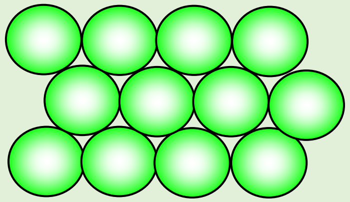
Molten – mobile
Does conduct

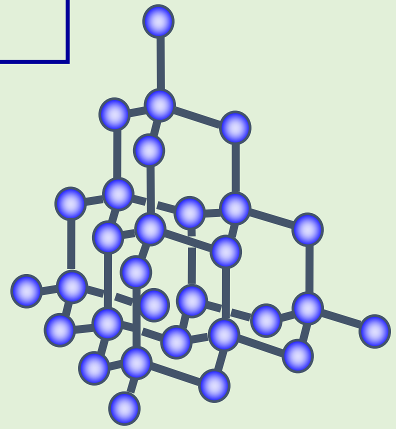
Bonding and physical properties

- Generally substances with giant structures have high melting points and boiling points.

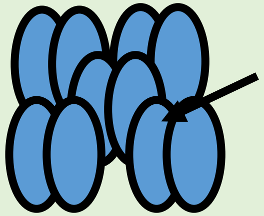
- Small molecules have melting points and boiling points that increase as the size of the molecule increases.

In giant structures all the atoms are tightly bonded together. Usually they are high melting-point solids.

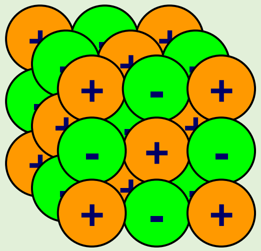




Small molecules tend to be gas, liquid solids with low melting points.



weak forces between molecules

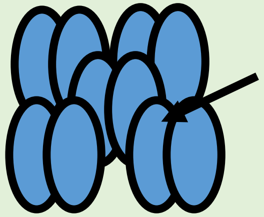


Bonding and physical properties

- Generally substances with giant structures do not dissolve easily (although many ionic compounds dissolve in water for a special reason).

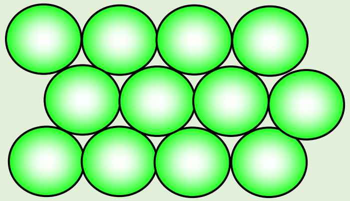
- Again this is because in giant structures separating the particles involves breaking chemical bonds.

Small molecules usually dissolve in a range of solvents. We just separate one molecule from another.

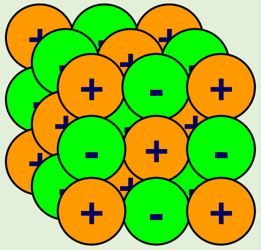


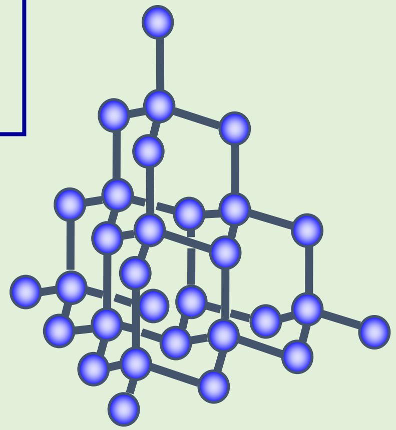
weak forces between molecules

Giant structures generally don't dissolve easily.



strong bonds between the atoms/ions





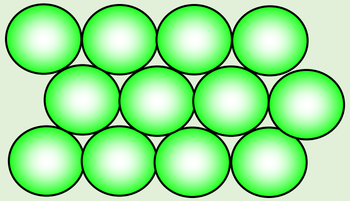
Bonding and physical properties

- The density of substances depends upon how closely the atoms are packed together.

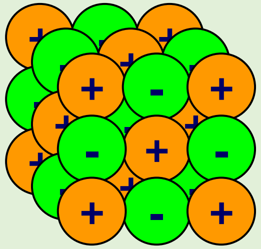
- Giant structures, metals especially, tend to be dense because all atoms/ions are pulled tightly together.

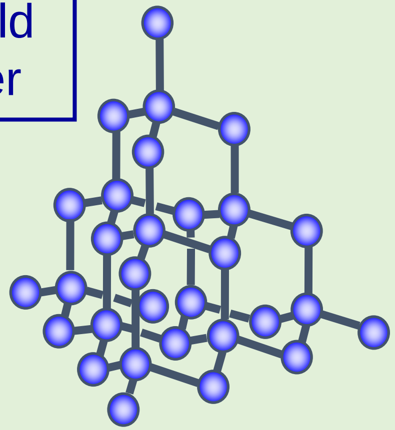
- Small molecules often have lower densities.

Giant structures generally have high densities.

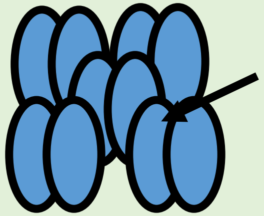


atoms / ions held closely together





Small molecules tend to have low densities because of space wasted between the molecules.



weak forces between molecules

What type of bonding will the substances have?

Substance	Bonding
Brass (Alloy copper + zinc)	Metallic
Copper oxide	Ionic
Sulphur dioxide	Covalent
Iron	Metallic
Sodium fluoride	Ionic
Nitrogen chloride	Covalent

04 January 2022

Polymer molecules

Polymers

- **<u>Polymer</u>**s comes from Greek –

  "**Poly = many**", "**mers = parts**".
- Polymers are built up from **<u>monomer</u>** units.
- Monomers undergo a **<u>polymerisation</u>** reaction to make a polymer.

"A polymer is…

a very large molecule made from many repeating units / monomers".

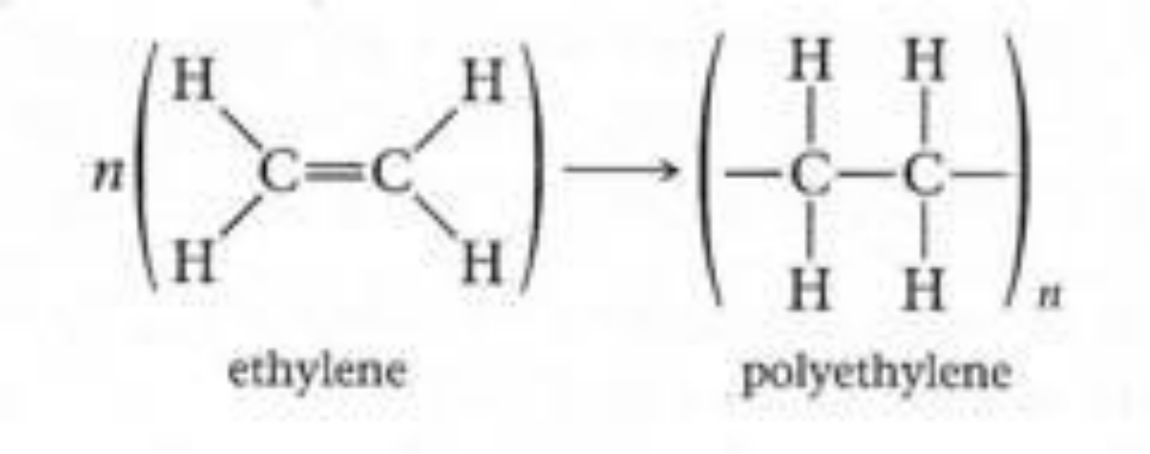
Addition Polymers

<u>Addition</u> polymers –

formed when **<u>alkenes</u>**

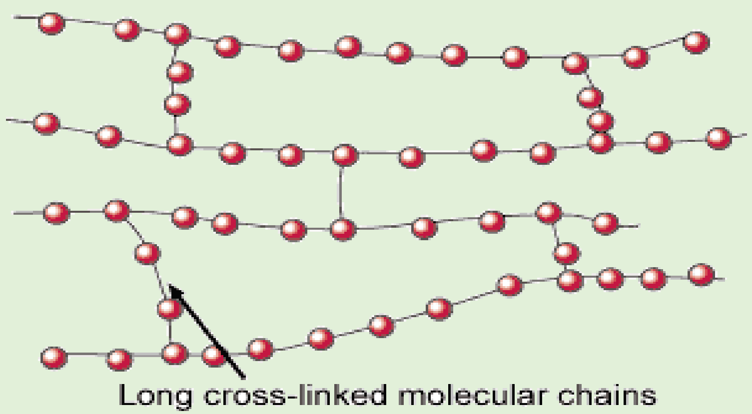
polymerise

e.g. poly(ethene).



Thermosetting plastics

- Once set they will not melt again.
- They burn before they melt.
- Rubber.
- Bakelite.



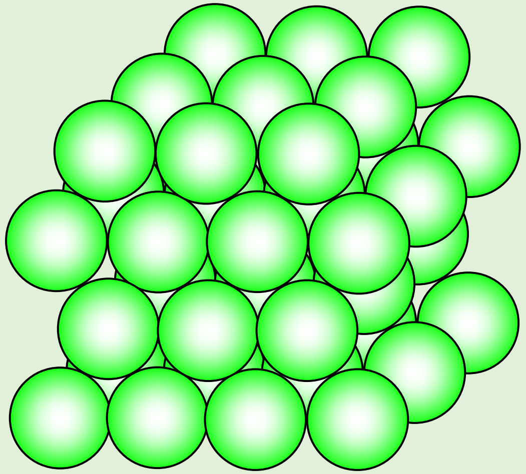
Thermosoftening plastics

- No cross links.
- Can be melted and softens and can be re-formed.
- Polyethene.

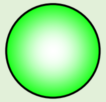
Structure of metals

Metallic bonding

- Metal atoms form a giant lattice similar to ionic compounds.
- The outermost electrons on each metal are free to move throughout the structure and form a “sea of electrons”.
- Having released electrons into this “sea” the metal atoms are left with a + charge.



Metallic bonding is the attraction of + metal ions for the “sea of delocalised electrons.”



= positively charged metal ion

Metallic Bonding

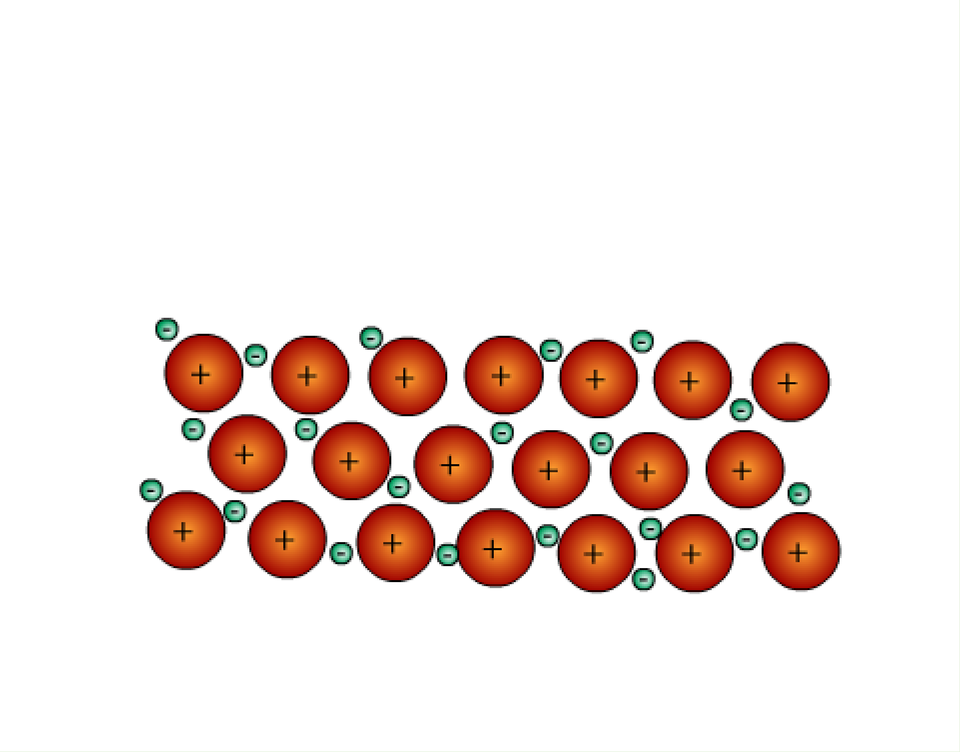
Metal properties

Giant Metallic structure

- Outer electrons are free to move around the metal

- The positive ions left are attracted to the free (delocalized)electrons

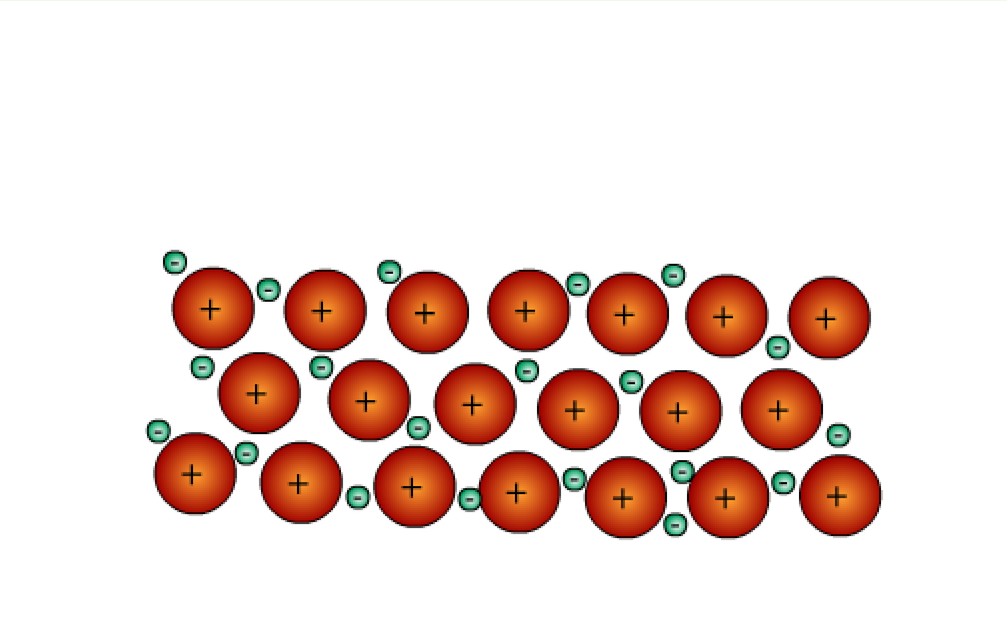
- This attraction holds the giant metal structure together



Conducting electricity

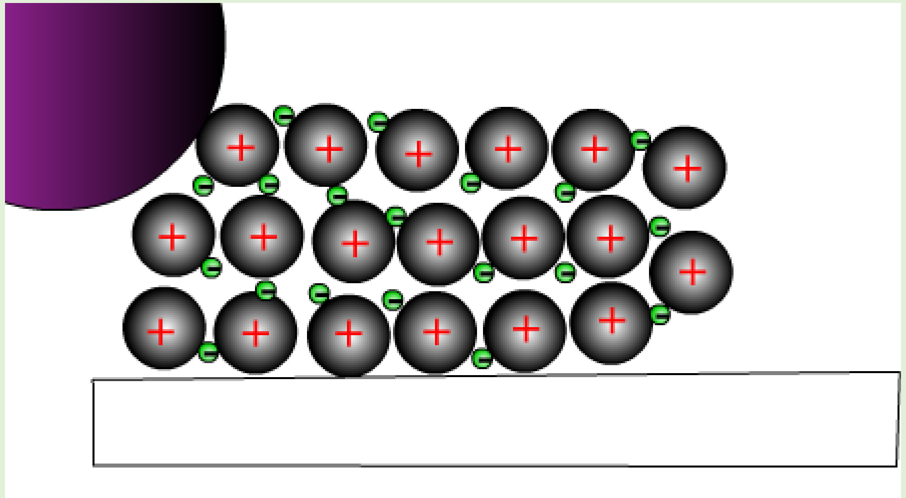
- Applying a voltage makes the delocalized electrons flow in one direction (a current).

- If a battery replaces the electrons then the current continues to flow.

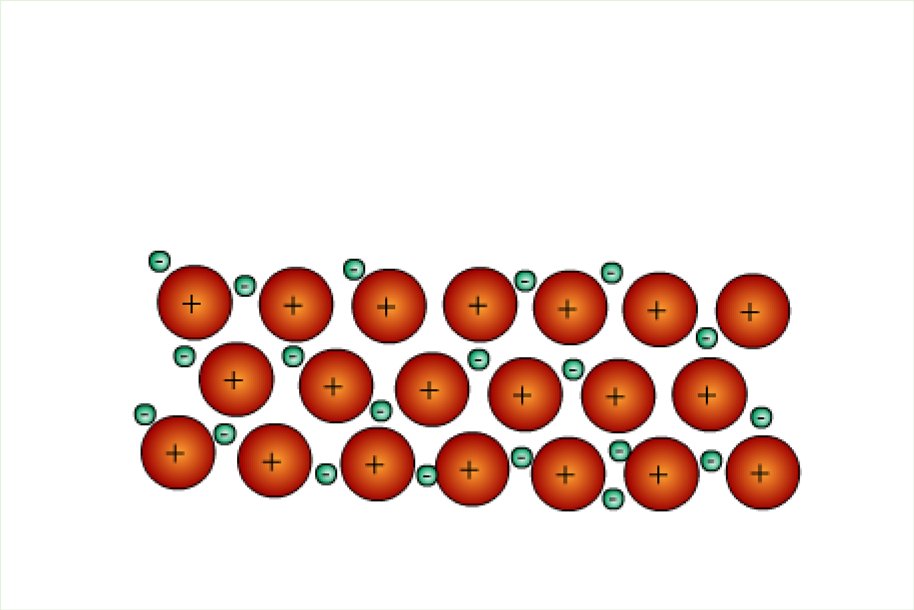


Ductile (making into wires and sheets)

- Delocalised electrons allow the ions to move past each other.

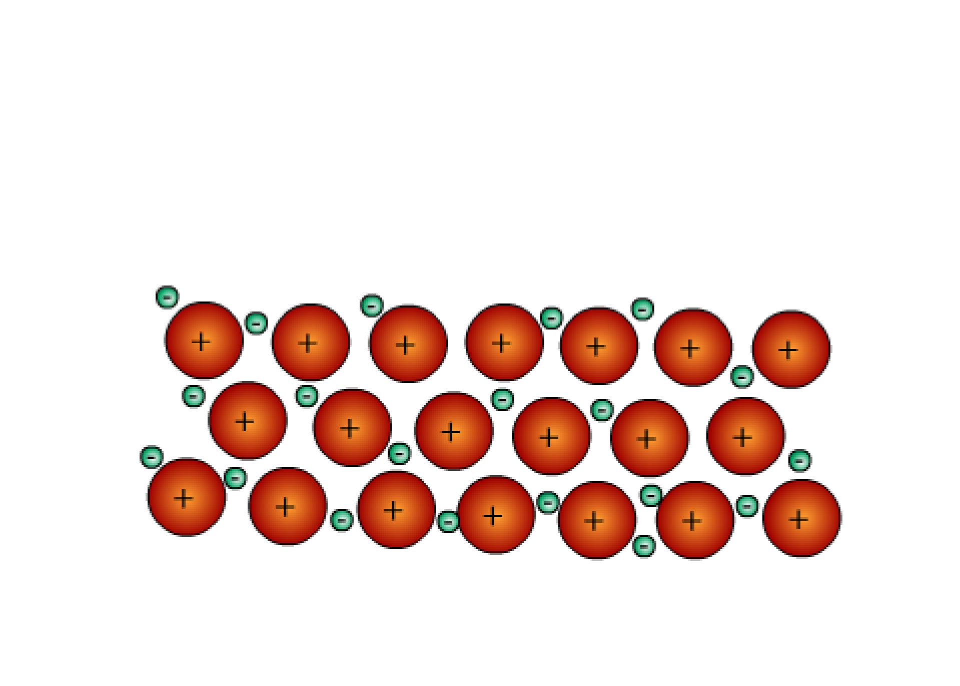


- The ions take up new positions while still being attracted to the rest of the metal.



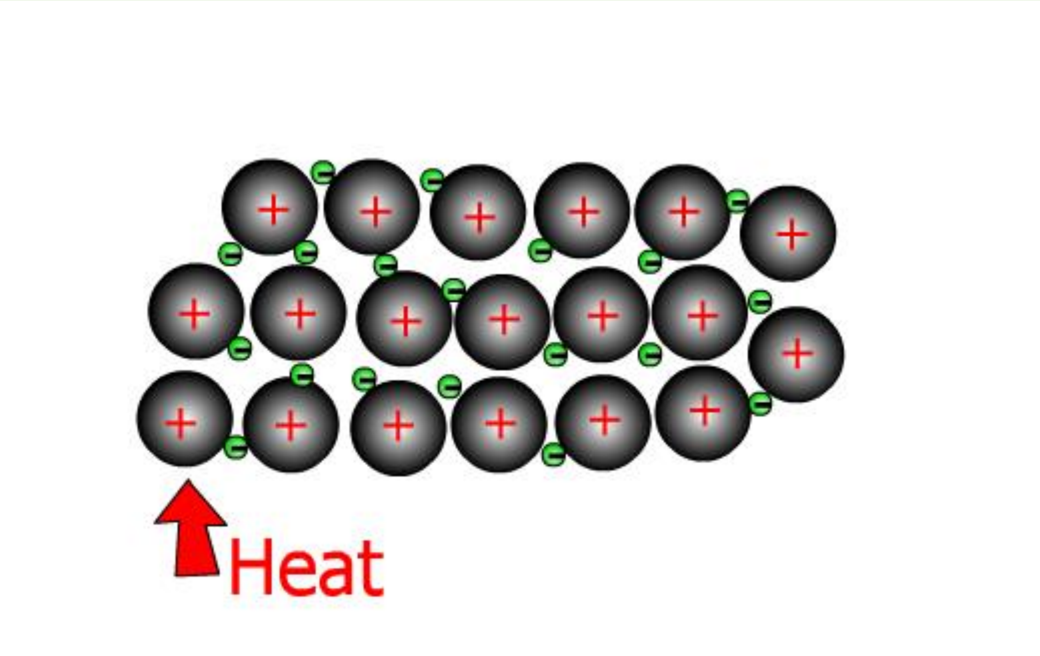
Malleable (bend into shape)

- The delocalised electrons hold the metal together while it is pushed into new shapes.



Conducting Heat

Why do you think metals are good at conducting heat?



Atomic structure and the Periodic Table

Rules to remember

The ***smaller number*** on an element's box is the ***atomic number*** – these increase one by one as you go up the Periodic Table. It is the ***number of protons*** in the nucleus of one atom of that element.

The ***larger number*** is the ***atomic mass*** - this is calculated by adding up how many ***protons and neutrons*** there are in the nucleus of each atom of that element.

In a standard ***atom*** of an element, the ***protons and electrons have equal numbers***.

So…

Which number is the atomic number?

Which number is the atomic mass?

How many protons?

How many neutrons?

How many electrons?

Atomic mass = 24 = protons + neutrons

12 protons + ? neutrons = 24

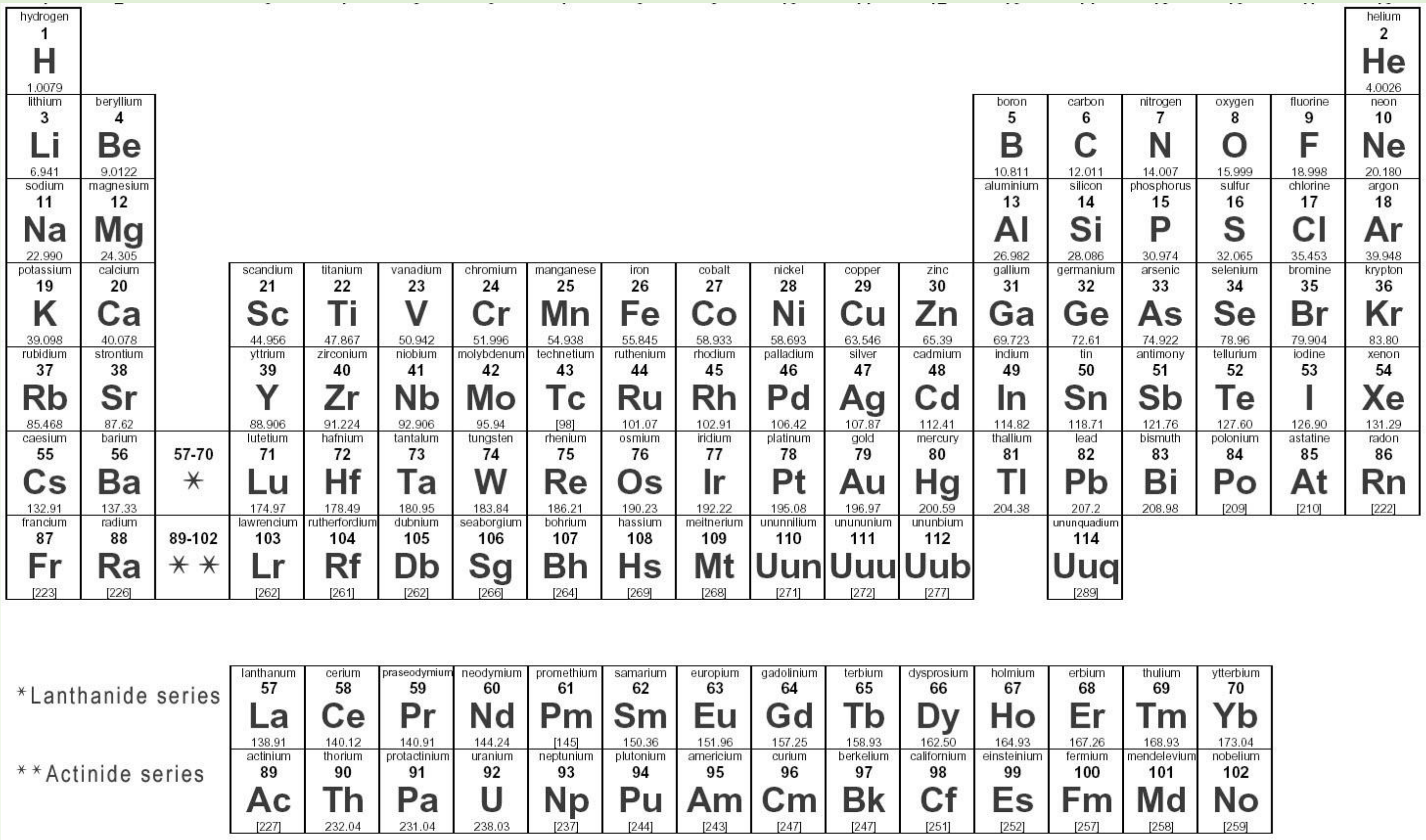
24 – 12 = neutrons

12 neutrons

Atomic number = 12 protons

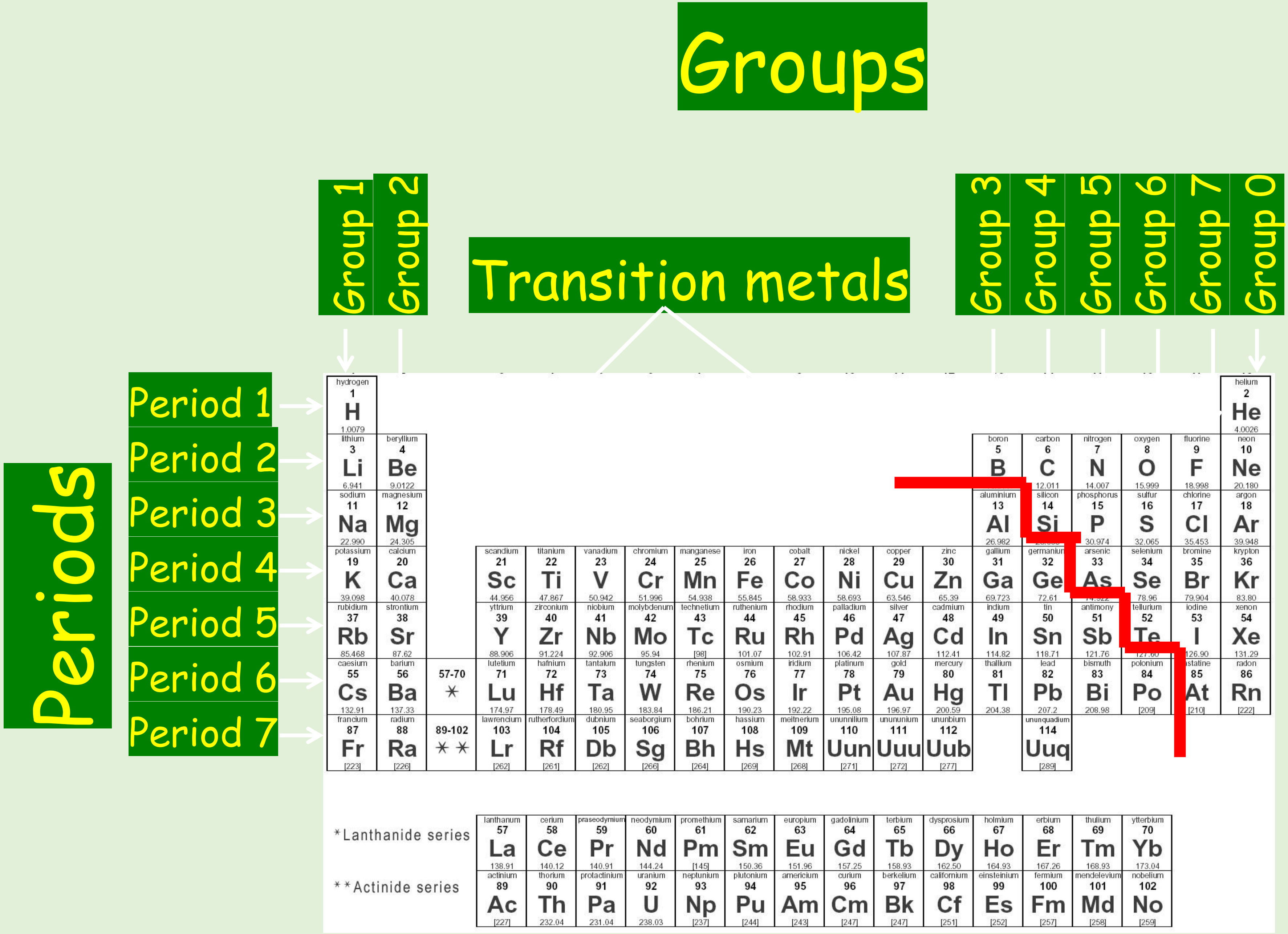
Protons = electrons → 12 electrons

Rules – ***Groups*** are ***vertical columns*** but only the taller ones (not the transition metals in the middle) they are numbered from 1 on the left to 8 (or correctly 0) on the right



More rules – ***Periods*** are ***horizontal rows*** starting with 1 for the top row down to period 7 at the bottom

Annotate your Periodic Table with period and group numbers



Groups
Group 1
Group 2
Transition metals
Group 3
Group 4
Group 5
Group 6
Group 7
Group 0
Periods
Period 1
Period 2
Period 3
Period 4
Period 5
Period 6
Period 7
*Lanthanide series
**Actinide series

Unit 3

Formula of Elements and Molecules

What is a word equation?

During a **CHEMICAL** reaction **REACTANTS** form **PRODUCTS**

reactants → products

A **word equation** uses the names of the reactants and products to show what happens in a chemical reaction.

For example, when a piece of sulfur is burned in oxygen gas it produces a colourless gas called sulfur dioxide.

The word equation for this reaction is:

sulfur + oxygen → sulfur dioxide

What is a symbol equation?

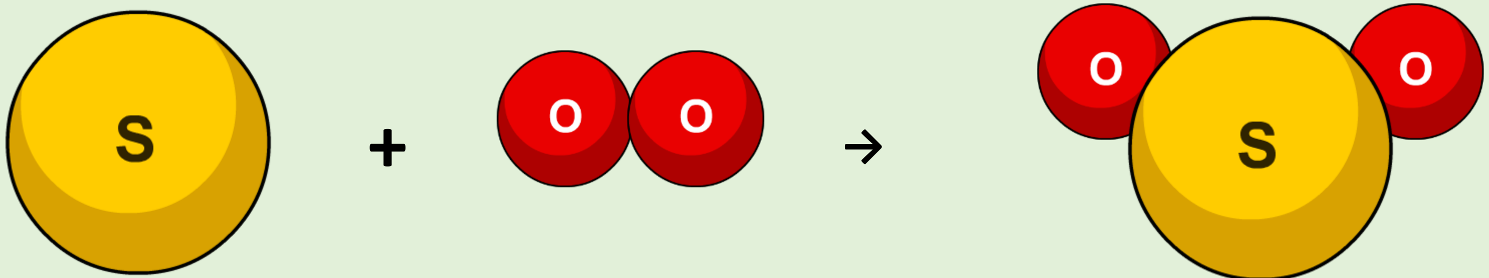
A **symbol equation** uses the formulae of the reactants and products to show what happens in a chemical reaction.

$$S + O_2 \rightarrow SO_2$$

This equation shows that one **atom** of sulfur (**S**) reacts with one **molecule** of oxygen (**O_2**) to make one **molecule** of sulfur dioxide (**SO_2**).

A *Molecular Formula* shows the number of each type of atom in a molecule.

A *Diatomic Molecule* contains two atoms; e.g. O_2



Interpreting Molecular Formulae

How many elements are in this compound, and what are they?

- NaCl = two elements: sodium and chlorine

How many atoms are in this formula, and what are they?

- Fe_2O_3 = iron (Fe) two atoms
  oxygen (O) three atoms

How many atoms are in this formula, and what are they?

- $Ca(OH)_2$ = calcium (Ca) one atom
  oxygen (O) two atoms
  hydrogen (H) two atoms

What do state symbols show?

State symbols are added to a symbol equation to show whether the reactants and products are:

- **solid** – symbol is **(s)**
- **liquid** – symbol is **(l)**
- **gas** – symbol is **(g)**
- **dissolved in water (aqueous)–** symbol is **(aq)**.

$$\mathbf{S}(s) \quad + \quad \mathbf{O_2}(g) \rightarrow \quad \mathbf{SO_2}(g)$$

With state symbols in place, this symbol equation now shows that the sulfur is a solid, the oxygen is a gas and the sulfur dioxide is also a gas.

Ionic formulae and equations

REMEMBER:Writing word equations

- In any chemical reaction we combine reactants to make products

- Reactants $\longrightarrow$ Products
- A + B $\longrightarrow$ C + D

- Example - Methane gas burns in oxygen to produce carbon dioxide and water.
- Methane + Oxygen $\longrightarrow$ Carbon dioxide + water

Word equation practise – Have a go at these!

- 1) Lead nitrate reacts with potassium iodide. This produces a yellow solid called lead iodide and a clear liquid called potassium nitrate

- 2) In a car, when the engine is started enough energy is produced to combine nitrogen from the air with oxygen to form nitrogen dioxide.

- 3) To make sodium chloride one possible method is to react hydrochloric acid and sodium hydroxide together. This reaction also produces water

Introduction

- The periodic table can help us assign charges to elements in groups 1-3 (positive) and 5-7 (negative)

1+	2+											3+		3-	2-	1-	
					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 sulphur 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							

key: atomic mass / symbol / name / atomic number

Elements with atomic numbers 112–116 have been reported but not fully authenticated

* The Lanthanides (atomic numbers 58–71) and the Actinides (atomic numbers 90–103) have been omitted
Cu and Cl have not been rounded to the nearest whole number

Naming ionic compounds

- Almost all ionic compounds have a two word name

Cation	Anion	Ionic compound
Sodium	Chloride	Sodium chloride
Lithium	Bromide	Lithium bromide

The ending of the anion is usually the part that changes, mostly to –ide. This tells us they contain only atoms from one element

Note : Cations have a positive charge and Anions have a negative charge

Writing chemical formula

Cations			Anions	
+1	**+2**	**+3**	**-1**	**-2**
Lithium - Li^{+}	Magnesium - Mg^{2+}	Aluminium - Al^{3+}	Chloride - Cl^{-}	Oxide - O^{2-}
Sodium - Na^{+}	Calcium - Ca^{2+}		Bromide - Br^{-}	Sulphide - S^{2-}
Potassium - K^{+}	Barium - Ba^{2+}		Iodide - I^{-}	Sulfite - SO_3^{2-}
Ammonium - NH_4	Iron (II) - Fe^{2+}		Nitrate - NO_3^{-}	Sulphate - SO_4^{2-}
Hydrogen - H^{+}	Copper - Cu^{2+}		Hydroxide - OH^{-}	Carbonate - CO_3^{2-}

Ionic compounds in symbols are called chemical formula, these save time and give us the correct ratio of elements present. We need the charge of both anion and cation to write them

Writing chemical formula

- Step 1: write down the name
- Step 2 Write down the correct formula for each ion
- Step 3 balance the charges so that the compound is neutral

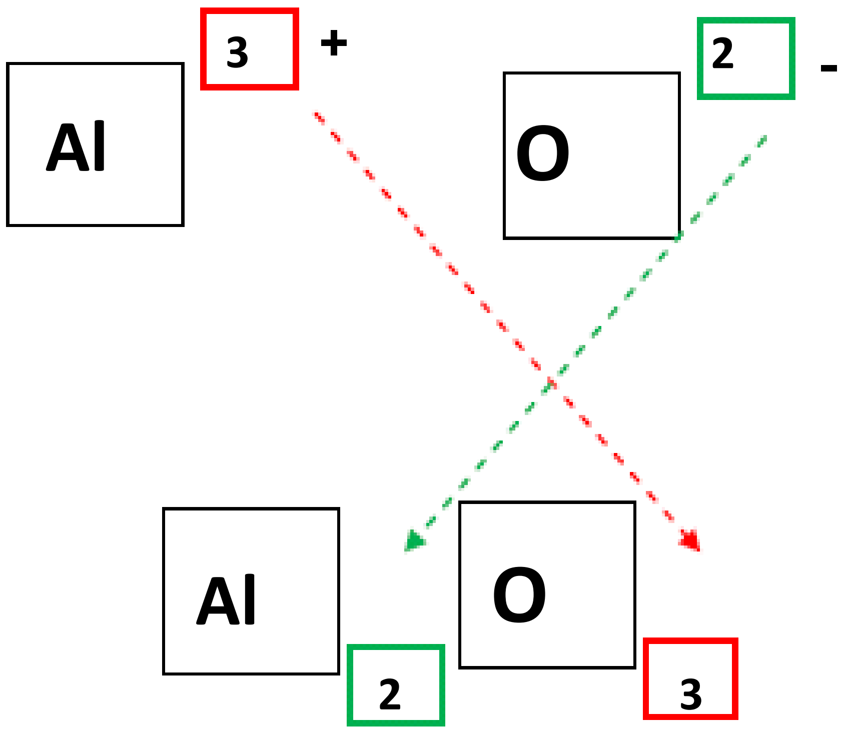
- Example
- Step 1: sodium sulphate
- Step 2: Na^+ and SO_4^{2-}
- Step 3: Na^+ Na^+ SO_4^{2-} (Notice we need 2 sodium to balance)
- Step 4: Na_2 SO_4

Practice writing ionic formula

- You will need your table of common ions to do this.

- Copper Chloride
- Iron (II) Oxide
- Magnesium carbonate
- Barium Iodide
- Ammonium nitrate
- Calcium carbonate

Alternative method – cross over



1. Write the symbols for each ion in the large boxes and their charges in the small boxes.
2. Write the symbols of the ions in the large boxes (these cannot change)
3. Swap the numbers on the charges over as shown by the arrows and colours.
4. Write the formula without the boxes, using brackets if needed.

Chemical Equations

Word equation practise – Have a go at these!

- 1)Carbon burns on oxygen (from air!) to produce carbon dioxide
- 2)During respiration, glucose reacts with oxygen to produce carbon dioxide and water
- 3)During photosynthesis, plants make their own food (glucose) by reacting carbon dioxide and water together to produce glucose plus oxygen.

Balancing equations

- You can never change a formula when balancing an equation. You can only put a number in front of a formula.
- Lets consider how to show 2 molecules of CO_2:
  - C_2O ✗
  - CO_4 ✗
  - C_2O_2 ✗
  - $2CO_2$ ✓

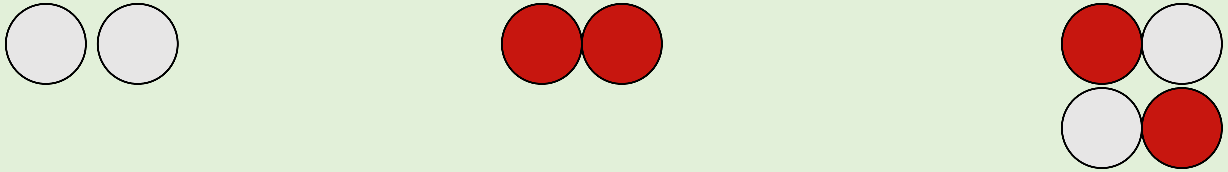
Balancing equations

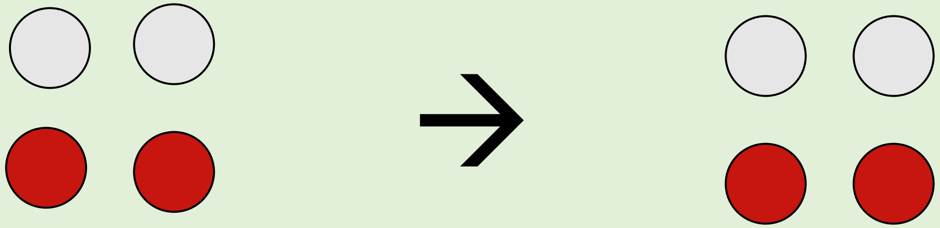
- No atoms are created or destroyed in a chemical reaction.
- There should be the same number of atoms of each type of element each side of the equation.

  - Magnesium + Oxygen $\longrightarrow$ Magnesium oxide

Balancing equations

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





Symbols $Mg + O_2$

You try – balance these equations

1. $C + O_2 \longrightarrow CO_2$

2. $H_2 + O_2 \longrightarrow H_2O$

3. $N_2 + H_2 \longrightarrow NH_3$

4. $CH_4 + O_2 \longrightarrow CO_2 + H_2O$

Half Equations and Ionic Equations

HT only

Types of equations you have met in Chemistry include:

- word equations

e.g.

aluminium + iron(III) oxide $\longrightarrow$ aluminium oxide + iron

- chemical (or balanced) equations

e.g.

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

Others you will meet are:

- ionic equations
- half equations

Half Equations

- A half equation is a model for the change that happens to **ONE REACTANT** in an equation equations showing the SEPARATE oxidation (loss of e^-) and reduction (gain of e^-) processes in any redox reaction

Eg 1 $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

Ca atoms - $0 \rightarrow +2$ $\rightarrow$ Ca oxidised (lose electrons)

O_2 mols - $0 \rightarrow -2$ $\rightarrow$ O_2 reduced (gain electrons)

HALF EQUATIONS :

Oxidation : $Ca \rightarrow Ca^{2+} + 2e^-$

Reduction : $O_2 + 4e^- \rightarrow 2O^{2-}$

Eg 2 $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Na atoms - $0 \rightarrow +1$ $\rightarrow$ Na oxidised

H_2O mols - $H(+1) \rightarrow H(0)$ $\rightarrow$ H_2O reduced

HALF EQUATIONS :

Oxidation : $Na \rightarrow Na^+ + e^-$

Reduction : $2H_2O + 2e^- \rightarrow 2OH^- + H_2$

Ionic equations have been developed to remove chemicals (IONS) that might otherwise over complicate the chemistry of the reaction.

Which ions need to be removed?

These are known as SPECTATOR IONS

Deciding on the ions to be removed can be quite difficult.

A knowledge of the chemicals (bonding) and their states of matter is essential.

Knowledge of the states of matter is essential as this decides whether these particles are separated out.

There are FOUR states of matter.

- SOLID
- LIQUID
- GAS
- & AQUEOUS

The spectator ions must be able to move at room temp – so they must exist in the aqueous state.

Ions cannot exist as gases or liquids at room temperature. This means that if a solid or gas is present it cannot contain spectator ions and should NOT be removed

Look at the following chemicals and decide if they could contain spectator ions or not. Give a reason.

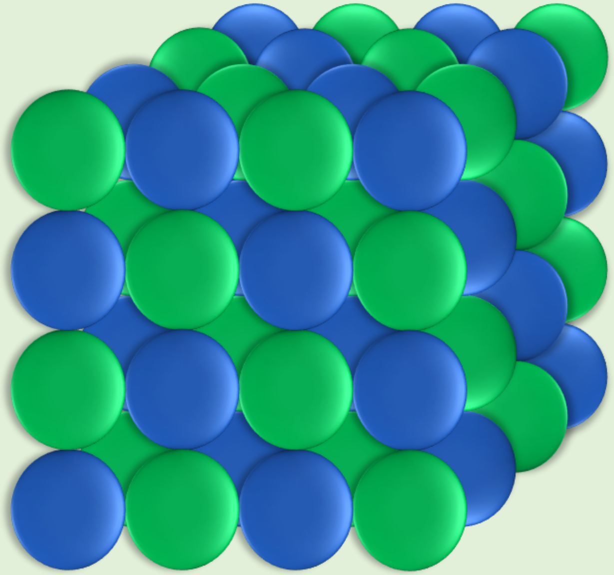
Sodium	NONE	metal element
Chlorine	NONE	non-metal element
Carbon dioxide	NONE	covalent compound
Water	NONE	covalent compound
Solid sodium chloride	NONE	solid ionic compound
Molten sodium chloride	NONE	molten ionic compound
Aqueous sodium chloride	YES	aqueous ionic compound
Aqueous glucose	NONE	aqueous covalent cmpd
Copper(II) sulfate solution	YES	aqueous ionic compound

Dissociation of compounds into ions

Dissociation is a term used to state that particles in a substance 'spread out' (when heated or dissolved)

Particles in a solid:

- just vibrate
- are close together
- have a regular arrangement (LATTICE)



You always need to include a STATE symbol!

i.e. For sodium chloride

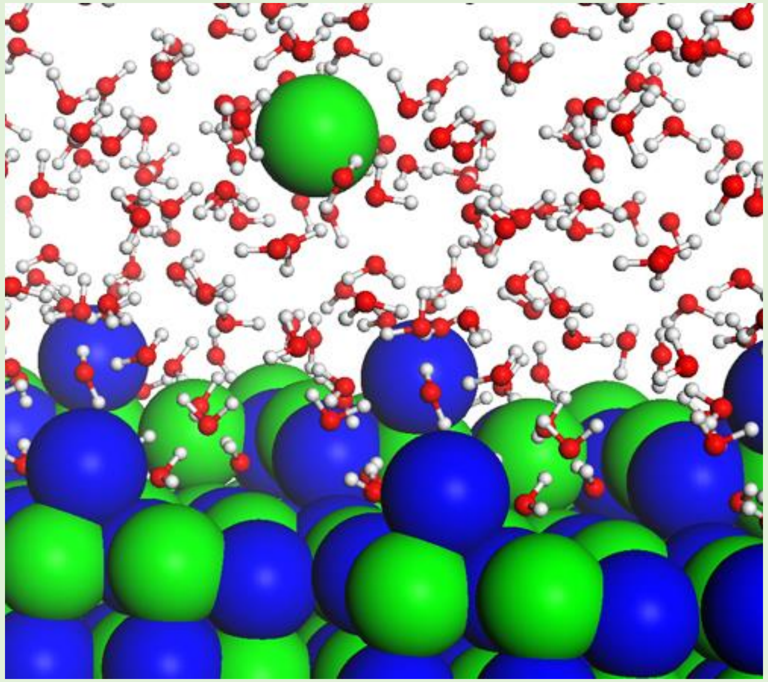
$NaCl_{(s)}$

Dissociation of compounds into ions

When the solid is dissolved in water

an aqueous solution is formed

The IONS spread out (or DISSOCIATE)



So for sodium chloride dissolving in water the word equation is:

sodium chloride(s) + aq

↓

sodium ions(aq) + chloride ions(aq)

The chemical equation is:

$$NaCl(s) + aq \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$$

Chemical equations showing dissociation of into ions

The word equation will tell you which ions are produced:

sodium chloride(s) + aq → sodium ions(aq) + chloride ions(aq)

The resulting chemical equation is:

$$NaCl(s) + aq \longrightarrow Na^{+}(aq) + Cl^{-}(aq)$$

If the formula has more than one particular ion, balancing is needed in the chemical equation.

Again the word equation will tell you which ions are produced:

barium chloride(s) + aq → barium ions(aq) + chloride ions(aq)

However, the resulting chemical equation is:

$$BaCl_2(s) + aq \longrightarrow Ba^{2+}(aq) + 2Cl^{-}(aq)$$

Notice the number in the formula now comes before the ion

Chemical equations showing dissociation of into ions

If the formula has complicated ion

e.g ammonium, nitrate, hydroxide etc.

this ion is NOT split into the individual elements

The word equation will still tell you which ions are produced:

sodium hydroxide(s) + aq → sodium ions(aq) + hydroxide ions(aq)

The resulting chemical equation is:

$$NaOH\,(s) + aq \longrightarrow Na^{+}(aq) + OH^{-}(aq)$$

Notice the OH^{-} ion is not split into the individual elements

With a more complicated compound start with the word equation

ammonium sulfate(s) + aq → ammonium ions(aq) + sulfate ions(aq)

The resulting chemical equation is:

$$(NH_4)_2SO_4\,(s) + aq \longrightarrow 2NH_4^{+}(aq) + SO_4^{2-}(aq)$$

Chemical equations showing dissociation of into ions

Acids, even though they consist of non-metals only
do dissociate into ions when dissolved in water

The word equation can help

sulfuric acid(l) + aq ⟶ hydrogen ions(aq) + sulfate ions(aq)

The resulting chemical equation is:

$$H_2SO_4\,(l) + aq \longrightarrow 2H^+\,(aq) + SO_4^{2-}(aq)$$

Notice the sulfate ion is not split into the individual elements

Try dissociating the following acids:

nitric acid(l) + aq ⟶ hydrogen ions (aq) + nitrate ions(aq)

The resulting chemical equation is:

$$HNO_3\,(l) + aq \longrightarrow H^+(aq) + NO_3^-\,(aq)$$

Chemical equations showing dissociation of into ions

Look at the following substances and dissociate them into ions completing the word equation and writing a chemical equation

magnesium bromide(s) + aq ⟶ magnesium ions(aq) + chloride ions(aq)

$$MgBr_2 + aq \longrightarrow Mg^{2+}(aq) + 2Br^-(aq)$$

calcium hydroxide(s) + aq ⟶ calcium ions(aq) + hydroxide ions(aq)

$$Ca(OH)_2 + aq \longrightarrow Ca^{2+}(aq) + 2OH^-(aq)$$

aluminium nitrate(s) + aq ⟶ aluminium ions(aq) + nitrate ions (aq)

$$Al(NO_3)_3 + aq \longrightarrow Al^{3+}(aq) + 3NO_3^-(aq)$$

sodium carbonate(s) + aq ⟶ sodium ions(aq) + carbonate ions(aq)

$$Na_2CO_3 + aq \longrightarrow 2Na^+(aq) + CO_3^{2-}(aq)$$

phosphoric acid(s) + aq ⟶ hydrogen ions(aq) + phosphate ions(aq)

$$H_3PO_4 + aq \longrightarrow 3H^+(aq) + PO_4^{3-}(aq)$$

Producing Ionic Equations for Neutralisation Reactions

Neutralisation reactions can be thought of as

- removal of acidity, by reaction of acid with base (alkali)
- reaction of acid to produce water as one of the products

There are three common strong acids.

Hydrochloric acid	HCl
Nitric acid	HNO_3
Sulfuric acid	H_2SO_4

When added to water *ALL* acids produce $H^+(aq)$

Three common strong alkalis are:

Sodium hydroxide	$NaOH$
Potassium hydroxide	KOH
Calcium hydroxide	$Ca(OH)_2$

When dissolved in water *ALL* alkalis produce $OH^-(aq)$

Producing Ionic Equations for Neutralisation Reactions

Recognising Neutralisation reactions

A neutralisation reaction will involve:

- an ACID (reactant)
- WATER (product)
- NO elements

An example of a neutralisation reaction is:

sodium hydroxide(aq) + nitric acid(aq) $\longrightarrow$ sodium nitrate(aq) + water(l)

Notice: WATER is a liquid (l) NOT aqueous

Aqueous (aq) is present in all other chemicals

Spectator ions MUST exist in the (aq) state

In this case there are TWO: Na^{+} (aq) & NO_3^{-} (aq)

The ionic word equation must be:

hydroxide ions(aq) + hydrogen ions(aq) $\longrightarrow$ water(l)

The ionic equation for neutralisation is:

$$H^{+}(aq) + OH^{-}(aq) \longrightarrow H_2O(l)$$

Producing Ionic Equations for Neutralisation Reactions

Another example of a neutralisation reaction is:

calcium hydroxide(aq) + hydrochloric acid(aq) $\longrightarrow$ calcium chloride(aq) + water(l)

The spectator ions are: Ca^{2+}(aq) & Cl^{-} (aq)

The ionic word equation must be:

hydroxide ions(aq) + hydrogen ions(aq) $\longrightarrow$ water(l)

The ionic equation for neutralisation is:

$$H^{+}(aq) + OH^{-}(aq) \longrightarrow H_2O(l)$$

Notice it is the same as the previous example.

Producing Ionic Equations for Neutralisation Reactions

An example of a neutralisation reaction using carbonates is:

Sodium carbonate(aq) + hydrochloric acid(aq) ⟶
sodium chloride(aq) + water(l) + carbon dioxide(g)

The spectator ions are: Na^+ (aq) & Cl^- (aq)

The ionic word equation must be:

carbonate ions(aq) + hydrogen ions(aq) ⟶ water(l) + carbon dioxide(g)

The ionic equation for this reaction is:

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

NOTICE the need to balance this equation.

Producing Ionic Equations for Neutralisation Reactions

Try producing ionic equations for the following reactions

potassium hydroxide(aq) + sulfuric acid(aq) ⟶ potassium sulfate(aq) + water(l)

The spectator ions are: K^+ (aq) & SO_4^{2-} (aq)

The ionic word equation must be:

hydroxide ions(aq) + hydrogen ions(aq) ⟶ water(l)

The ionic equation for neutralisation is:

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

caesium carbonate(aq) + nitric acid(aq) ⟶
caesium nitrate(aq) + water(l) + carbon dioxide(g)

The spectator ions are: Cs^+ (aq) & NO_3^- (aq)

The ionic word equation must be:

carbonate ions(aq) + hydrogen ions(aq) ⟶ water(l) + carbon dioxide(g)

The ionic equation for this reaction is:

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

Producing Ionic Equations for Precipitation Reactions

Recognising PRECIPITATION reactions

A precipitation reaction will involve:

- TWO SOLUTIONS (reactants)
- A SOLID product (precipitate)
- NO elements

An example of a precipitation reaction is:

lead nitrate(aq) + sodium iodide(aq) $\longrightarrow$ sodium nitrate(aq) + lead iodide(s)

Notice: There is only ONE solid (s) product

Aqueous (aq) is present in all other chemicals

Spectator ions MUST exist in the (aq) state

In this case there are TWO: Na^{+} (aq) & NO_3^{-} (aq)

The ionic word equation must be:

lead ions(aq) + iodide ions(aq) $\longrightarrow$ lead iodide(s)

The ionic equation for precipitation is:

$$Pb^{2+}(aq) + 2I^{-}(aq) \longrightarrow PbI_2(s)$$

Producing Ionic Equations for Precipitation Reactions

precipitation reactions may involve acids:

An example of such a precipitation reaction is:

barium chloride(aq) + sulfuric acid(aq) $\longrightarrow$
hydrochloric acid(aq) + barium sulfate(s)

Spectator ions MUST exist in the (aq) state

In this case there are TWO: H^+(aq) & Cl^-(aq)

The ionic word equation must be:

barium ions(aq) + sulfate ions(aq) $\longrightarrow$ barium sulfate(s)

The ionic equation for precipitation is:

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

Notice: the hydrogen ion is a SPECTATOR ion

Exothermic Reactions

Exothermic reactions increase in temperature.

- Examples include:
  - Burning reactions including the combustion of fuels.
  - Detonation of explosives.
  - Reaction of acids with metals.

Magnesium reacting with acid

Thermit reaction

Activity

Say whether these processes are exothermic.

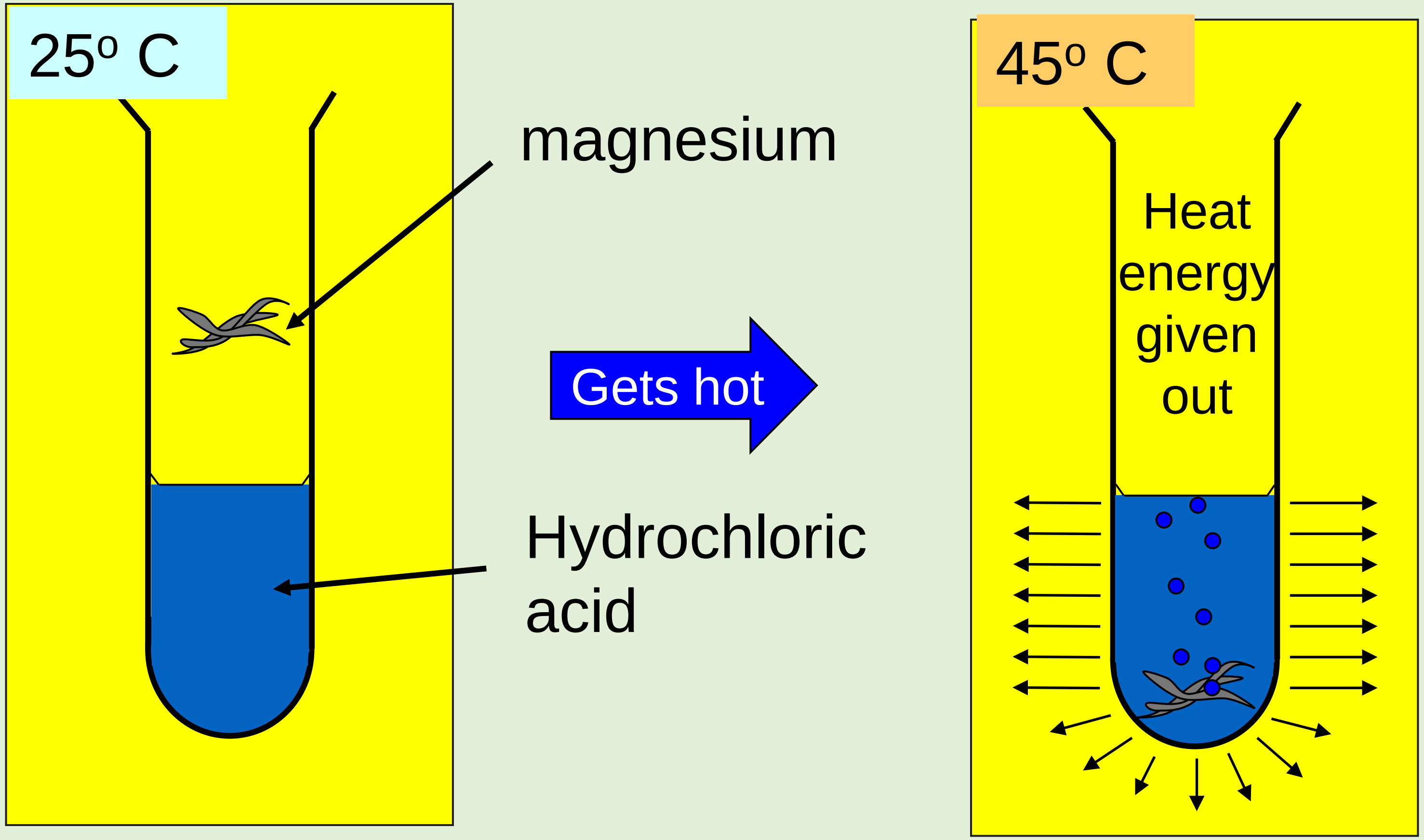
1. Charcoal burning — yes
2. A candle burning. — yes
3. A kettle boiling — no
4. Ice melting — no
5. A firework exploding — yes

You have to put heat *in* for boiling and melting.

You get heat *out* from all the other processes

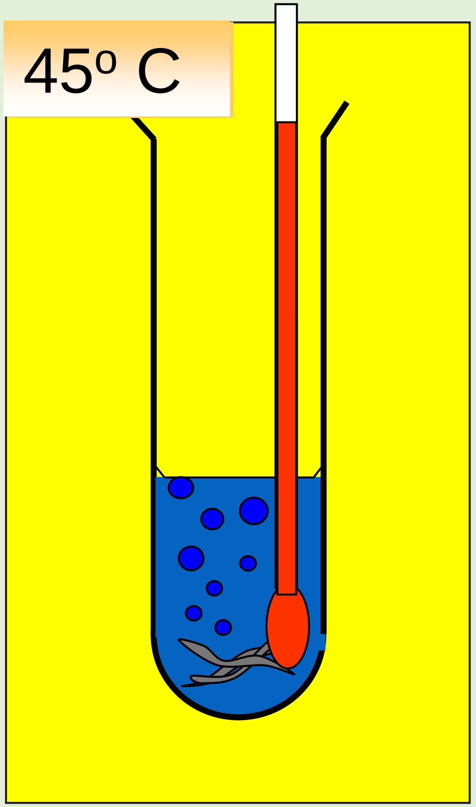
Exothermic Reactions

- Magnesium + Hydrochloric acid



Exothermic Reactions

- If heat is given out this energy must have come from chemical energy in the starting materials (reactants).

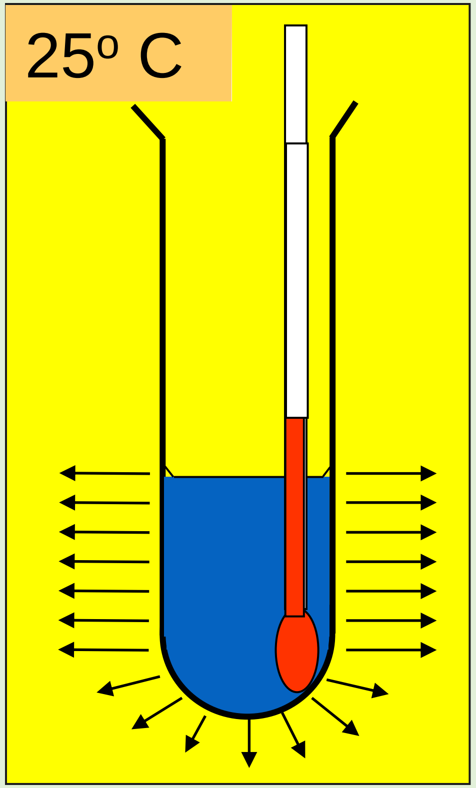


Reactants convert chemical energy to heat energy.

The temperature rises.

Exothermic Reactions

- Almost immediately the hot reaction products start to lose heat to the surroundings and eventually they return to room temperature.



Chemical energy becomes heat energy.

The reaction mixture gets hotter.

Eventually this heat is lost to the surroundings.

It follows that reaction products have less chemical energy than the reactants had to start with.

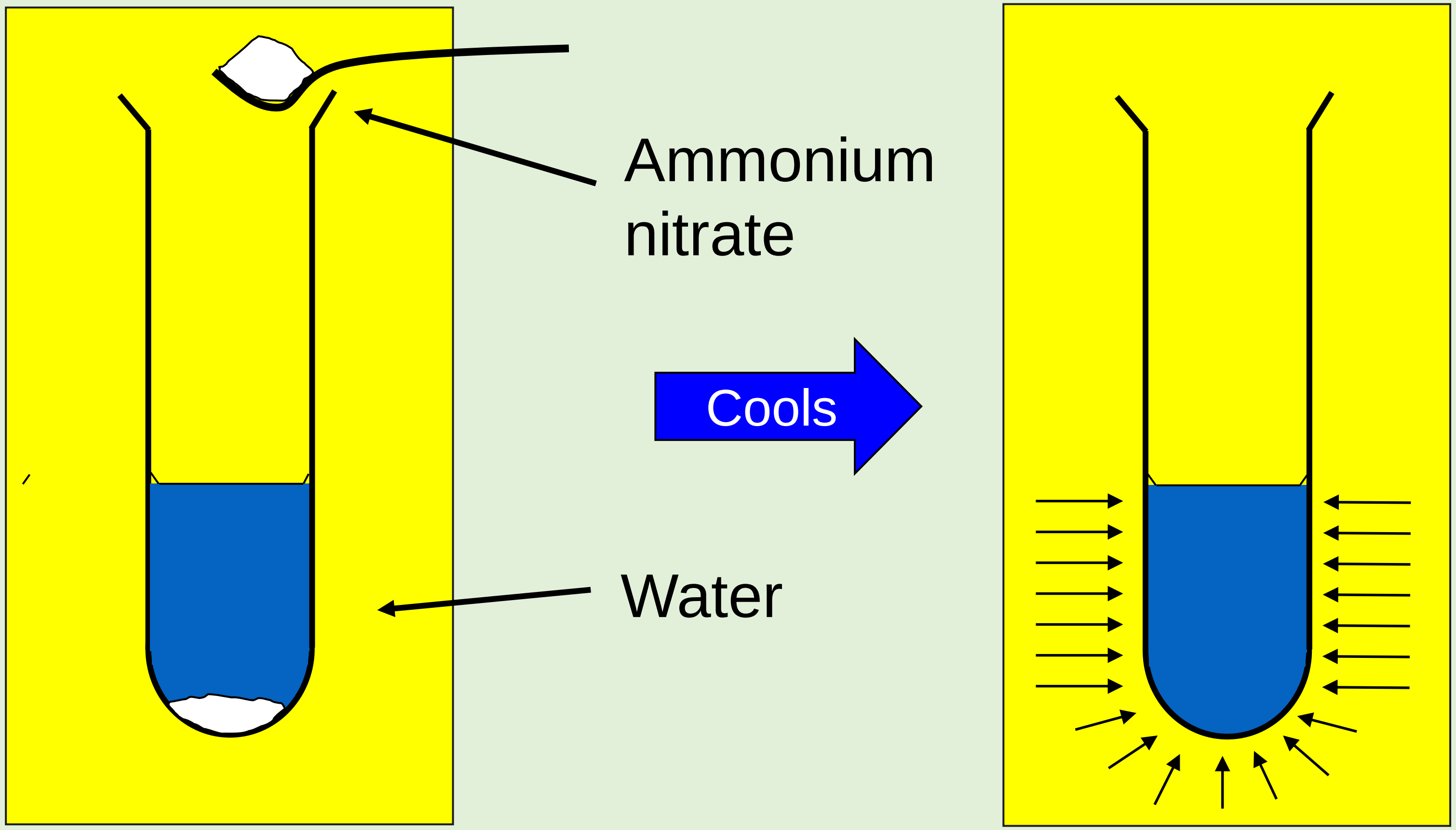
Endothermic Reactions

Endothermic reactions cause a decrease in temperature.

- Endothermic chemical reactions are relatively rare.
- A few reactions that give off gases are highly endothermic - get very cold.
- Dissolving salts in water is another process that is often endothermic.

Endothermic Reactions

Endothermic reactions cause a decrease in temperature.



Heat energy taken in as the mixture returns back to room temp.

Starts 25°C

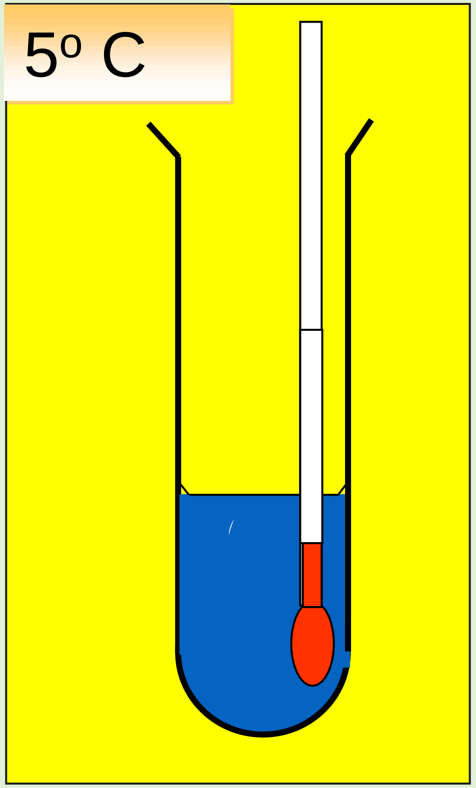
Cools to 5°C

Returns to 25°C

Endothermic Reactions

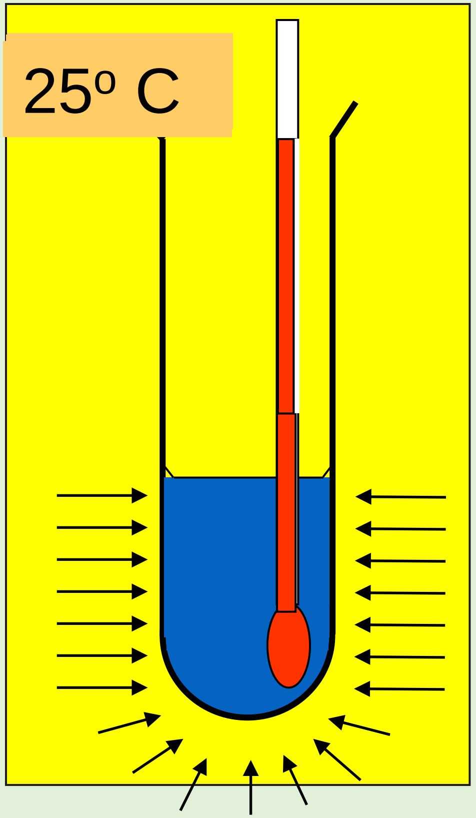
- Extra energy is needed in order for endothermic reactions to occur.
- This comes from the thermal energy of the reaction mixture which consequently gets colder.

Reactants convert heat energy into chemical energy as they change into products. The temperature drops.



Endothermic Reactions

- The cold reaction products start to gain heat from the surroundings and eventually return to room temperature.



The reactants gain energy.

This comes from the substances used in the reaction and the reaction gets cold.

Eventually heat is absorbed from the surroundings and the mixture returns to room temperature.

Overall the chemicals have gained energy.

Activity

Are these endothermic or exothermic?

1. A red glow spread throughout the mixture and the temperature rose.

exo

2. The mixture bubbled vigorously but the temperature dropped 15^0C.

endo

3. Hydrazine and hydrogen peroxide react so explosively and powerfully that they are used to power rockets into space.

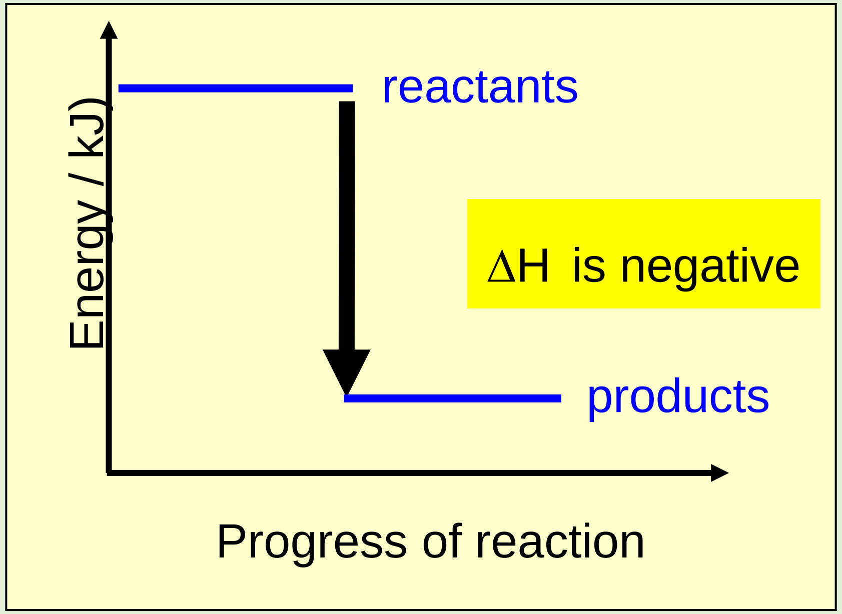
exo

4. The decaying grass in the compost maker was considerably above the outside temperature.

exo

Exothermic Reaction - Definition

Exothermic reactions give out energy. There is a temperature rise and ΔH is negative.



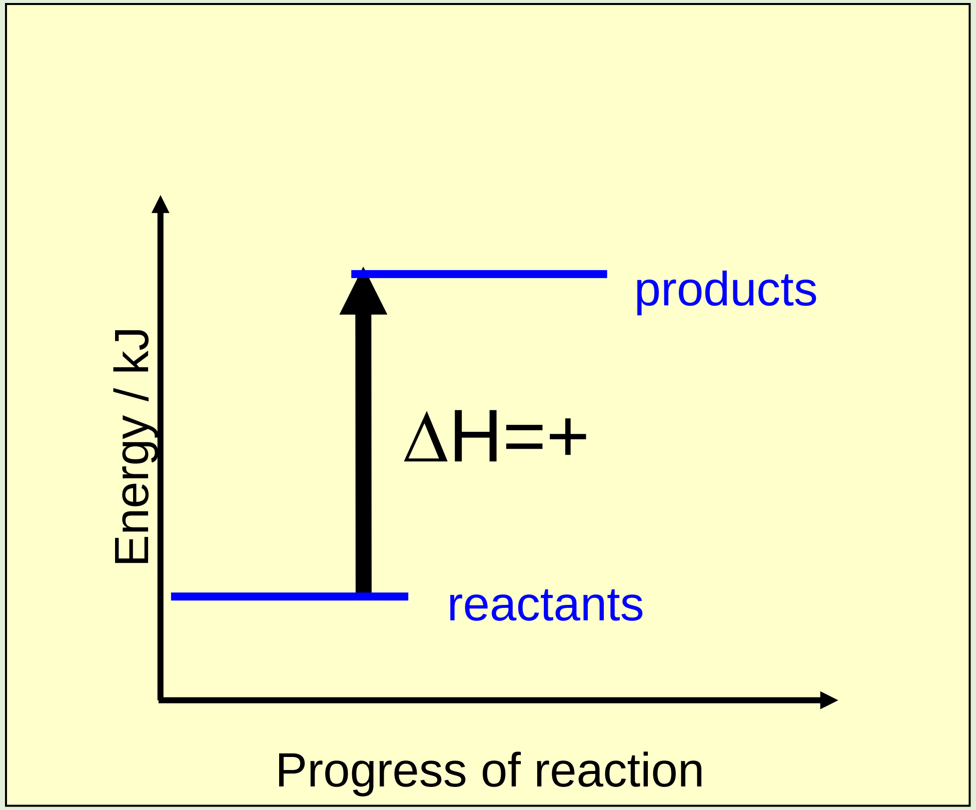
Activity

Using your knowledge of exothermic reactions decide whether the following statements are likely to be TRUE or FALSE.

1. $\triangle$H for burning magnesium is positive
2. Water is a substance that is high in chemical energy
3. Exothermic reactions convert chemical energy to heat energy
4. In an explosion the products have more energy than the reactants

Endothermic Reaction Definition

Endothermic reactions take in energy. There is a temperature drop and ΔH is positive.



Using your knowledge of endothermic reactions decide whether the following statements are likely to be TRUE or FALSE.

1. Endothermic reactions get hot — T F
2. Endothermic reactions are less common than exothermic ones — T F
3. ΔH for endothermic reactions is positive — T F
4. The products have more energy than the reactants — T F

Activity

Are these endothermic or exothermic?

1. A red glow spread throughout the mixture and the temperature rose.

exo

2. The mixture bubbled vigorously but the temperature dropped 15^0C.

endo

3. Hydrazine and hydrogen peroxide react so explosively and powerfully that they are used to power rockets into space.

exo

4. The decaying grass in the compost maker was considerably above the outside temperature.

exo

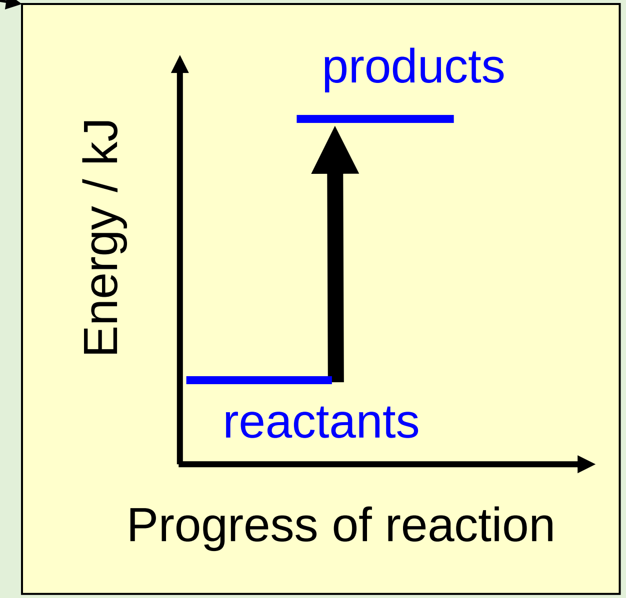
Activity

- The formation of nitrogen (IV) oxide (formula NO_2) from reaction of nitrogen with oxygen in car engines has a ΔH value of +33.2kJ per mol of nitrogen oxide.

1. Write a word equation for the reaction.
2. Write a chemical equation for the reaction.
3. Is ΔH positive or negative?
4. Is the reaction exothermic or endothermic?
5. Draw an simple energy diagram for the reaction (not showing bond breaking and forming.)
6. Which involves the biggest energy change: bond breaking or bond forming?

Answer

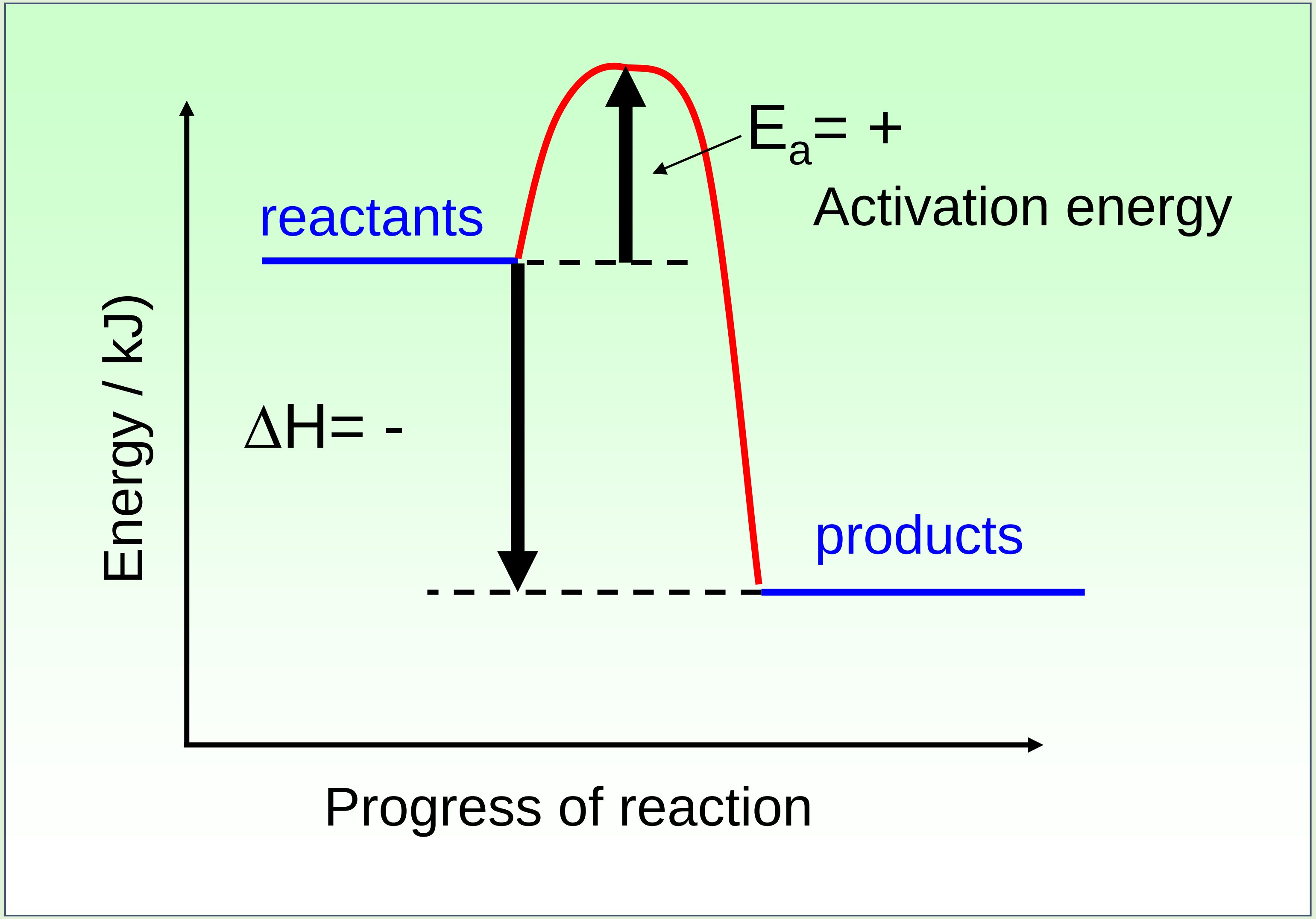
1. Nitrogen + oxygen ⇨ nitrogen(IV)oxide
2. $N_2 + 2O_2 \Rightarrow 2NO_2$.
3. ΔH positive (+33.2kJ/mol).
4. The reaction is endothermic.
5. Energy diagram
6. Bond breaking involves the biggest energy change.



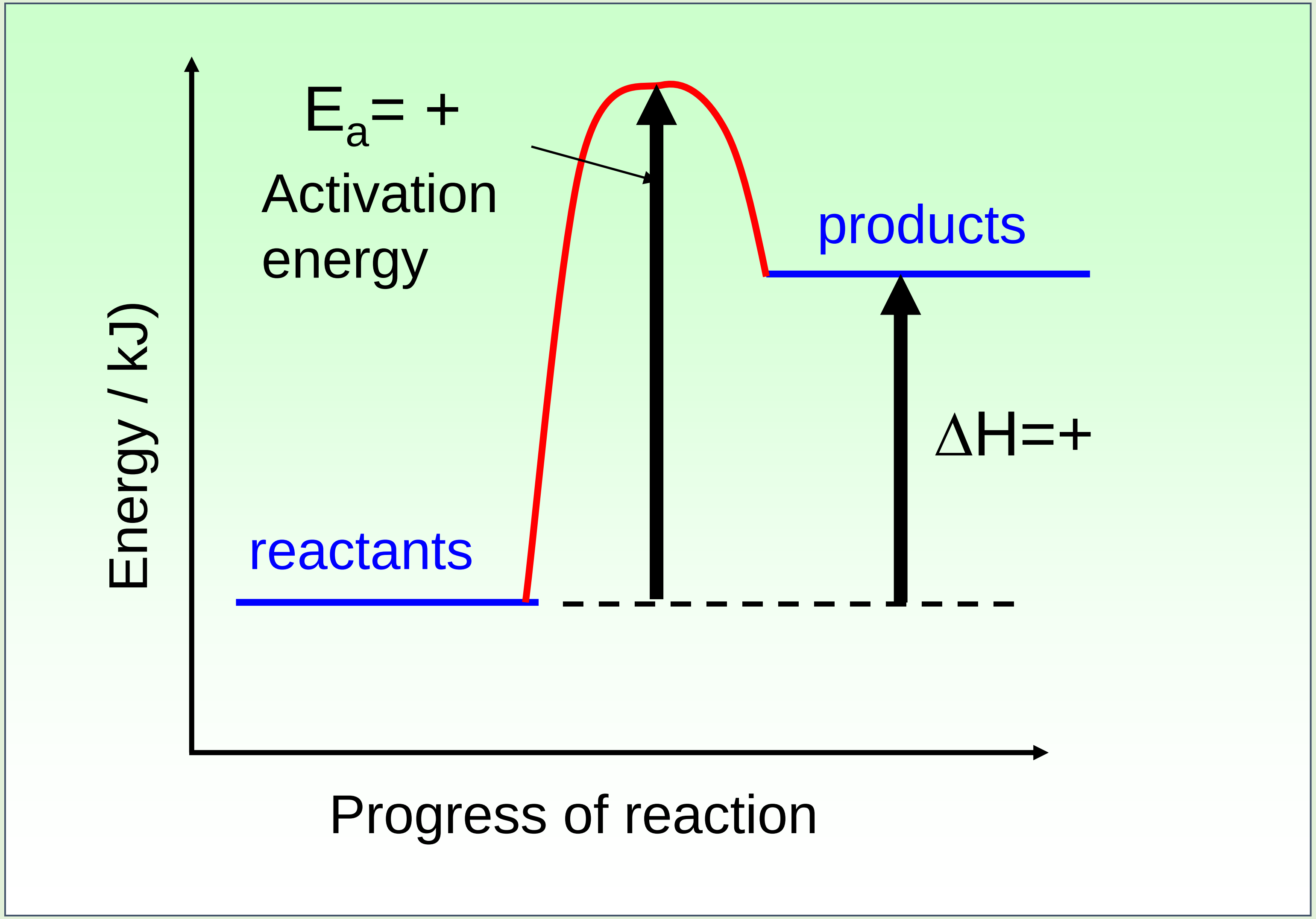
Activation Energy.

- Most chemical reactions, including exothermic reactions, seem to need an input of energy to get the reaction started.
- This fits completely with what we have already explained:
  - Before new bonds can be formed we need to break at least some existing chemical bonds.
  - This requires an energy input –known as the *activation energy* (E_a or E_{act})
  - Once an exothermic reaction is underway it can provide its own activation energy (from the bond forming stage) and so sustains the reaction.

Activation Energy and Exothermic Reactions



Activation Energy and Endothermic Reactions



Activity

Copy the summary using the words from the box to fill in the gaps:

endothermic	lose	positive
exothermic	common	

1. Exothermic reactions are common
2. Reactions that get cold are called endothermic
3. Bond forming is an exothermic process.
4. Endothermic reactions have a positive ΔH.
5. In exothermic reactions the chemicals lose chemical energy.

continued

Activity

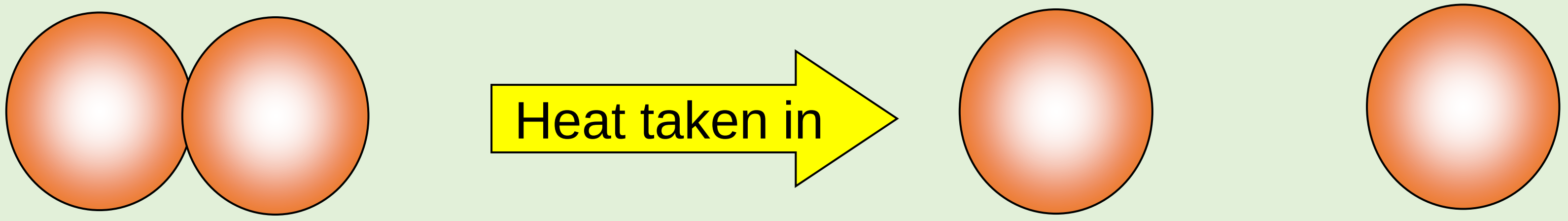
Copy the summary using the words from the box to fill in the gaps:

more	endothermic	activation

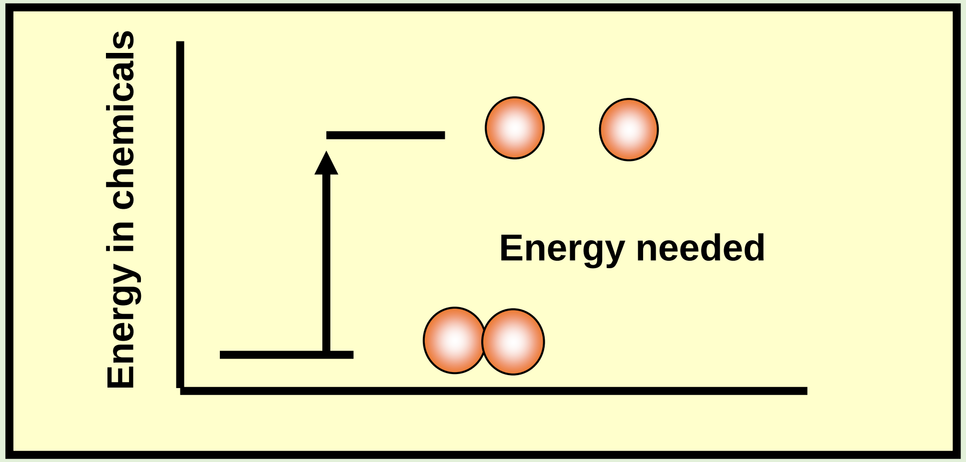
1. The energy needed to start off a reaction is called the activation energy
2. In endothermic reactions bond breaking requires more energy than is produced by bond forming.
3. Bond breaking is an endothermic process.

Breaking chemical bonds

- Most chemicals will decompose (break up) if we heat them strongly enough.
- This indicates that breaking chemical bonds requires energy – is an endothermic process.



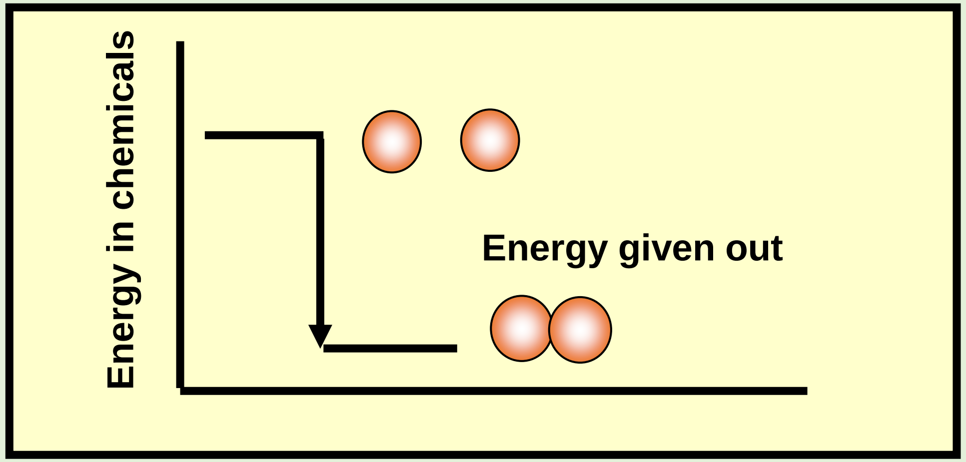
Energy needed to overcome the bonding between the atoms



Making chemical bonds

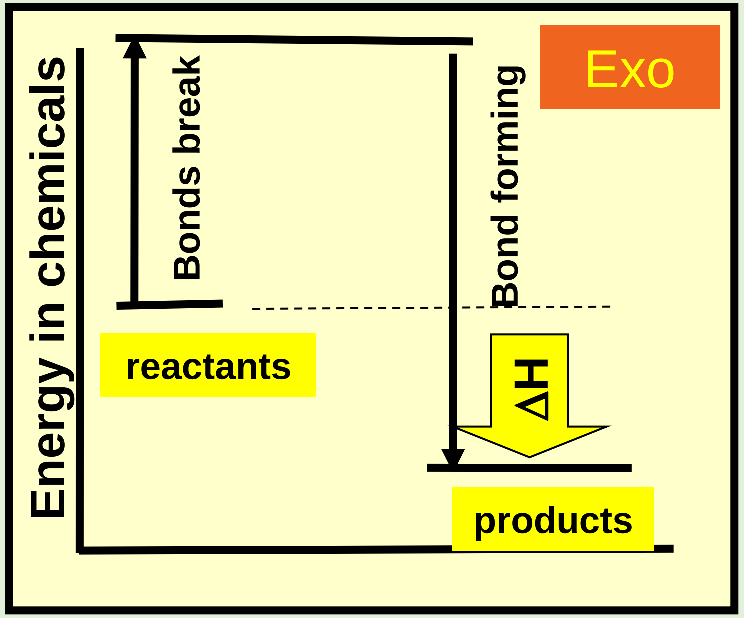
- It is reasonable to assume that bond making will be the opposite of bond breaking
- Energy will be given out in an exothermic process when bonds are formed.

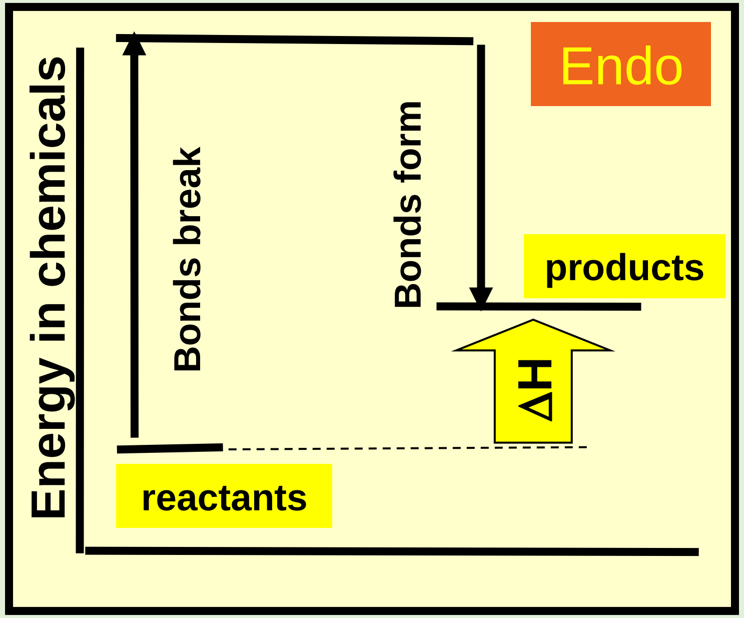
Energy given out as bonds form between atoms



Summary – Bond Changes

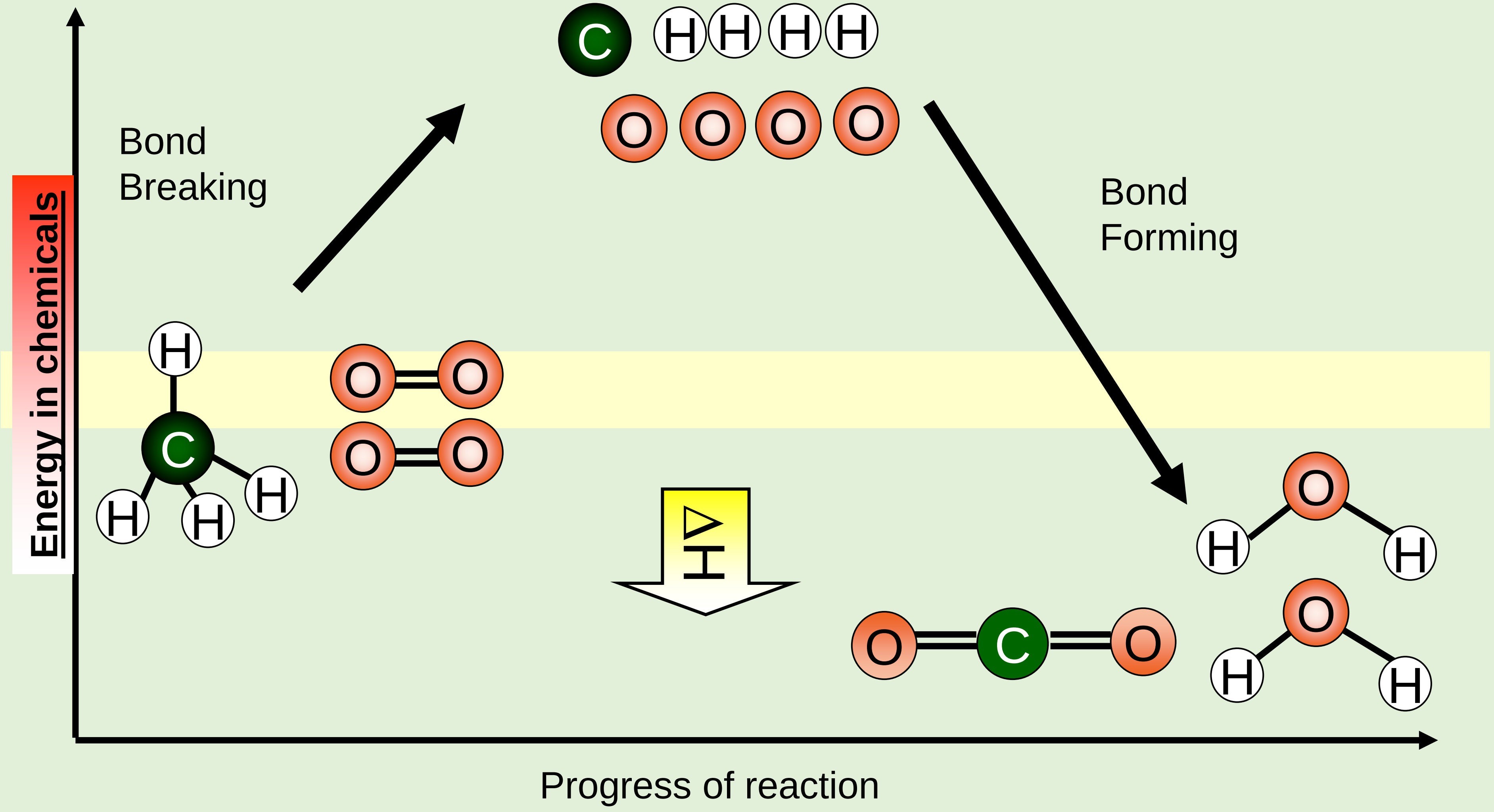
- Where the energy from bond forming exceeds that needed for bond breaking the reaction is exothermic.
- Where the energy for bond breaking exceeds that from bond forming the reaction is endothermic.





Burning Methane

- This is an exothermic reaction

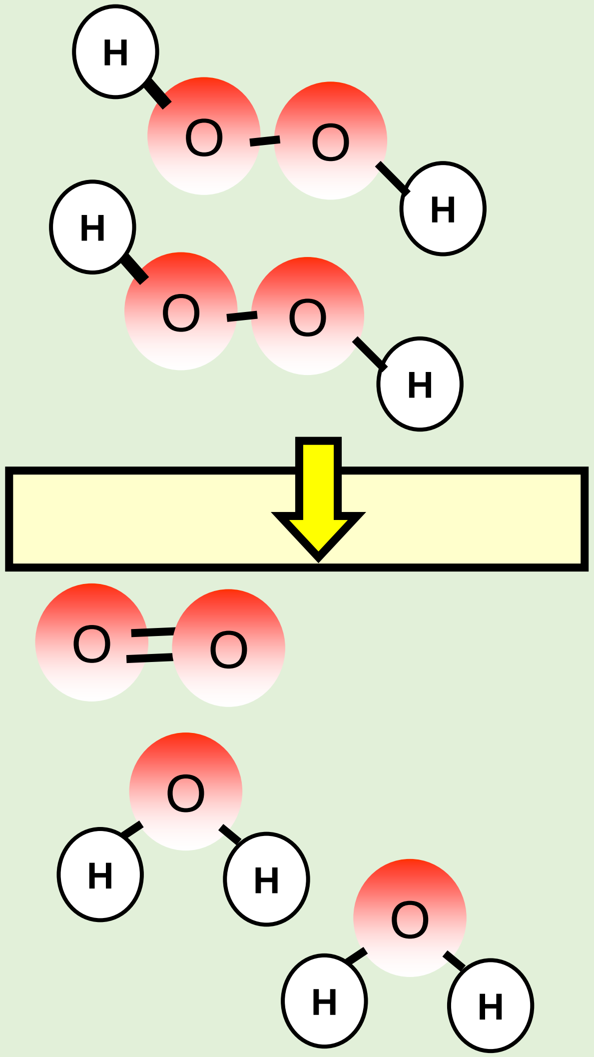


Activity

- Hydrogen peroxide decomposes as shown:

Bond	Energy (kJ)
H-O	464
O-O	146
O=O	498

1. Calculate energy for bond breaking.
2. Calculate the energy from bond making
3. What is the value of ΔH for the reaction shown



Answer

Bond	Energy (kJ)
H-O	464
O-O	146
O=O	498

Bond breaking. (endothermic)

4(O-H) + 2(O-O)

=1856+292 = +2148kJ

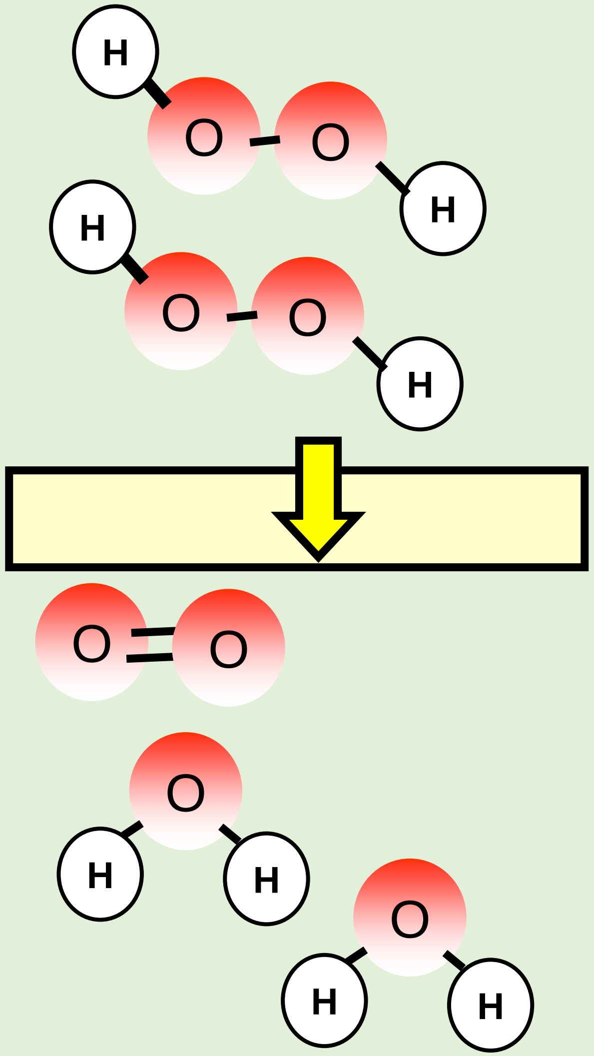
Bond forming: (exothermic)

4(O-H) + 1(O=O)

=1856+498 = -2354kJ

ΔH = +2148 – 2354 = -206kJ

(Exothermic)



Combustion of methane

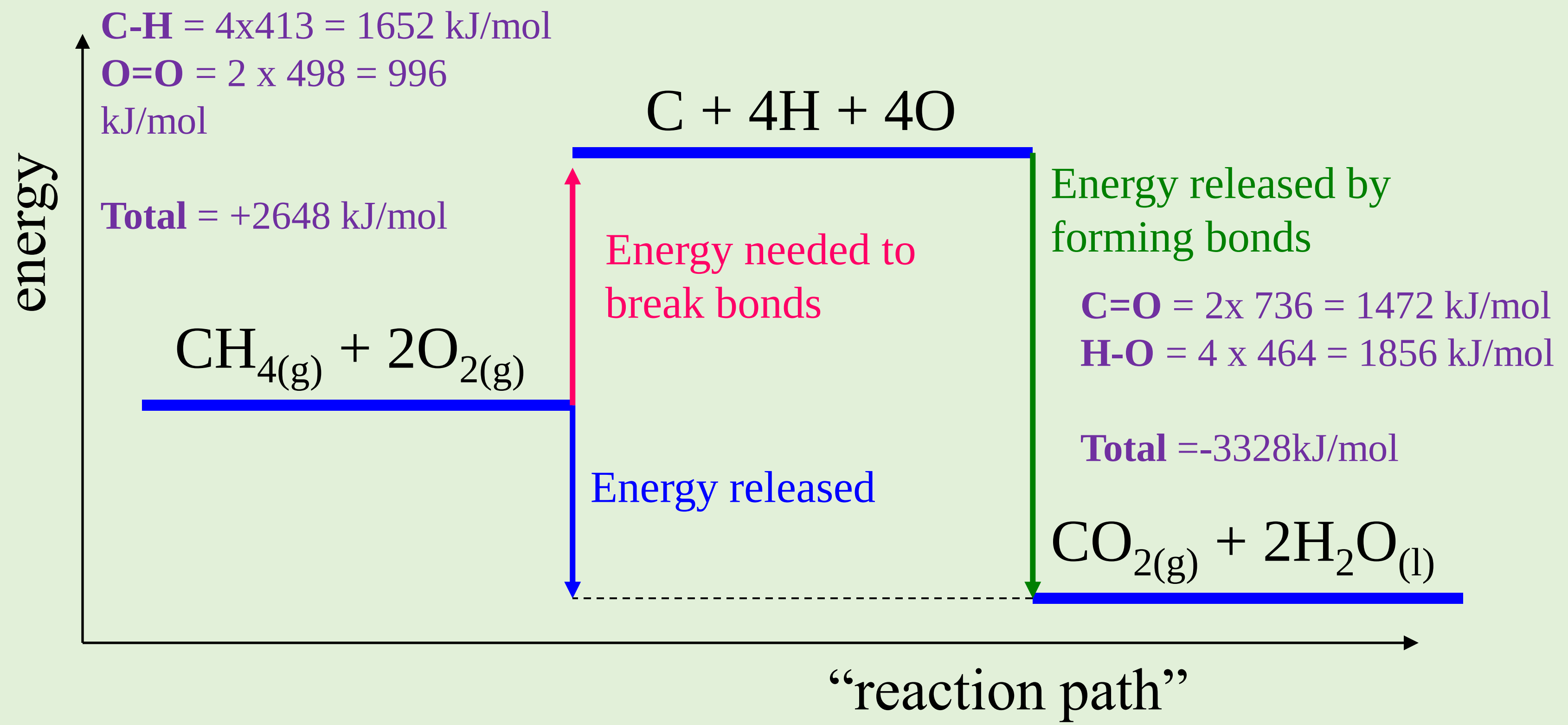
$$CH_{4(g)} + 2O_{2(g)} \longrightarrow 2H_2O_{(l)} + CO_{2(g)}$$

What bonds do we break and what bonds do we make?

Exothermic reaction

The energy need to break the bonds is less than the energy released when new bonds are made

Bond	Energy (kJ/mol)
C-H	413
O=O	498
C=O	736
H-O	464



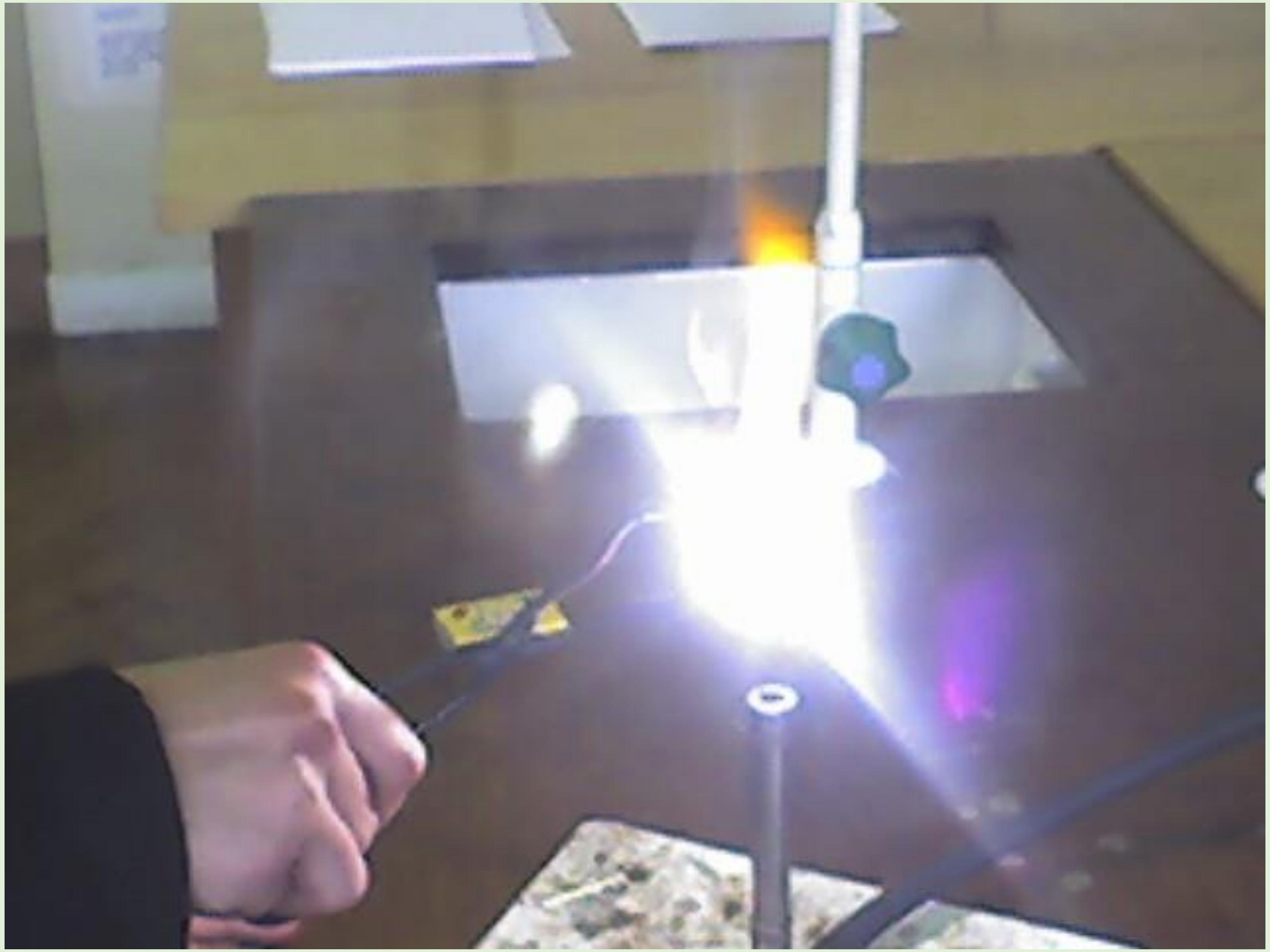
Overall energy change = (+2648) + (-3328) = ΔH -680 kJ/mol

Summary

- Exothermic reactions:
  - Are common,
  - Give out heat.
  - Have a negative ΔH.
  - Bond forming gives out more energy than bond breaking consumes.
  - Have reactants that contain more chemical energy than the products.
- Endothermic reactions – are the opposite!
- Bond breaking is endothermic.
- Bond forming is exothermic.
- Reactions require an activation energy to help start the bond breaking process.

REDOX

Burning Magnesium in air



Oxidation is gain of oxygen

- When magnesium reacts with oxygen in the air:

- $2Mg + O_2 \longrightarrow 2MgO$

- It is easy to see that the magnesium has gained oxygen and an oxidation reaction has occurred.

- The magnesium has been OXIDISED.

Reduction of lead(II)oxide

Reduction is loss of oxygen

- When lead is extracted from lead (II) oxide , carbon is used to remove the oxygen

- $2PbO + C \longrightarrow 2Pb + CO_2$

- The oxygen has been removed from the lead (II) oxide

- The lead (II) oxide has been REDUCED.

Lets look at this further…

- What actually happens when the magnesium joins with oxygen?
- The magnesium atom LOSES ELECTRONS to become a magnesium ion
- We can write what happens like this:
- $Mg \longrightarrow Mg^{2+} + 2e^-$

Oxidation is loss of electrons

- So, when magnesium takes part in an oxidation reaction it loses electrons.

- When magnesium reacts with chlorine it also loses electrons.

- $Mg + Cl_2 \longrightarrow MgCl_2$

- This is an oxidation reaction too!

Reduction is gain of electrons

- In the compound lead (II) oxide the lead is an ion, Pb^{2+}

- In order to form lead the Pb^{2+} ion has to GAIN 2 electrons

- We can write this as follows:

- $Pb^{2+} + 2e^- \longrightarrow Pb$

Redox

- You may have noticed that when the magnesium loses its electrons they are gained by the oxygen
- Oxidation (loss of electrons) and reduction (gain of electrons) happen simultaneously.
- We call these REDOX reactions

OILRIG

Oxidation

Is

Loss

Reduction

Is

Gain

(of electrons)

REDOX

- For a reaction to be a redox reaction, different reactants have to be oxidised or reduced.

- **OXIDISING AGENTS** oxidise other chemicals and are themselves **REDUCED**.
- **REDUCING AGENTS** reduce other chemicals and are themselves **OXIDISED**.


Hey, Litmus Paper, why so blue?
Basic Humor

The pH scale

Types of Chemical Reaction

How we identify acids and alkali’s

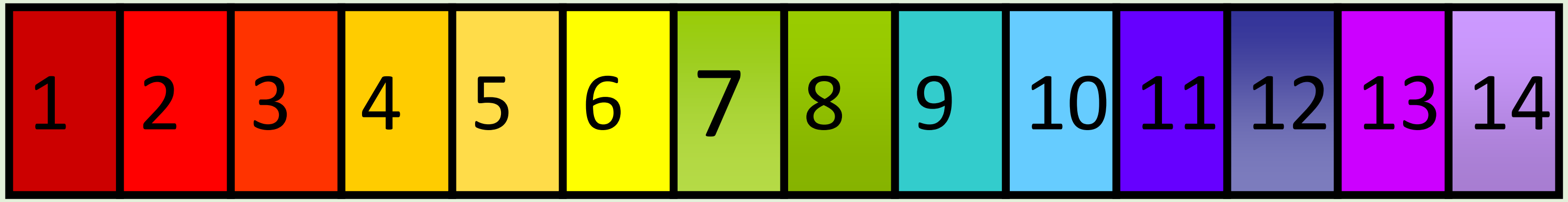
- We identify acids and alkali’s based on their pH using indicators.
- Indicators change colour depending on the pH
- E.g.
  - Universal indicator (range of colours)
    - Paper
    - Solution
  - Litmus paper
    - Blue (turns red in acid)
    - Red (turns blue in alkali)

Indicators: the pH scale

This attaches a number called the pH value to each universal indicator colour.

- pH7 is neutral
- pH 1 is strongly acid
- pH14 is strongly alkali

This means we can quickly say how acid or alkali a substance is by quoting a single number.



Strong acid Neutral Strong alkali

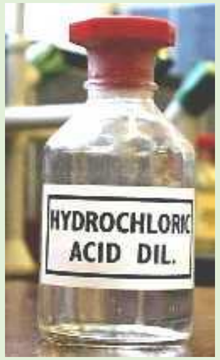
Weak acid Weak alkali

Some common acids

Three acids are particularly common in the laboratory .

Acid	Formula
hydrochloric	HCl
sulphuric	H_2SO_4
nitric	HNO_3

These are strong acids that should be treated with the greatest respect.



Bases

Bases are substances that neutralise acids.

Bases are usually:

- Metal hydroxides contain **OH**
- Metal oxides contain **O**
- Metal carbonates contain $\mathbf{CO_3}$

The following general word equation describes neutralisations:

acid + base → a salt + water

In the case of carbonates we also get carbon dioxide.

Chemical Reactions: Neutralisation

Neutralisation reactions

- **Acids** are substances that:
- Turn litmus red.
- Turn universal indicator yellow, orange or red.
- Have a pH below 7.
- Form solutions containing H^+ ions.
- **Bases** are substances that:
- Turn litmus blue.
- Turn universal indicator dark green, blue or purple.
- React with the H^+ ions in acids.
- Are called alkalis if they dissolve in water.

1	2	3	4	5	6	7	8	9	10	11	12	13	14

Increasingly acid **Increasingly alkali**

Neutralisation reactions: acids

- Common Acids are

Name of acid	Formula	Strong or Weak?
Sulphuric	H_2SO_4	strong
Hydrochloric	HCl	strong
Nitric	HNO_3	strong
Ethanoic (vinegar)	CH_3COOH	weak

- Salts

Sulphuric acid → Sulphates

Nitric acid → Nitrates

Hydrochloric acid → Chlorides

Neutralisation reactions: bases

- Common alkalis are

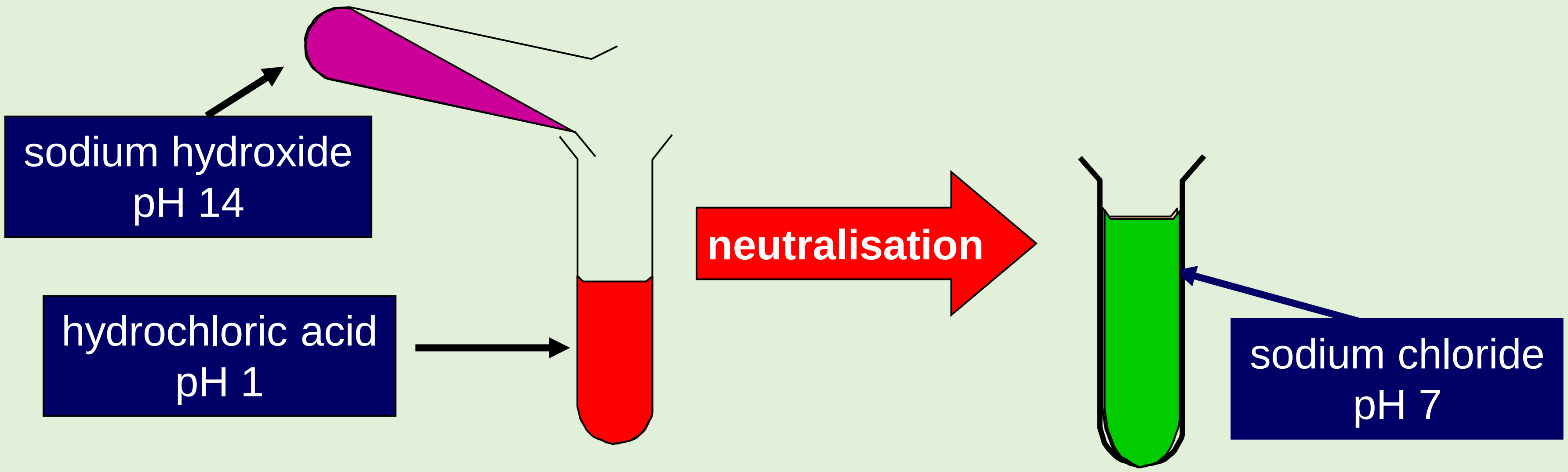
Name of alkali	Formula	Strong or Weak?
Sodium Hydroxide	$NaOH$	strong
Potassium Hydroxide	KOH	strong
Calcium Hydroxide	$Ca(OH)_2$	strong
Ammonium Hydroxide	NH_4OH	weak

- Common bases (neutralise acids but don’t dissolve) are

Type of compound	Contain	React with acids to give
Metal Hydroxides	OH^-	water + a salt
Metal Oxides	O^{2-}	water + a salt
Metal Carbonates	CO_3^{2-}	water + a salt + CO_2

Neutralisation reactions: acid + base

A neutralisation reaction is where an **acid** reacts with a **base** to produce a neutral solution of **a salt** and **water**.



1	2	3	4	5	6	7	8	9	10	11	12	13	14

← **Increasingly acid** **Increasingly alkali** →

Neutralisation - naming salts

To name the salt formed in a neutralisation:

1 The first part of the name of the salt comes from the first name of the base

So Ammonium hydroxide gives ammonium …………
Magnesium oxide gives magnesium …………...

2 The acid gives the last part of the name of the salt.

So Sulphuric acid make **sulphates**
Nitric acid makes **nitrates**
Hydrochloric acid makes **chlorides**

Eg. Sodium hydroxide + nitric acid forms: **Sodium nitrate**

Calcium carbonate + sulphuric acid forms: **calcium sulphate**

Name the salt formed in these neutralisations:

Base	Acid	Salt?
Calcium hydroxide	Hydrochloric acid	Calcium chloride
Magnesium oxide	Nitric acid	Magnesium nitrate
Calcium carbonate	Sulphuric acid	Calcium sulphate
Aluminium hydroxide	Nitric acid	Aluminium nitrate
Potassium hydroxide	Sulphuric acid	Potassium sulphate

Neutralisation reactions: hydroxides

Each OH^- ion reacts with one H^+ ion.

Reaction with hydroxides: $H^+ + OH^- \rightarrow H_2O$

Eg.	Potassium hydroxide	+	hydrochloric acid	→	water +	potassium chloride
	KOH	+	HCl	→	H_2O +	KCl

Eg.	Calcium hydroxide	+	sulphuric acid	→	water +	calcium sulphate
	$Ca(OH)_2$	+	H_2SO4	→	$2H_2O$ +	$CaSO_4$

Neutralisation Reactions: oxides

Neutralisation reactions usually lead to water being formed.

Reaction with oxides: $2H^+ + O^{2-} \rightarrow H_2O$

Eg.	Calcium oxide	+	hydrochloric acid	→	water	+	calcium chloride
	CaO	+	$2HCl$	→	H_2O	+	$CaCl_2$

Eg.	Sodium oxide	+	sulphuric acid	→	water	+	sodium sulphate
	Na_2O	+	H_2SO4	→	H_2O	+	Na_2SO_4

Neutralisation Reactions: carbonates

Each carbonate ion provides one oxygen to join with two H+ ions. At the same time carbon dioxide is released.

Carbonates: $2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$

Eg.	Potassium carbonate	+	hydrochloric acid	→	water	+	carbon dioxide	+	potassium chloride
	K_2CO_3	+	$2HCl$	→	H_2O	+	CO_2	+	$2KCl$

Eg.	calcium carbonate	+	nitric acid	→	water	+	carbon dioxide	+	calcium nitrate
	$CaCO_3$	+	$2HNO_3$	→	H_2O	+	CO_2	+	$Ca(NO_3)_2$

Neutralisation equations

Complete the word equation

Eg.	Potassium hydroxide	+	hydrochloric acid	→	water	+	Potassium chloride

Replace the words with the **correct formula**

Eg.	KOH	+	HCl	→	H_2O	+	KCl

Check that it <u>balances</u> (same number of each type of atom each side).

Eg. KOH + HCl → H_2O + KCl ✓

Reactants				Products			
1*K	1*O	2*H	1*Cl	2*H	1*O	1*K	1*Cl

Neutralisation equations

Complete the word equation

Eg.	Magnesium oxide	+ nitric acid ➔	water	+	Magnesium nitrate

Replace the words with the **correct formula**

Eg.	MgO	+ HNO_3 ➔	H_2O	+	$Mg(NO_3)_2$

Check that it <u>balances</u> (Same number of each type of atom each side.

Eg.	**MgO**	+ **2 HNO_3** ➔	**2 H_2O**	+	**$Mg(NO_3)_2$** ✓

Reactants				Products			
1*Mg	1*O	***1*H***	***1*NO_3***	***2*H***	1*O	1*Mg	***2*NO_3***

Write balanced equations going through the same stages as the previous examples.

1. word equation
2. formulae
3. balance

a) sodium hydroxide + hydrochloric acid □

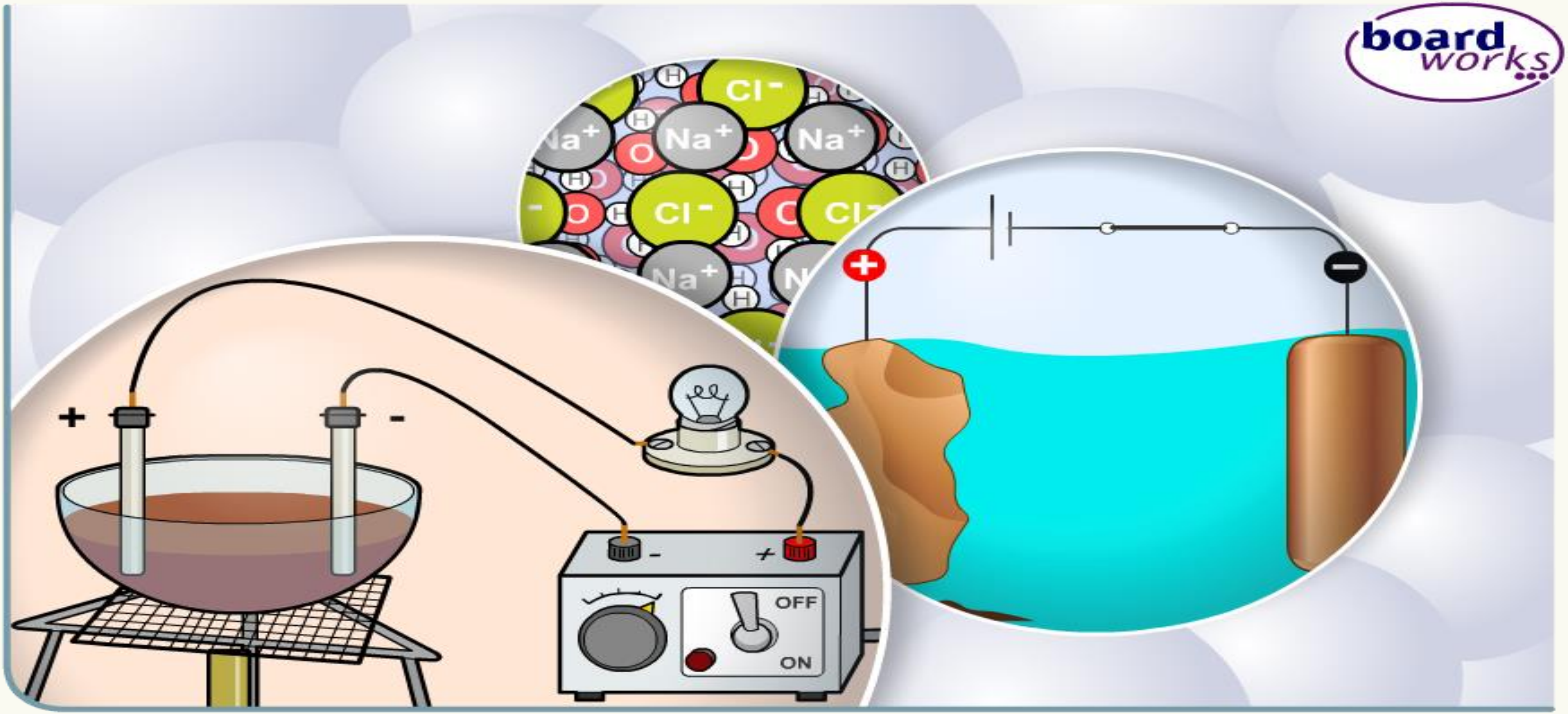
b) magnesium oxide + hydrochloric acid □

c) sodium hydroxide + sulphuric acid □

d) ammonium hydroxide + hydrochloric acid □

e) calcium hydroxide + nitric acid □

GCSE Chemistry



Electrolysis

Indicates a Flash activity.

Indicates an accompanying worksheet.

Indicates that 'How Science Works' skills are covered.

Indicates a virtual experiment.

Indicates that there are teacher's notes.

For more detailed instructions, see the *Getting Started* presentation.

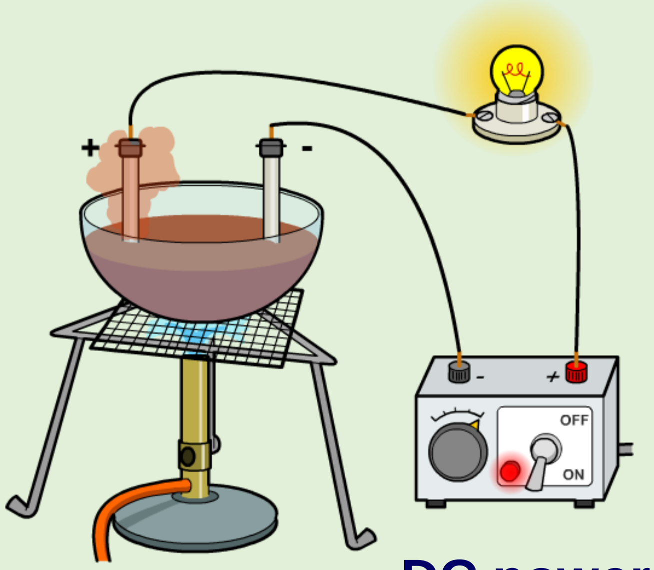
What is electrolysis?

Electrolysis is the decomposition of an **ionic substance** by passing an electric current through it. The ionic substance is called the **electrolyte**.

An ionic compound is formed of charged particles called **ions**.

- positive ions are called **cations**
- negative ions are called **anions**.

When solid, the ions are fixed in position within the compound.



DC power supply

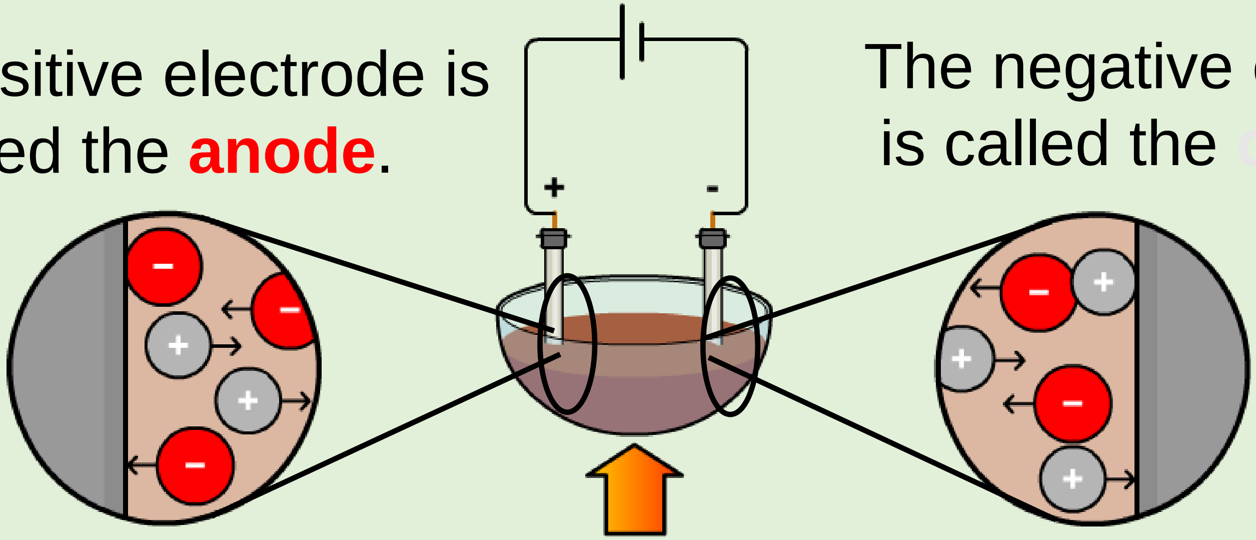
When molten or in solution, the ions are able to move freely through the liquid. The liquid can conduct electricity.

Oxidation and reduction

When a current is passed through an ionic compound that is molten or in solution, the ions move towards the electrodes and discharge.

The positive electrode is called the **anode**.

The negative electrode is called the cathode.



heat

Anions move to the anode and **lose electrons**. This is **oxidation**.

Cations move to the cathode and **gain electrons**. This is **reduction**.

Half equations

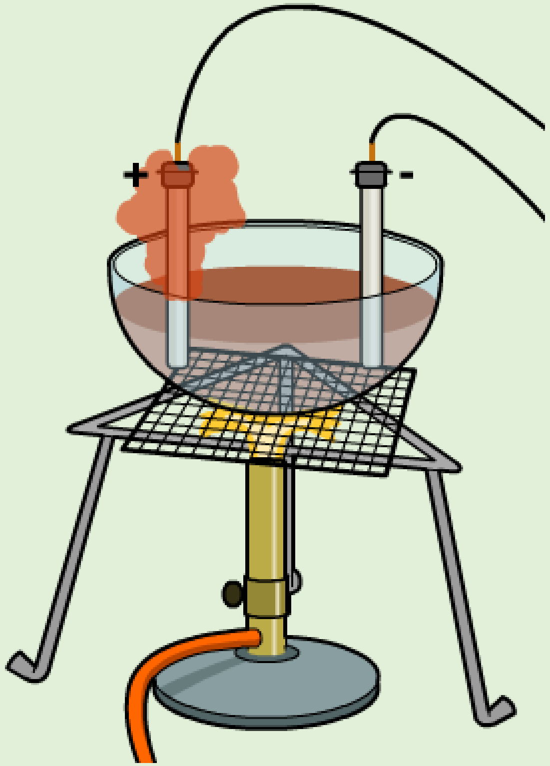
What happens at the electrodes during electrolysis can be described using **redox equations**, also called **half equations**.

What are the half equations for the electrolysis of molten lead bromide?

At the anode:

$2Br^{-} - 2e^{-} \rightarrow Br_2$

This means that each bromine ion has one extra electron.

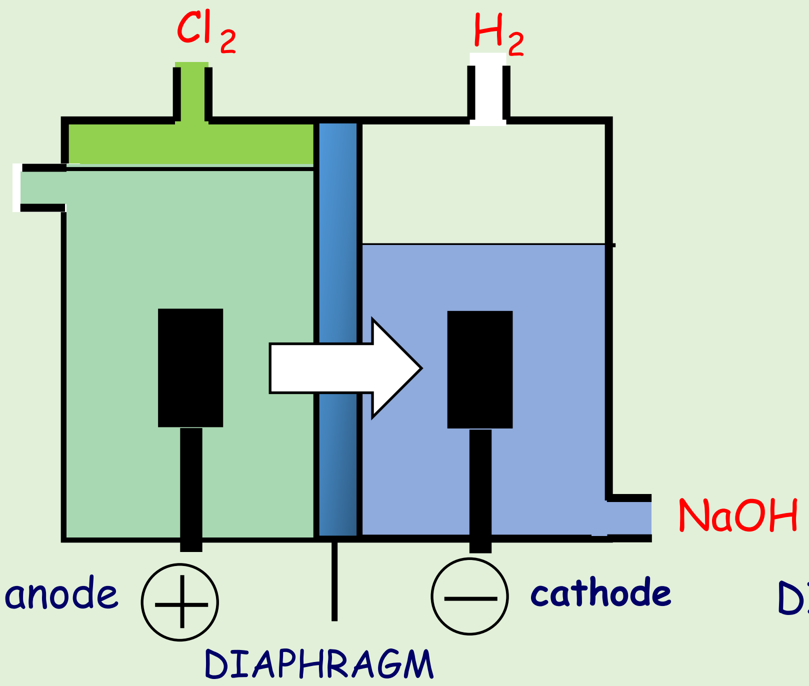


At the cathode:

$Pb^{2+} + 2e^{-} \rightarrow Pb$

This means that the lead ion is missing two electrons.

THE INDUSTRIAL ELECTROLYSIS OF SODIUM CHLORIDE



The electrolytic cell:
The diaphragm is a porous material
The brine seeps through from
the anode compartment to the cathode
The higher level of liquid maintains this
flow – preventing NaOH from finding
its way back into the chlorine section.

DISCHARGE of CHLORINE at the ANODE:

$$2Cl^-_{(l)} - 2e^- \longrightarrow Cl_{2(g)}$$

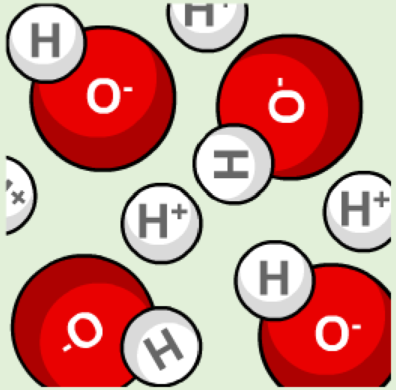
DISCHARGE of HYDROGEN at the CATHODE:

$$2H^+_{(aq)} + 2e^- \longrightarrow H_{2(g)}$$

Remaining ions in solution: $Na^+_{(aq)}$ $OH^-_{(aq)}$ producing **NaOH**

Why is sodium not formed?

In the electrolysis of sodium chloride solution, the Na^+ ions might be expected to form sodium at the negative electrode. Instead, hydrogen gas is produced here.



This is because the sodium chloride is in solution. The electrolyzed solution also contains H^+ ions from the water:

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

At the cathode, H^+ ions compete with Na^+ ions. The H^+ ions gain electrons; the Na^+ ions stay in solution.

For all ionic compounds containing a **metal that is more reactive than hydrogen**, electrolysis of a solution of the compound will produce hydrogen rather than the metal.

Manufacturing sodium

If sodium is not produced by the electrolysis of **sodium chloride solution**, can electrolysis still be used to manufacture sodium?

Sodium chloride can undergo electrolysis in either of two forms:

- in **solution**
- when **molten**.

In the electrolysis of molten sodium chloride, there are no hydrogen ions from water to be used instead of the sodium ions.

Pure sodium forms at the cathode.

Sodium separated in this way has many uses, from providing light in street lamps to being a coolant in nuclear reactors.

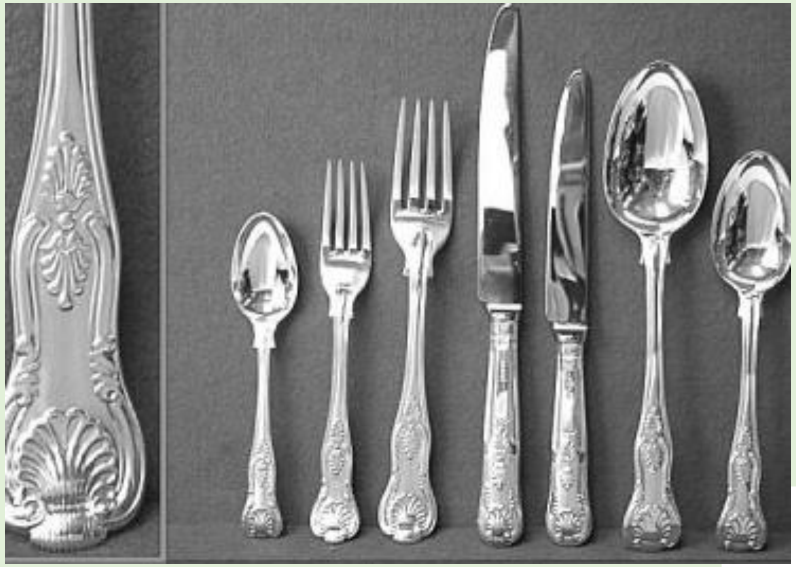
So what is electro(plating)? and why?

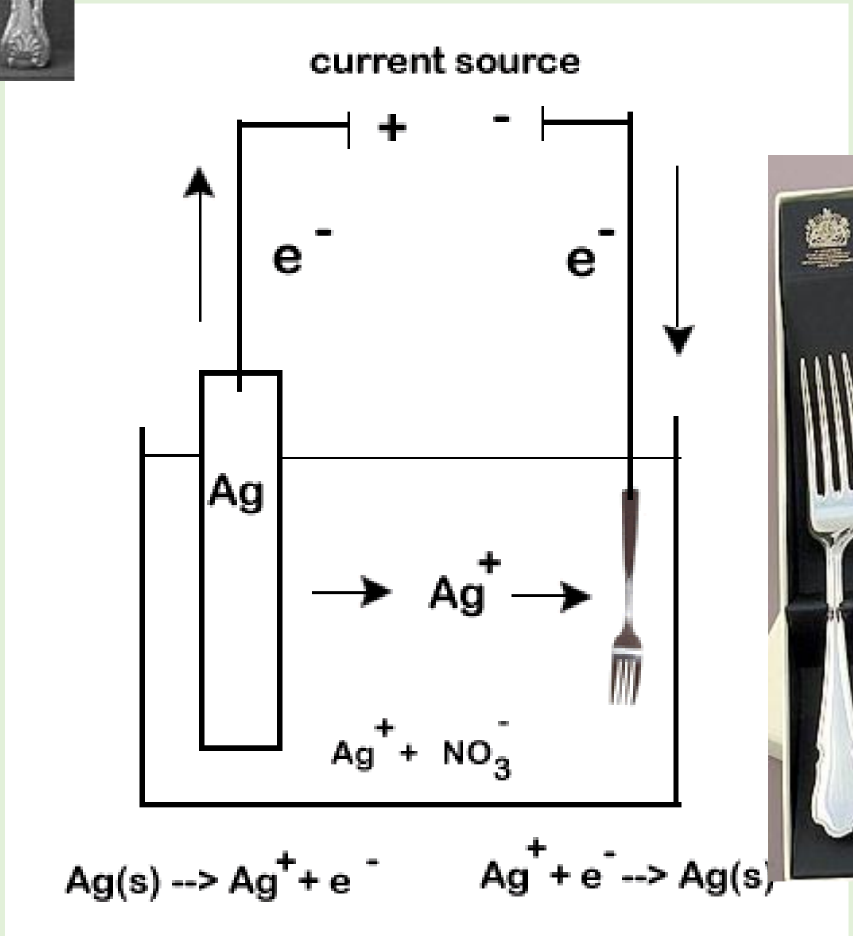
- Electroplating is a technique in which a t_ _ _ layer of a desired metal is used to coat (or "plate") another object. This process is often used to protect objects against corrosion or to improve their appearance.
- For our example we will exam the copper plating of flatware. In our example we will coat a fork made with an inexpensive metal with a thin layer of copper.

So where lies the connection?

- It's electrochemistry or electrolysis
- As with our other electrolytic cells we have three requirements:
- An E__________:- Our e____________ solution will need to contain ions of the plating metal. We will use $CuSO_4$ which will give us our required ___ ions.
- a source of current - a _________, and
- 2 E___________:- One of them will be the object to be coated (the fork), while the other must be the plating metal (a bar of copper)

Silver plated cutlery





Half equations for silver plating

- Cathode → reduction

  $Ag^+(aq) + e^- \rightarrow Ag(s)$
- Anode → oxidation

  $Ag(s) \rightarrow Ag^+(aq) + e^-$