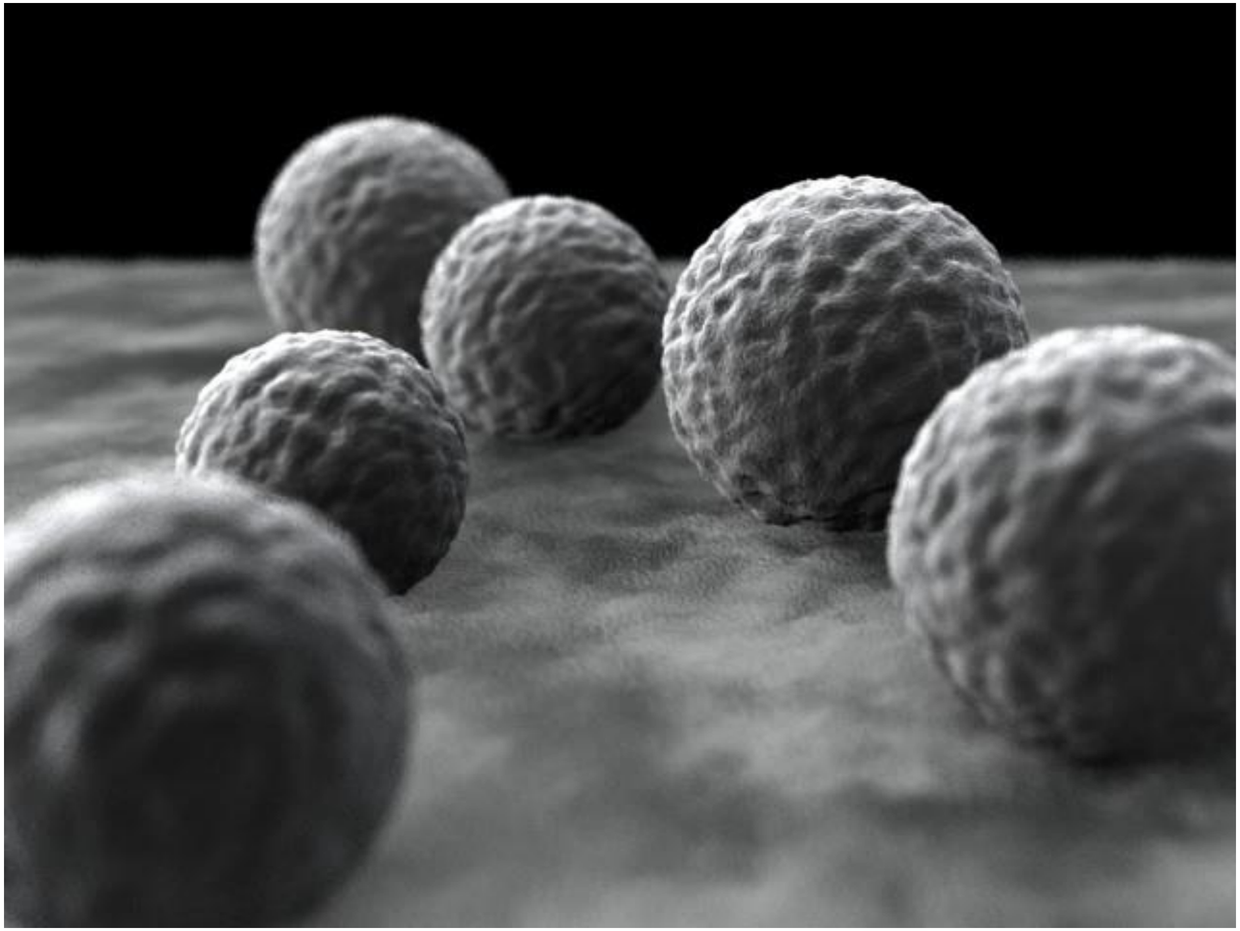
Module 10

Understanding chemical reactions

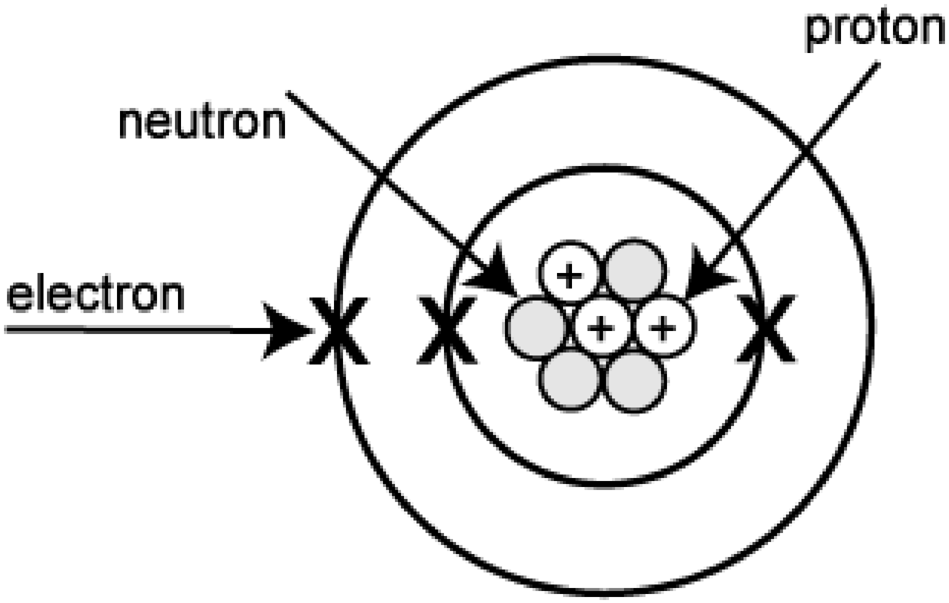


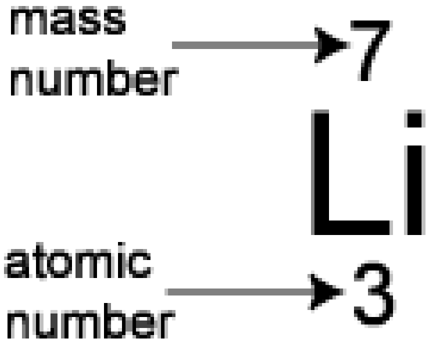
Lesson 1

Atomic structure

What’s in an atom?

An atom is made up of three <u>subatomic particles</u>.





Protons, neutrons and electrons

- ❑ The atomic number (proton number) tells you how many **protons** there are.

- ❑ Since an atom has no overall charge the number of **electrons** (-) is **equal** to the number of **protons** (+).

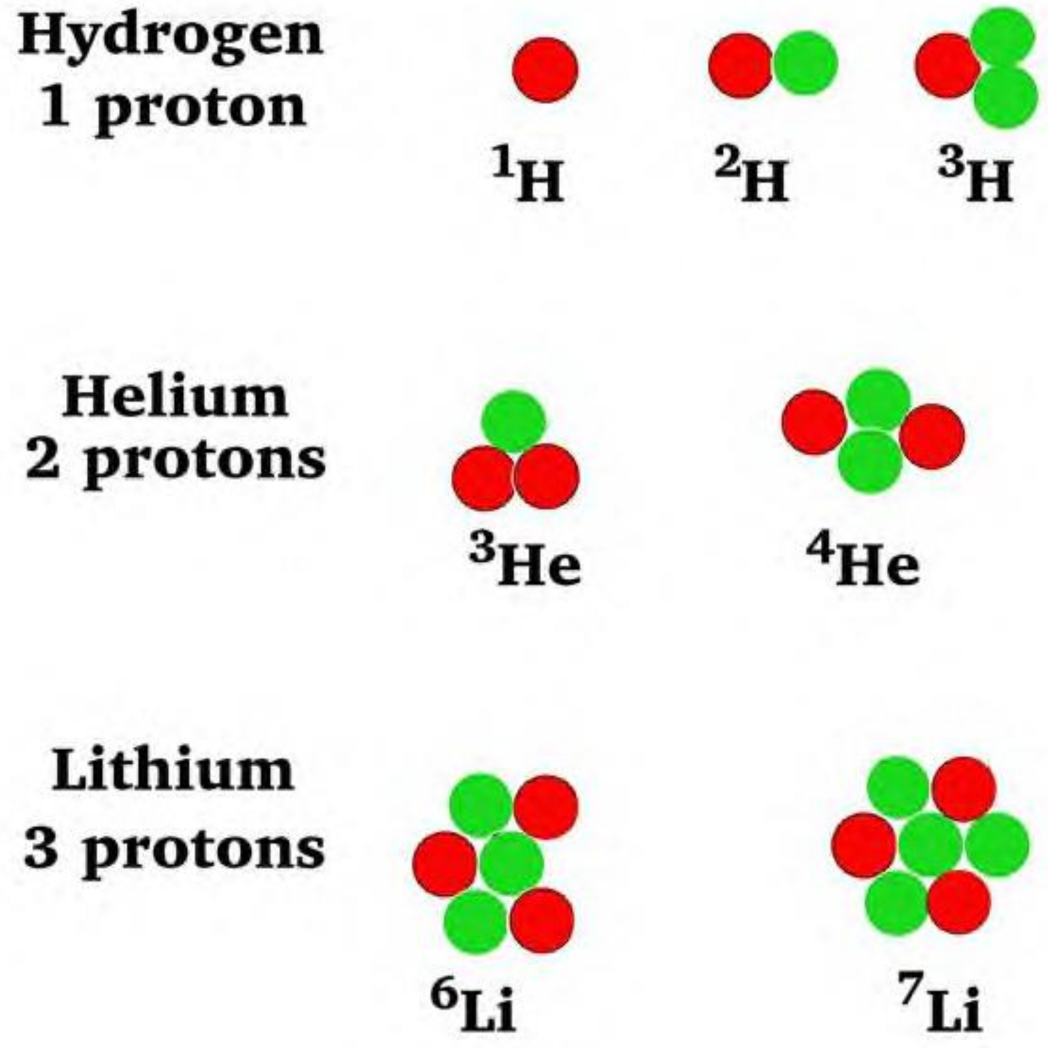
- ❑ The mass number is the total number of protons and neutrons.

- ❑ Since the atomic number tells you how many protons there are, the number of **neutrons** is calculated by:
mass number – atomic number.

Protons, neutrons and electrons

Subatomic particle	Mass	Charge	Position in atom
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus
Electron	0 (1/1800)	-1	Shells or orbits

	Protons	Electrons	Neutrons
$^{7}_{3}Li$			
$^{23}_{11}Na$			
$^{24}_{12}Mg$			
$^{103}_{45}Rh$			



Lesson 2

Isotopes and RAM

What's an isotope?

Many elements are made of atoms which have the same atomic number but a different atomic mass. These atoms are called **isotopes** of the element.

Since the atomic number is the same, the only difference in the structure of the atoms is the number of neutrons.

e.g.,

18 neutrons
17 protons
17 electrons

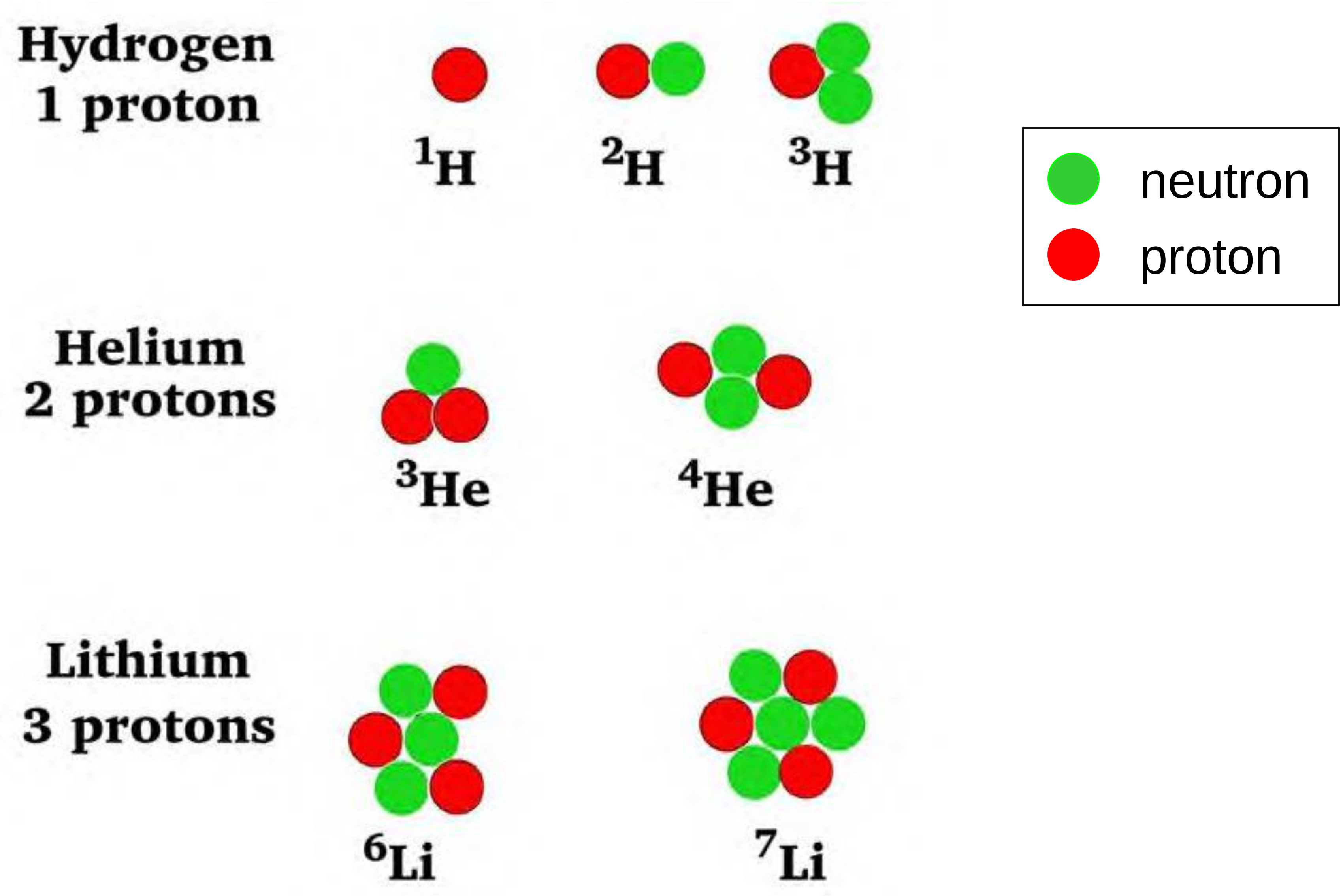
$^{35}_{17}\mathrm{Cl}$

$^{37}_{17}\mathrm{Cl}$

20 neutrons
17 protons
17 electrons

Chlorine has two isotopes. Both are atoms of chlorine because the atomic number is the same.

More isotopes



Relative atomic mass

18 neutrons
17 protons
17 electrons

$^{35}_{17}Cl$

$^{37}_{17}Cl$

20 neutrons
17 protons
17 electrons

Abundance 75 % 25 %

To calculate the relative atomic mass (RAM) you need to know how much (abundance) there is of each isotope.

The fraction of the mass contributed by each isotope is added together.

$$\left(\frac{75}{100} \times 35 = 26.25\right) + \left(\frac{25}{100} \times 37 = 9.25\right) = 35.5$$

more RAM

A sample of lithium contains two isotopes, **P** and **Q**, as shown in the table.

	percentage abundance	relative mass
P	10	6
Q	90	7

What is the relative atomic mass of lithium?

A 6.9
B 6.5
C 6.1
D 1.0

The table shows the relative mass and percentage abundance of two isotopes of gallium, R and S.

isotope	relative mass	percentage abundance
R	69.0	60
S	71.0	40

What is the relative atomic mass of gallium?

A 69.0
B 69.8
C 70.2
D 140.0

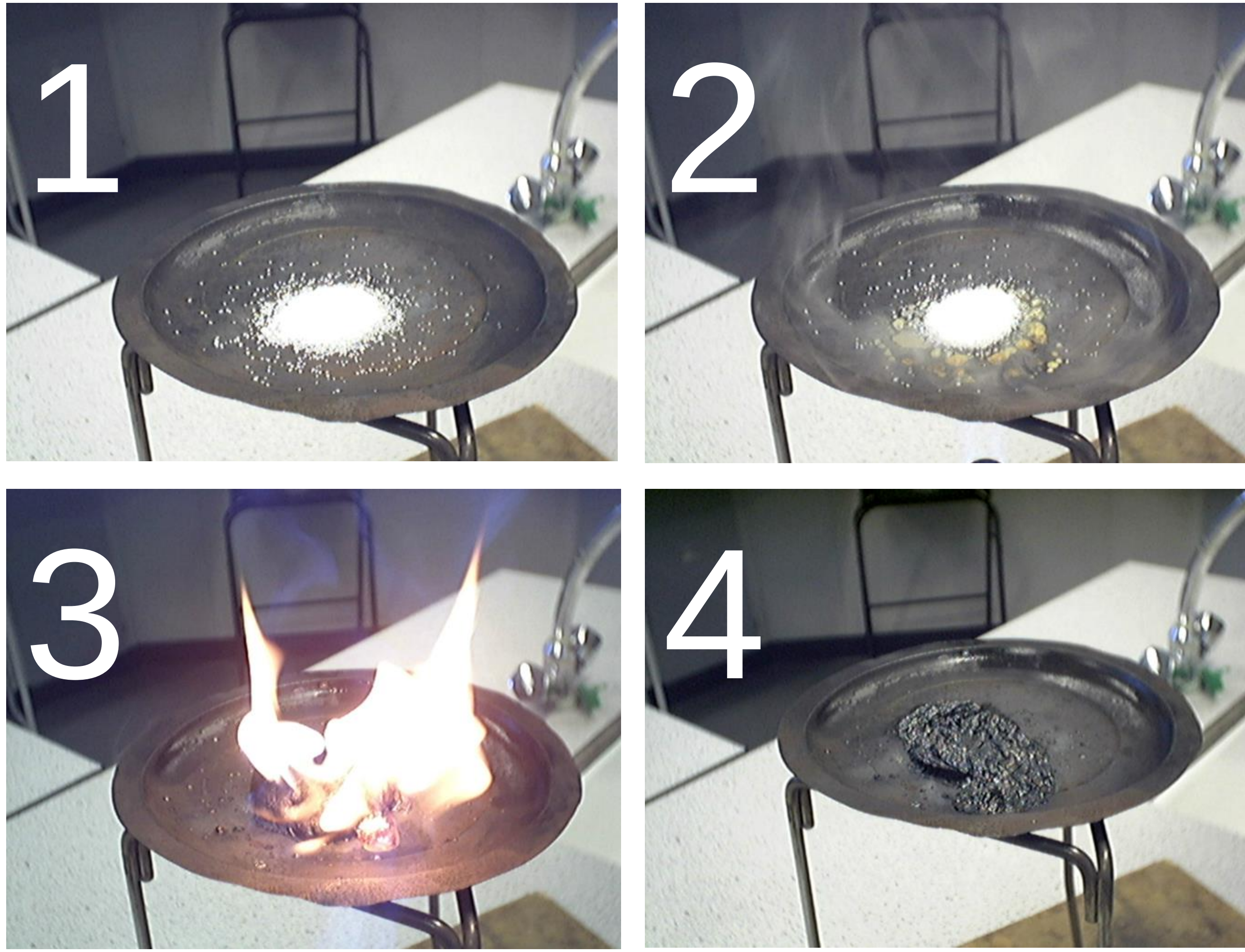
more RAM

A sample of copper consists of 69 % copper-63 and 31 % copper-65.
What is the relative atomic mass of this copper?

A 63.50
B 63.62
C 64.00
D 64.18

The relative atomic mass of chlorine is 35.5.
A sample of chlorine contains atoms of chlorine-35 and chlorine-37.
What is the percentage abundance of each isotope in the sample?

	percentage chlorine-35	percentage chlorine-37
A	35	37
B	75	25
C	50	50
D	25	75



Lesson 3

Chemical reactions

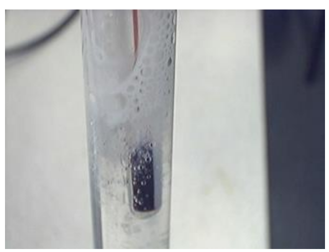
Types of chemical reactions

thermal decomposition
Heat causes a compound to split into new substances

combustion
A reaction in which water and/or carbon dioxide is made

neutralisation
An acid and alkali react to form a salt (and water)

exothermic
Increase in temperature

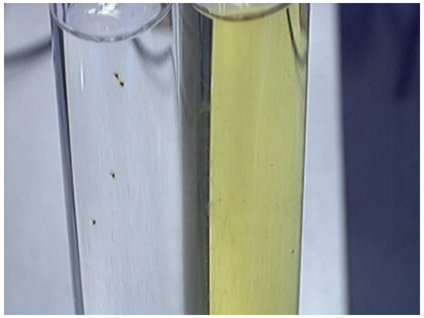


endothermic
Decrease in temperature

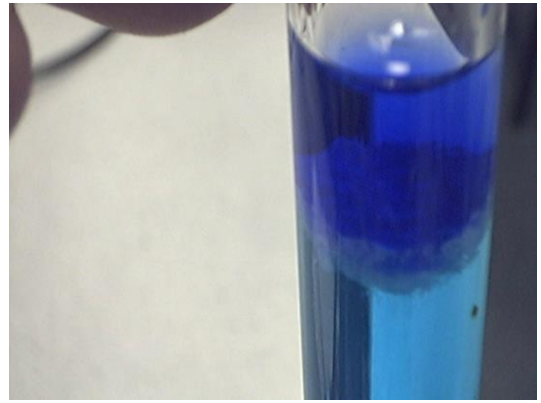
oxidation
Gain of oxygen or loss of electrons

reduction
Loss of oxygen or gain of electrons

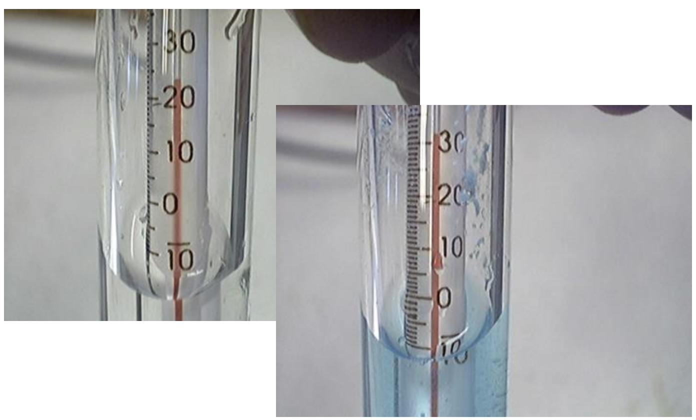
In a reaction.....



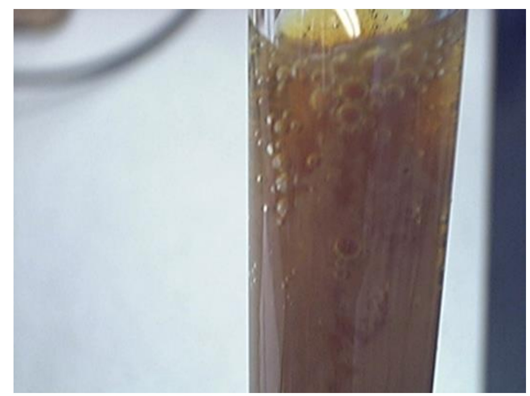
Colour changes



A solid or precipitate forms



A change in temperature



A gas is made (fizzing)

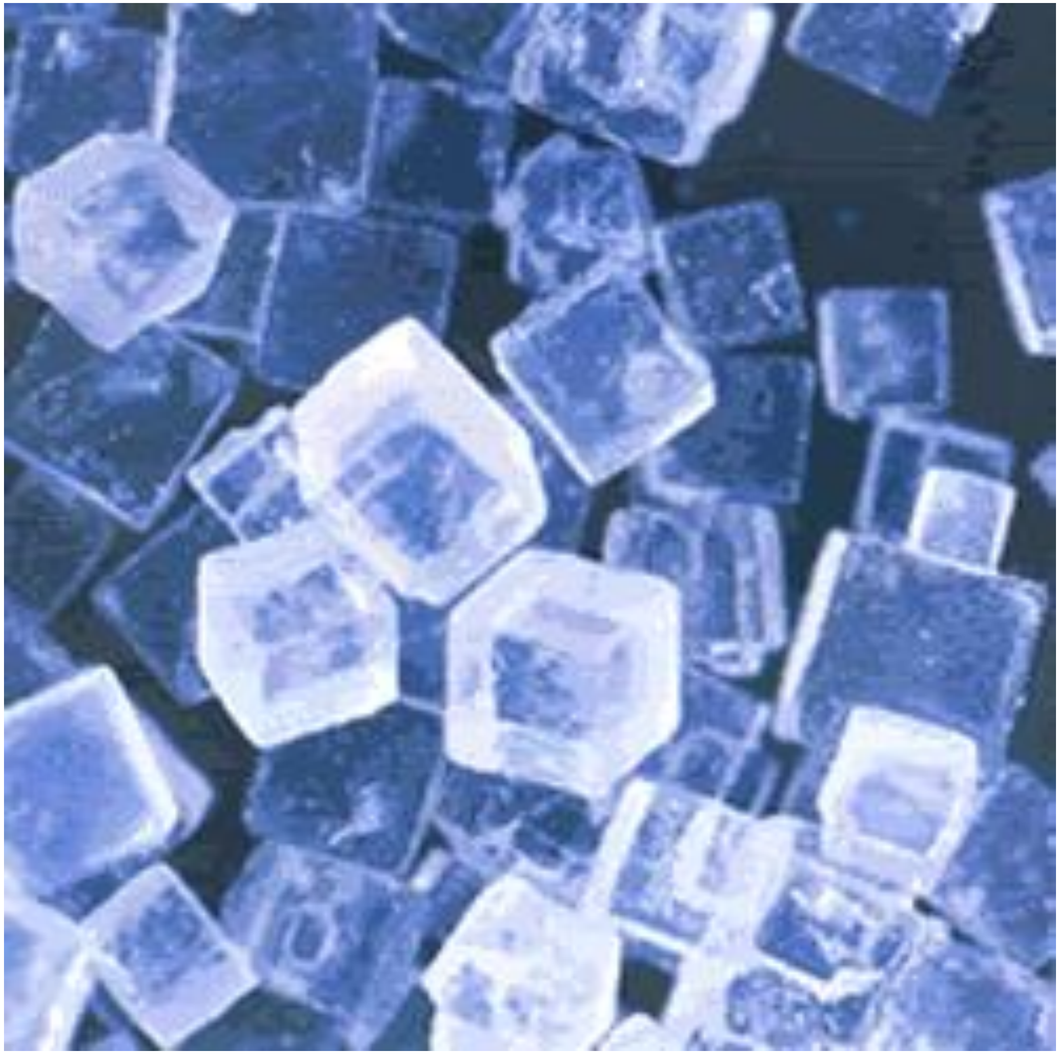
Chemical equations

$$\underbrace{CaCO_{3(s)} + \textcolor{red}{2}HCl_{(aq)}}_{\text{Reactants}} \rightarrow \underbrace{CaCl_{2\,(aq)} + CO_{2\,(g)} + H_2O_{\,(l)}}_{\text{Products}}$$

When two chemicals react the product formed is called a compound. The properties of the reactants and products are all very different.

Sodium (soft metal, very reactive) + **Chlorine** (green, poisonous gas) → **Sodium chloride** (white crystalline solid)

Lesson 4



Ionic bonding

Bonding

- When two atoms of elements come together with enough energy they will react to form a chemical bond. In doing so, each atom is trying to get a full outer shell of electrons.

- To understand bonding you must know how many electrons an atom has in its outer shell. Then using common sense you have to deduce if the atom will loose, gain or share electrons in order to achieve a full outer shell of electrons.

- When atoms gain/lose electrons to form a bond then the atoms become ions, which attract to form an IONIC bond.

Forming ions

- Ions are charged particles – this is shown by a '+' or '-' sign next to the symbol for the ion, e.g., Na^{+}, Cl^{-}, Ca^{2+}.

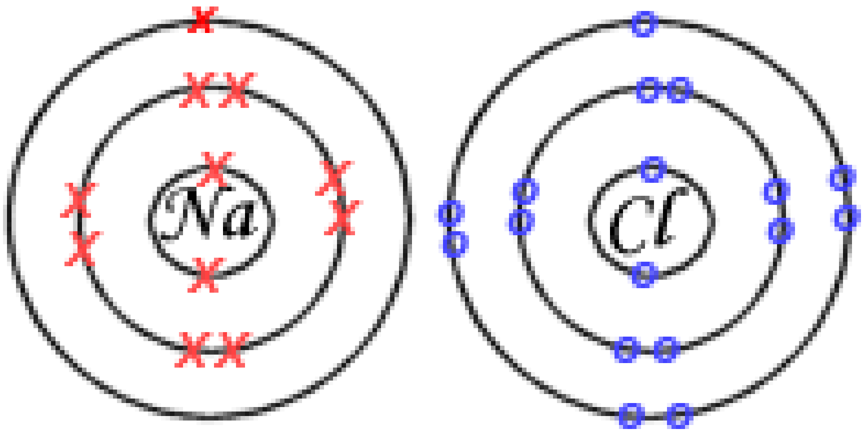
- If the atom loses electrons (negative charges) then it forms a positive ion.

- If the atom gains electrons then it forms a negative ion.

- The small number written next to the '+' or '-' sign tells you how many electrons were lost or gained., e.g., Ca^{2+}, 2 electrons were lost.

- **Ionic compounds form between metal and non-metal atoms.**

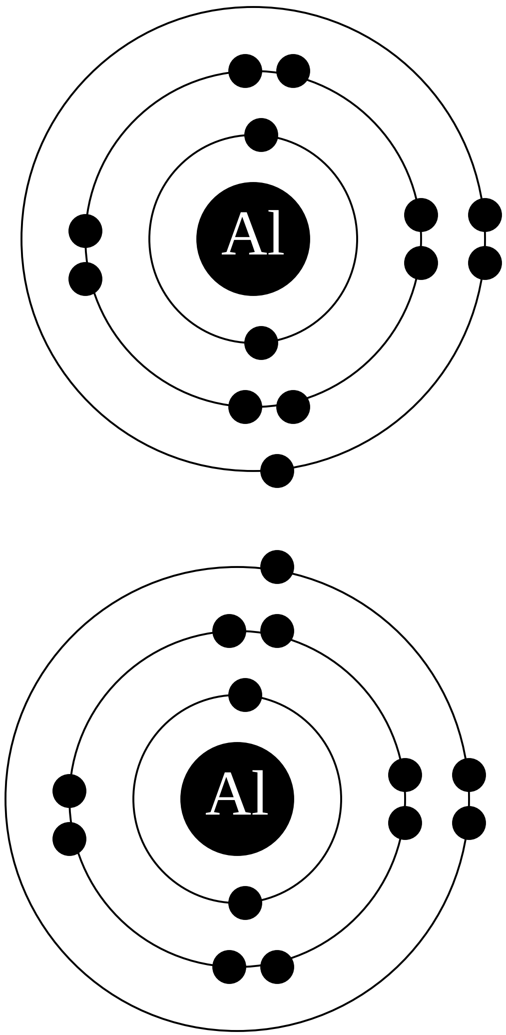
Ions attract to form an ionic bond



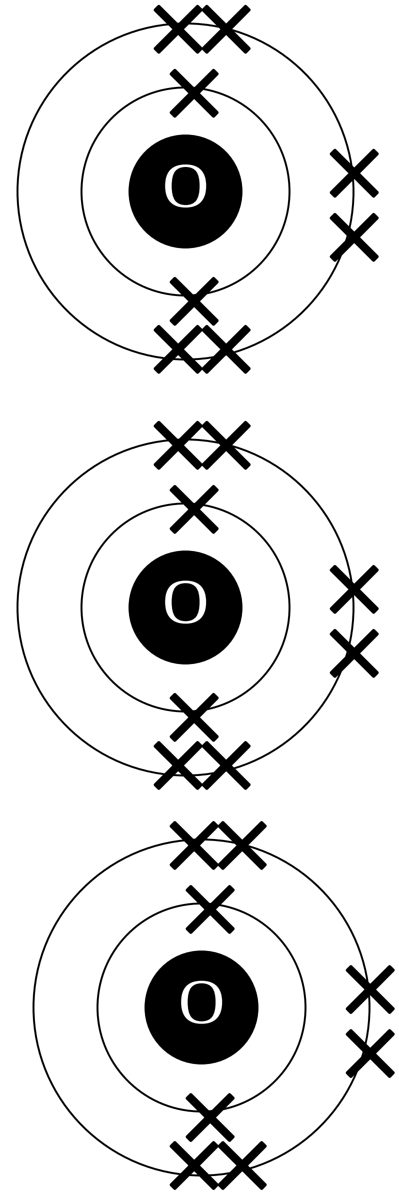
- This is a dot and cross diagram to represent the formation of ions that attract together to form the compound sodium chloride (dots and crosses are electrons).

- Sodium is in Group 1, so it has 1 electron in the outer shell. Chlorine is in Group 7 and has 7 electrons in the outer shell. Sodium loses one electron and chlorine attracts it. This results in the formation of sodium ions (+) and chloride (-) ions. Opposite charges attract and an ionic bond is formed.

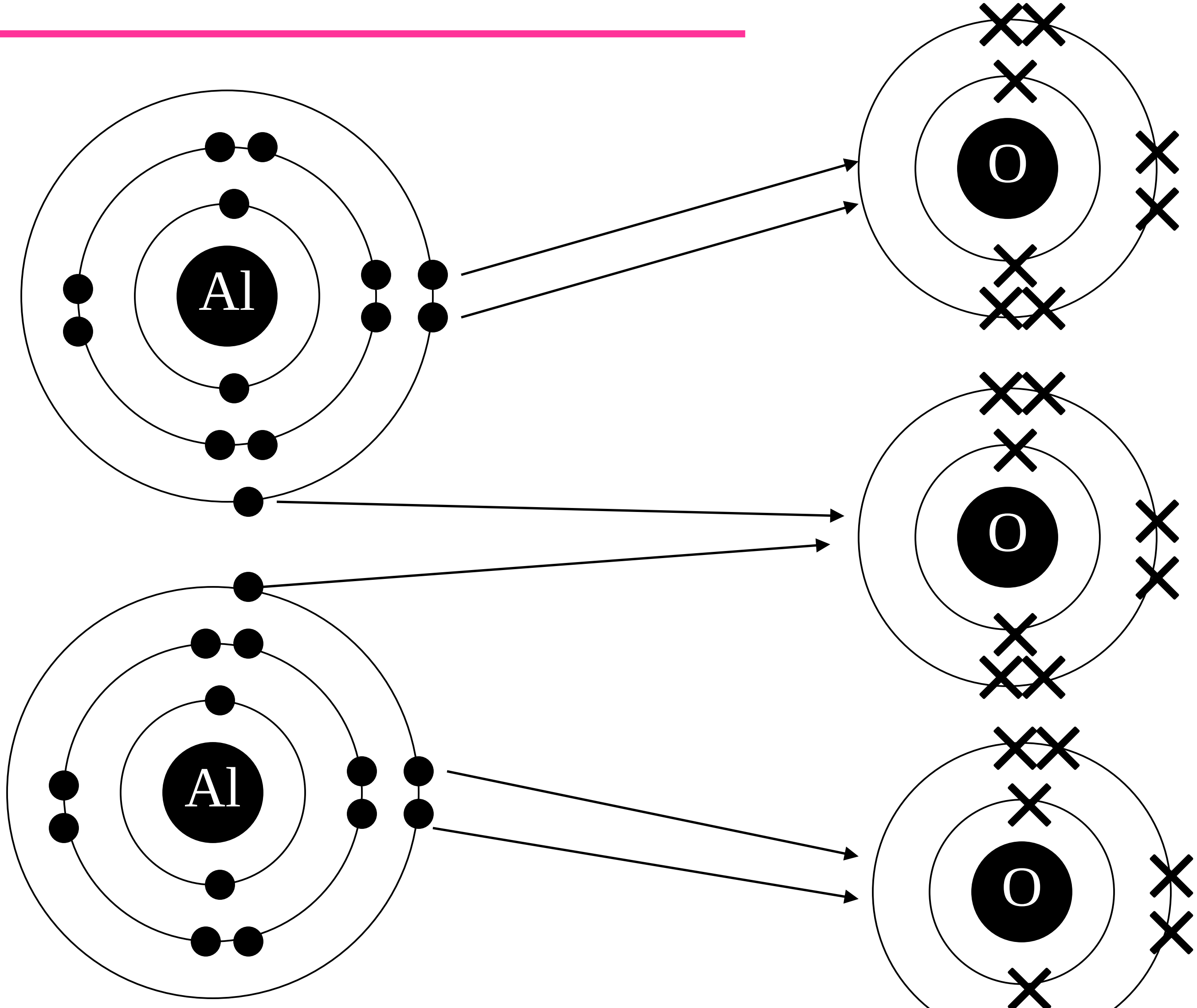
Aluminium oxide



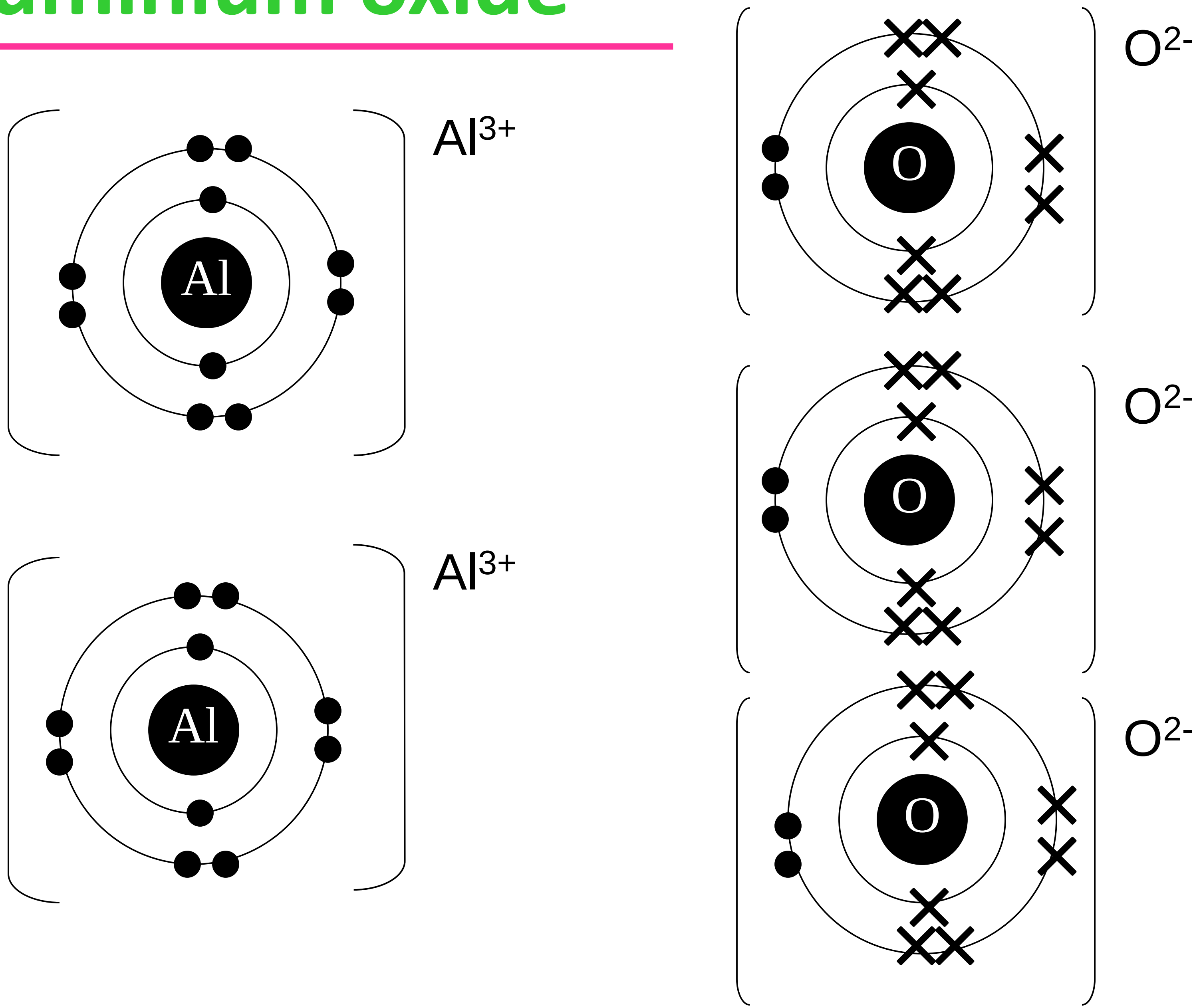
Electrons must have somewhere to go and all compounds have no overall charge. This means that the number of positive charges must balance the number of negative charges.



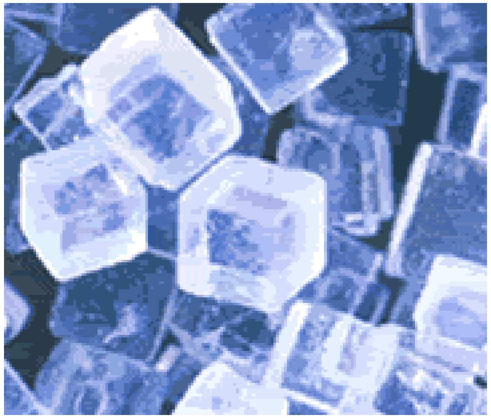
Aluminium oxide

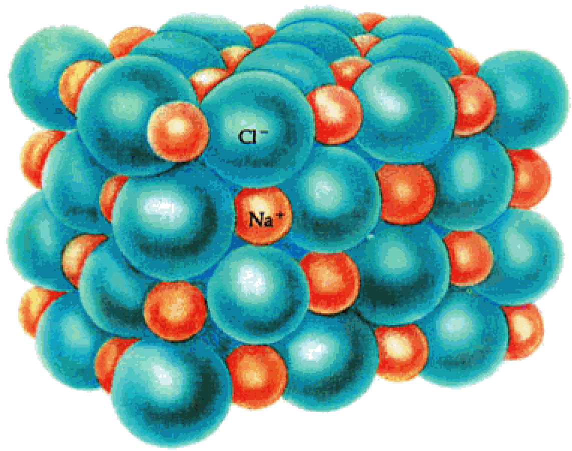


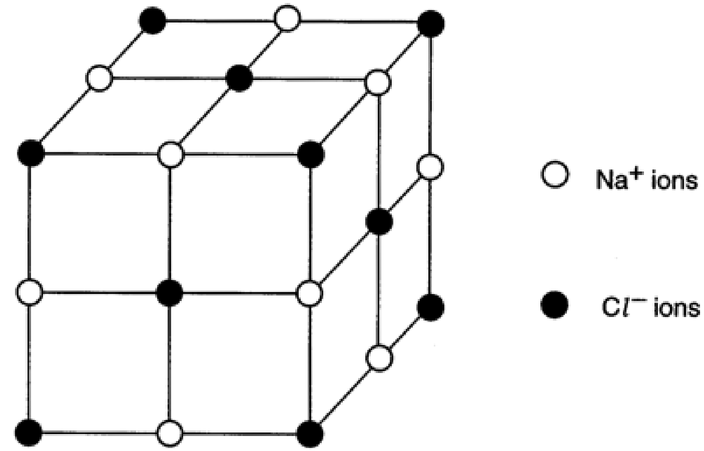
Aluminium oxide



Ionic compounds







- ❑ Negative and positive ions attract to form large crystals. All the ions are held together by strong ionic bonds. NaCl, MgO all from giant lattice structures.

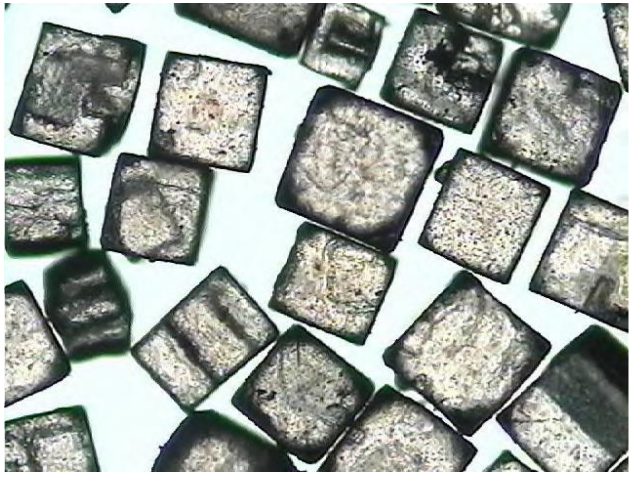
- ❑ The structure is called ‘Giant Ionic’ or ‘Giant lattice’.

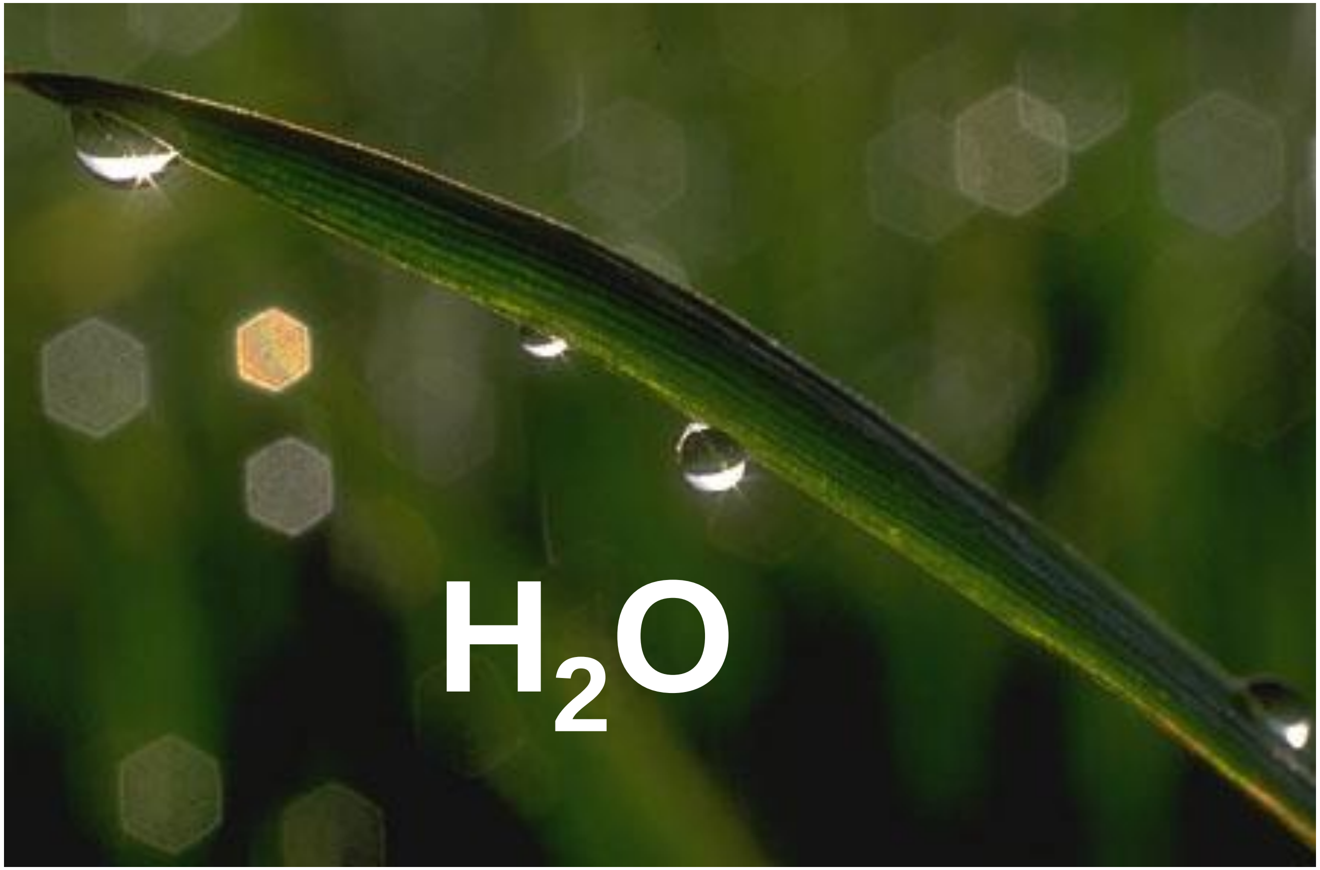
Properties of ionic compounds

- ❑ The giant lattice structure produces crystals which have high melting and boiling points.

- ❑ They dissolve in water.

- ❑ They will conduct electricity but only when dissolved in water or melted. This allows the ions to move about and conduct a current.



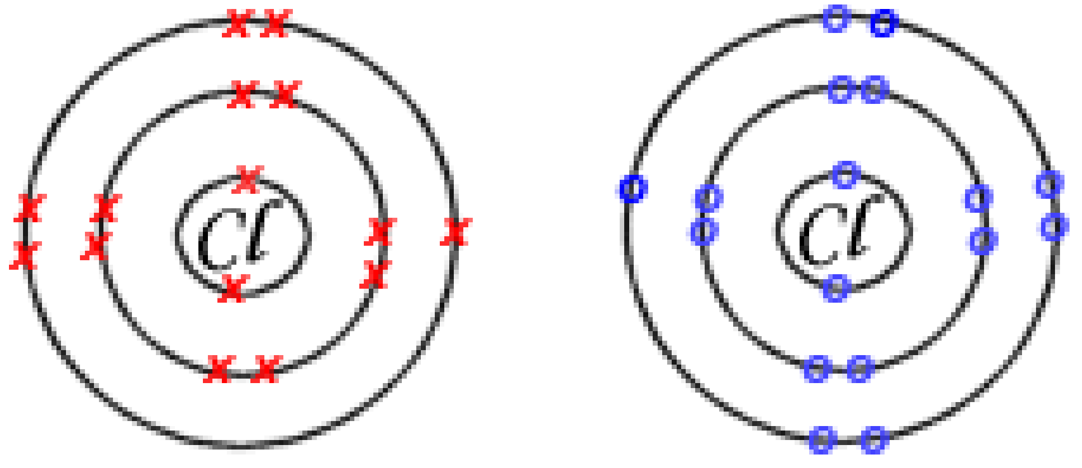


Lesson 5

Covalent bonding

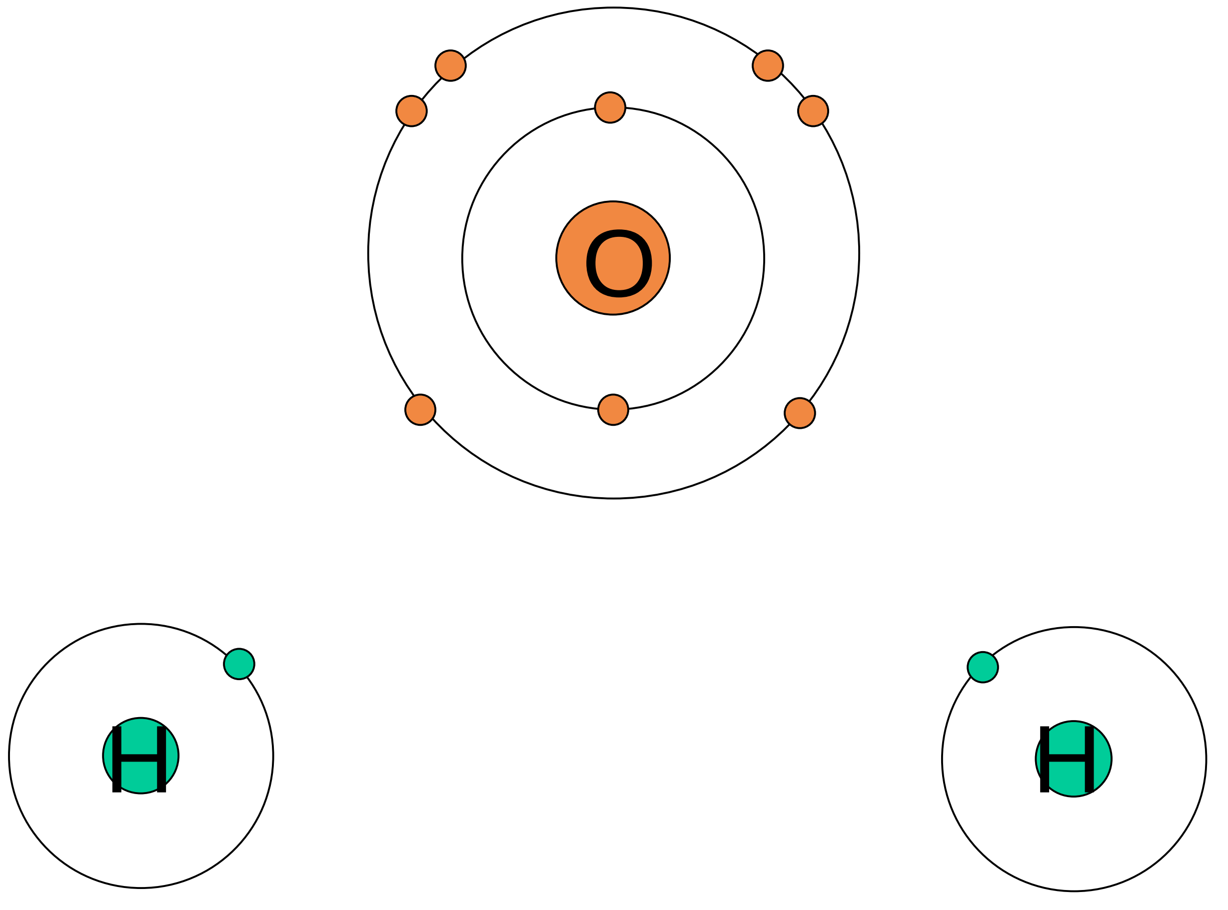
Sharing electrons

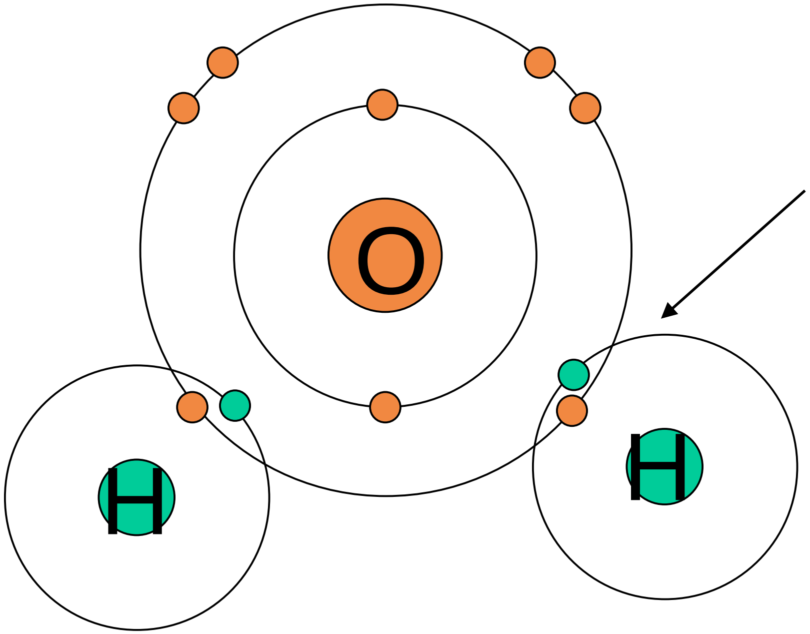
When non-metal atoms form bonds with other non-metal atoms, they have to share electrons. The bond formed is called COVALENT.



The outer shells overlap when the electrons share. A pair of shared electrons forms one covalent bond.

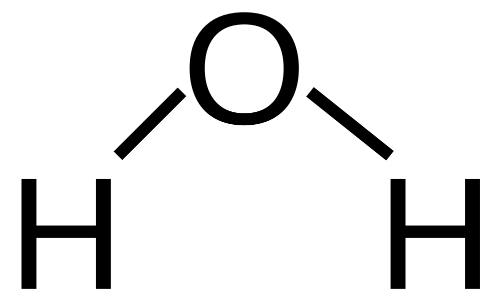
Simple molecules – water



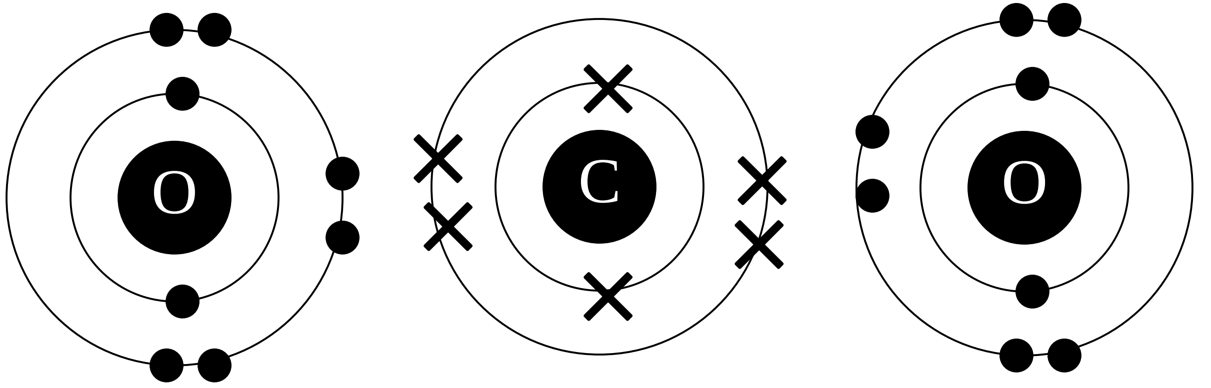


One pair of shared electrons = one covalent bond

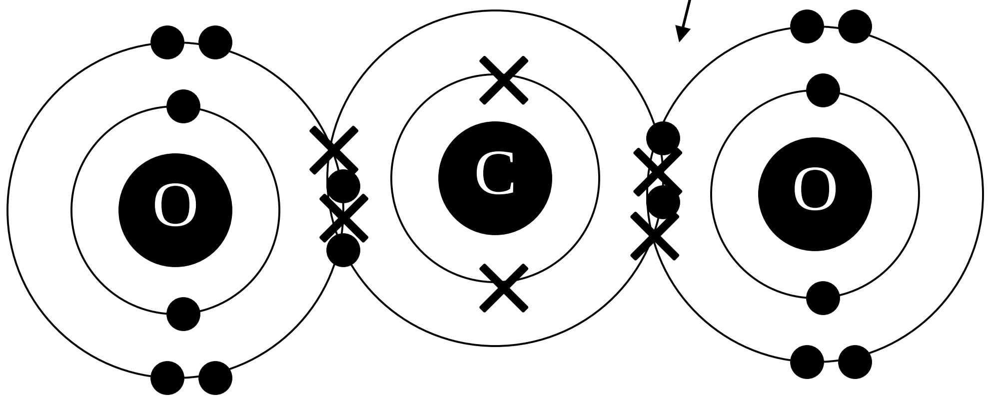
H_2O



Simple molecules – CO_2



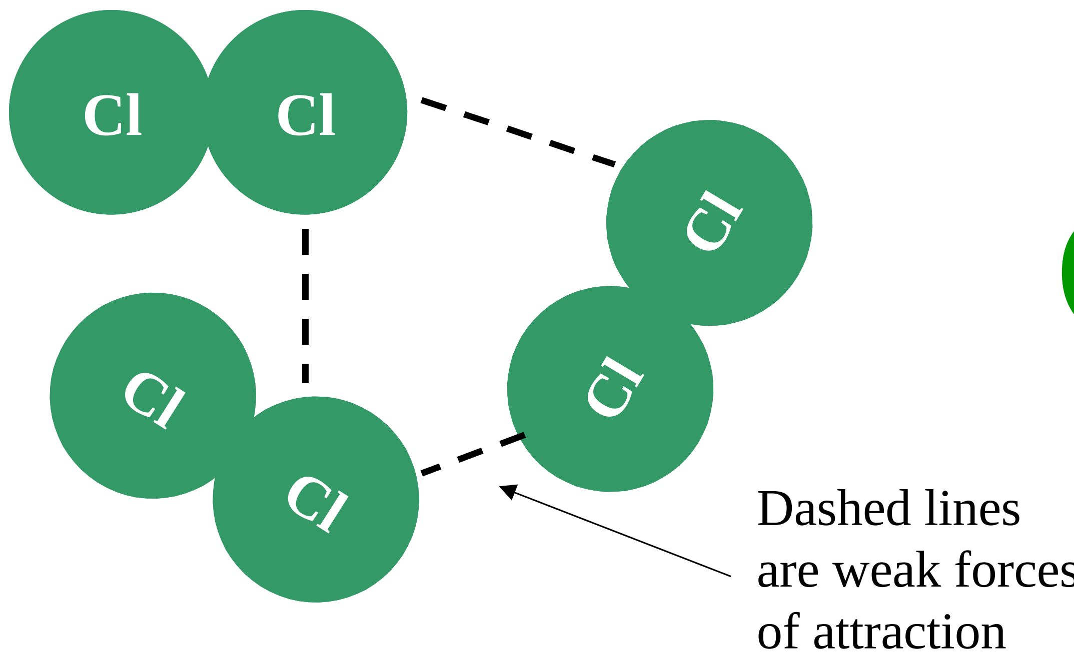
2 pairs of shared electrons produces a double covalent bond



O=C=O

Properties of simple molecules

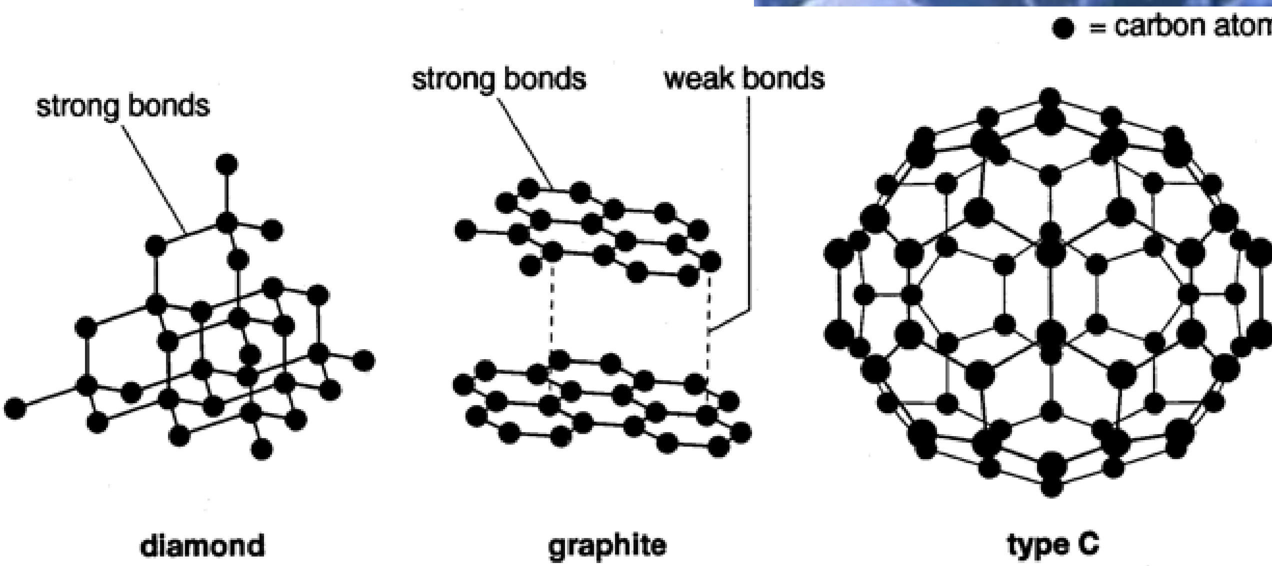
- They have low melting and boiling points because the simple molecules are attracted to each other by weak forces of attraction. To melt or boil covalent molecules you do not break the covalent bonds between atoms. You only separate whole molecules apart.

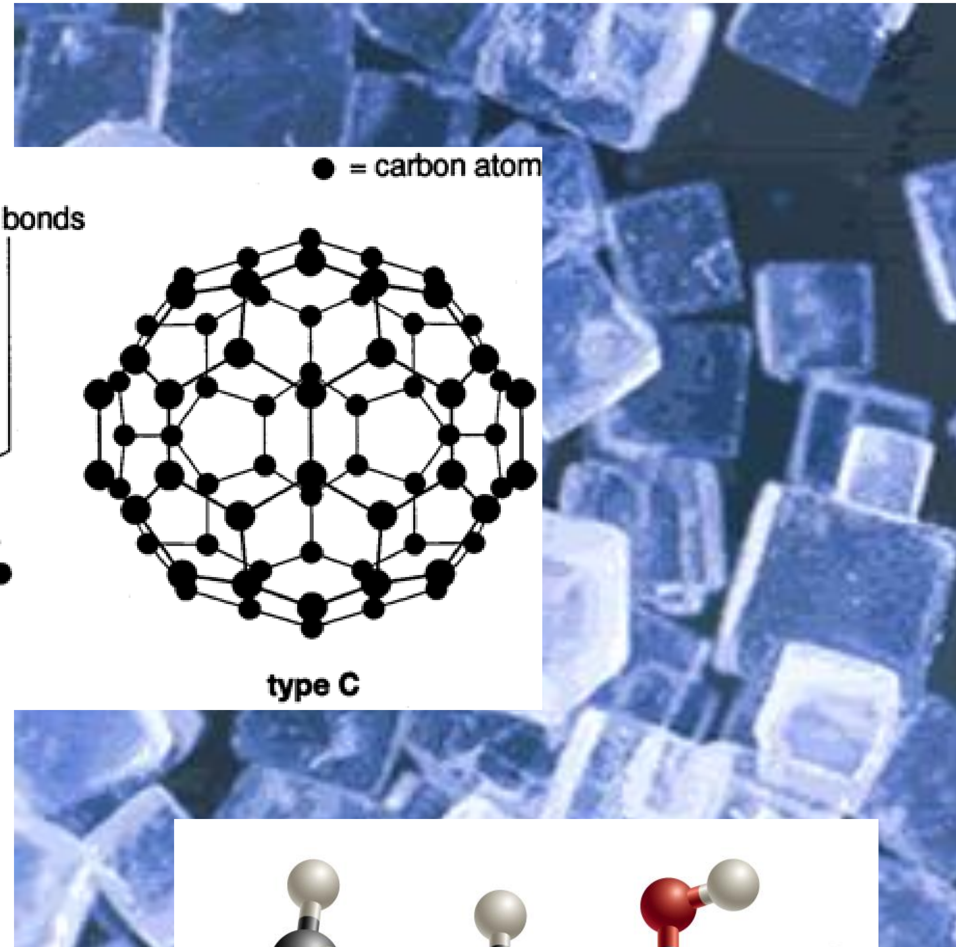


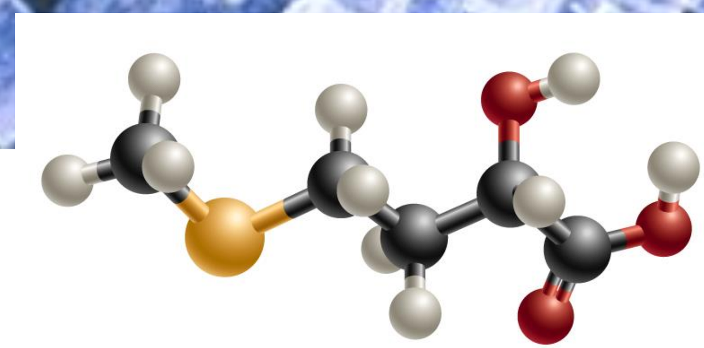
Cl_2

Properties of simple molecules

- Simple covalent molecules are not soluble water.
- They do not conduct electricity.







Lesson 6

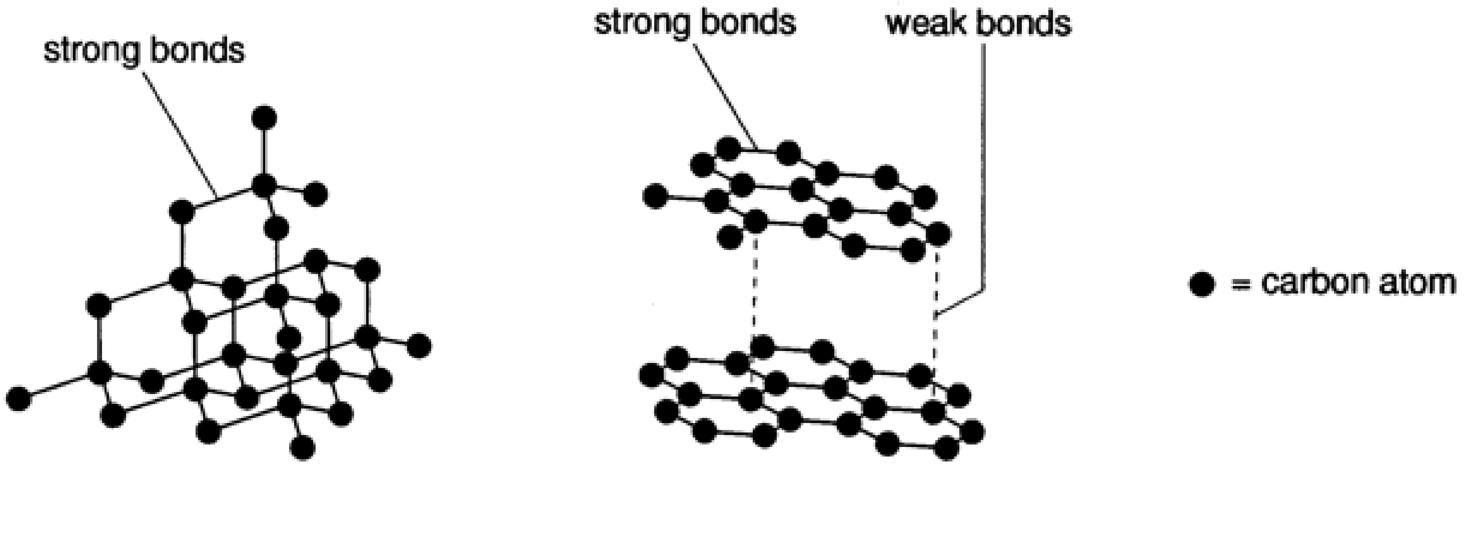
Structures

Types of structures

- Giant Ionic (lattice) – e.g., NaCl

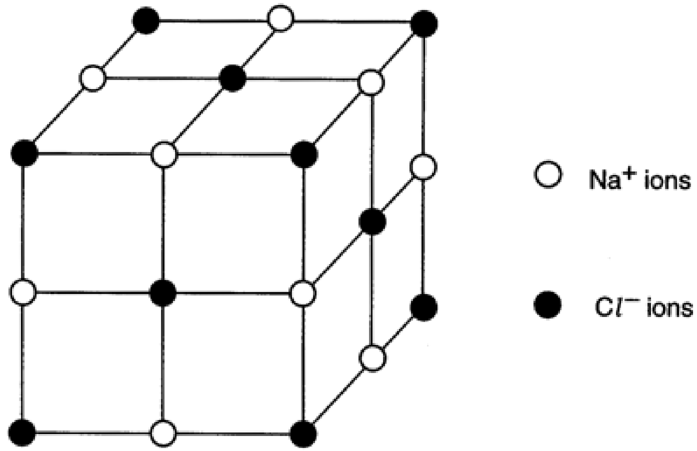
- Simple molecular – e.g., H_2O, CO_2, Cl_2

- Giant molecular – diamond and graphite. Both have high melting and boiling points, but do not dissolve in water. Graphite conducts electricity.



diamond

graphite

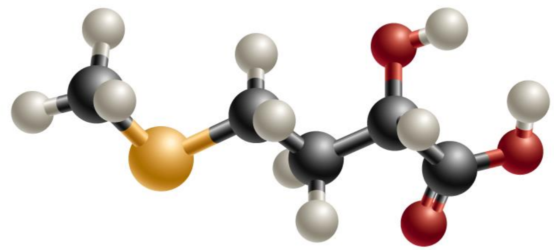


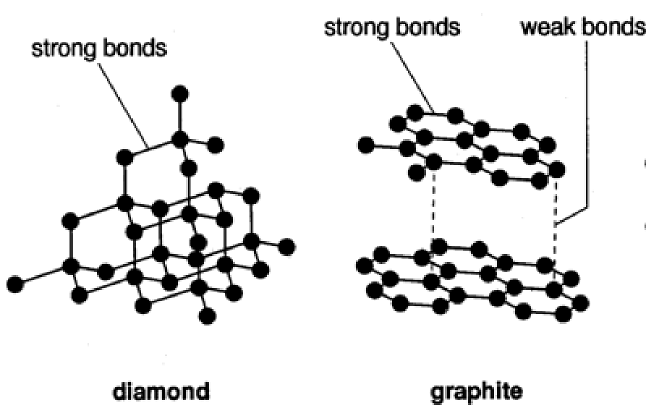
Giant Ionic (lattice)

Large numbers of ions held together by ionic bonds. They have high melting and boiling points, dissolve in water.

Simple molecular

A few atoms bonded together by covalent bonds. Simple molecules have low melting and boiling points, do not dissolve in water or conduct electricity.





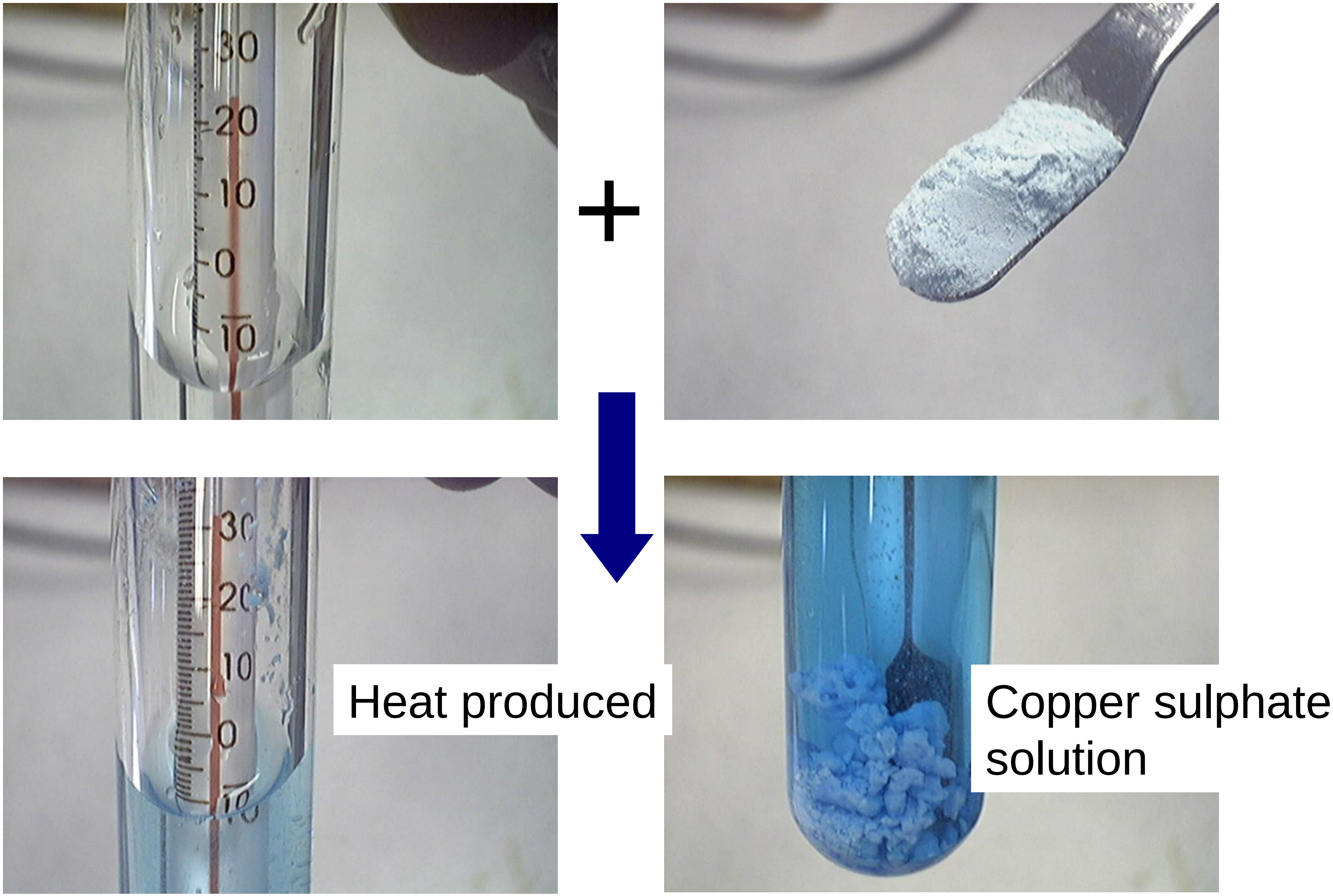
Giant molecular

Large numbers of atoms held together by covalent bonds. They have high melting and boiling points but do not dissolve in water.

Lesson 7

Energy changes in chemical reactions

Most reactions produce heat



Energy changes in reactions

Exothermic reactions

- Heat energy is produced in these reactions.
- Temperature increases.

Endothermic reactions

- Heat energy is taken in these reactions so the surroundings cool down.
- Temperature decreases.

Very few reactions are endothermic

Bond breaking and bond making

- All chemical reactions involve the breaking of bonds between atoms and the forming of new bonds.

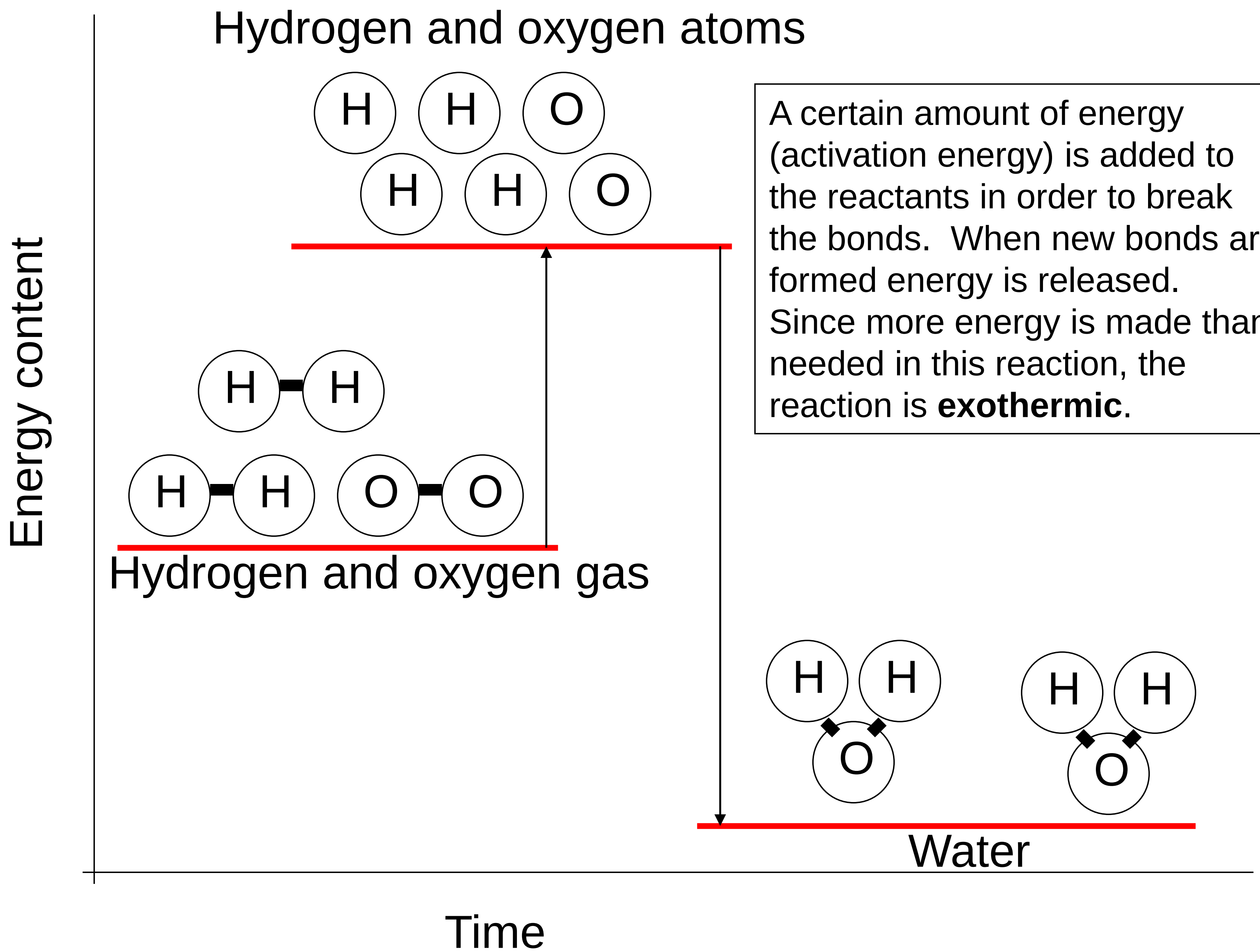
- Energy is needed to break bonds.

- Energy is made when new bonds are formed.

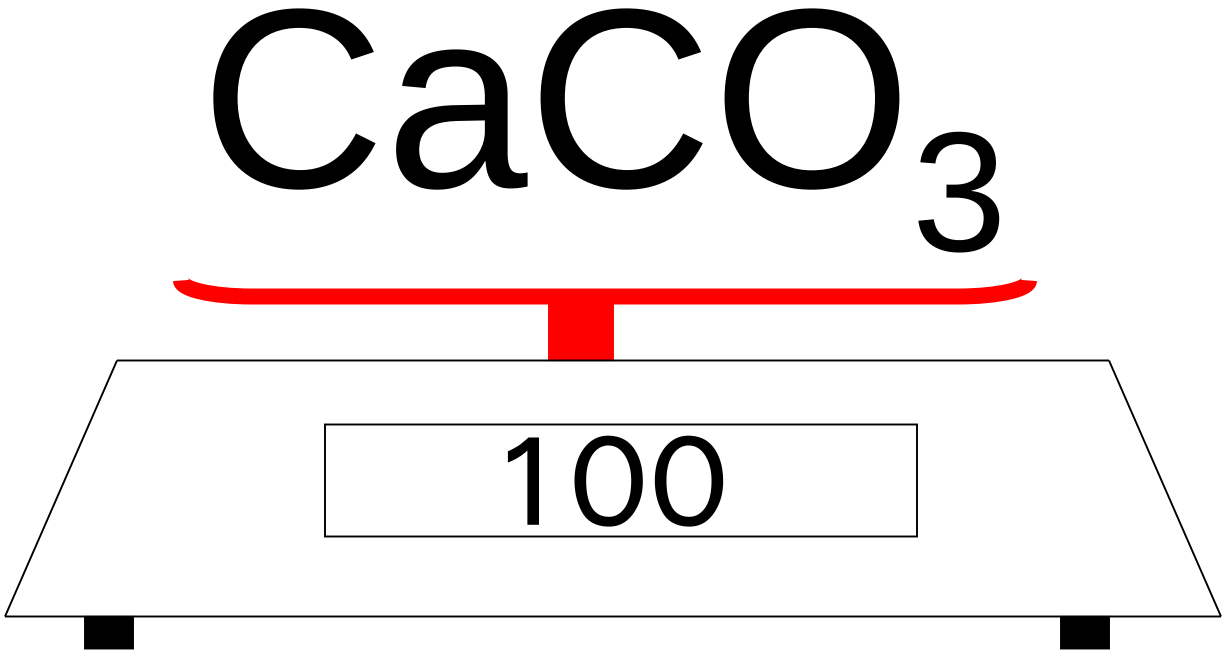
- If more energy is made when new bonds are formed than is needed when bonds are broken then the reaction is exothermic.

Example

	Energy needed to break bonds	Energy made when new bonds formed
Exothermic reaction	1000 kJ	5000 kJ
Endothermic reaction	5000 kJ	1000 kJ



A certain amount of energy (activation energy) is added to the reactants in order to break the bonds. When new bonds are formed energy is released. Since more energy is made than needed in this reaction, the reaction is **exothermic**.



Lesson 8

Relative formula mass

The **relative formula mass** is the sum of the atomic masses of all the atoms in the chemical formula. Just add up the mass numbers!!

What is the relative molecular mass of phosphorus(III) fluoride, PF_3?
(Relative atomic masses: F = 19; P = 31)

A	50
$\boxed{\textbf{B}}$	88
C	122
D	150

3 fluorine atoms

$$31 + (3 \times 19) = \mathbf{88}$$

What is the relative formula mass of magnesium oxide, MgO?
(Relative atomic masses: $O = 16$, $Mg = 24$)

A	8
B	20
C	40
D	384

What is the relative formula mass of methane, CH_4?
(Relative atomic masses: $H = 1$; $C = 12$)

A	4
B	12
C	16
D	48

What is the relative formula mass of carbon dioxide, CO_2?
(Relative atomic masses: $C = 12$; $O = 16$)

A	28
B	40
C	44
D	56

$$C_nH_{2n}$$

Lesson 9

Empirical formulae

The empirical formulae is the simplest ratio of atoms in the chemical formula of a compound.

An oxide of copper contains 32 g of copper and 4 g of oxygen.
What is its empirical formula?
(Relative atomic masses: O = 16; Cu = 64)

A CuO_2
B CuO
$\boxed{C}$ Cu_2O
D Cu_4O

$$\frac{\text{Mass in grams}}{\text{Relative atomic mass}}$$

$$\frac{32}{64} : \frac{4}{16}$$

$$= 0.5 : 0.25$$

$$= 2 : 1$$

2 Cu atoms

1 O atom

Divide the mass in grams by the relative atomic mass for each element. Then calculate the simplest ratio of the atoms.

An oxide of sodium contains 4.6 g of sodium and 1.6 g of oxygen.
What is its empirical formula?
(Relative atomic masses: O = 16; Na = 23)

A Na_2O
B Na_2O_2
C NaO
D NaO_2

2.4 g of magnesium reacted with 1.6 g of oxygen.
What is the empirical formula of the compound formed?
(Relative atomic masses: O = 16: Mg = 24)

A Mg_2O
B MgO
C Mg_2O_3
D MgO_2

$$\text{Moles} = \frac{\text{mass}}{\text{RAM}}$$

Lesson 10

Quantitative chemistry

Magnesium displaces iron from iron(II) sulphate solution.

$$Mg + FeSO_4 \longrightarrow MgSO_4 + Fe$$

$$\boxed{\text{Moles} = \frac{\text{mass}}{\text{RAM}}}$$

What mass of iron is formed from 12 g of magnesium?
(Relative atomic masses: Mg = 24; Fe = 56)

A 7 g
B 14 g
C 28 g
D 56 g

$$\text{Moles} = \frac{12}{24} = 0.5$$

Steps in calculation:

1. Calculate the number of moles of the substance for which you are given the mass data.
2. In the above example, 1 mole of magnesium forms 1 mole of iron. Therefore, 0.5 moles of magnesium produces 0.5 moles of iron.
3. Now that you have the moles of iron, you can calculate the mass produced by multiplying the moles of iron by the relative atomic mass of iron. Mass = 0.5 x 56 = 28 g.

4. The answer is **28 g**.

Copper carbonate decomposes to form copper oxide and carbon dioxide.

$$CuCO_3 \rightarrow CuO + CO_2$$

What mass of carbon dioxide is formed when 12.4 g of copper carbonate decomposes?
(Relative atomic masses: C = 12, O = 16, Cu = 64)

A 2.5 g
B 3.8 g
C 4.4 g
D 5.6 g

You will have to calculate the RFM for the compounds.

$$\text{Moles} = \frac{\text{mass}}{\text{RFM}}$$

When potassium chlorate is heated it forms potassium chloride.

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

What mass of potassium chloride is formed when 122.5 g of potassium chlorate decomposes completely?
(Relative atomic masses: O = 16; Cl = 35.5; K = 39)

A 74.5 g
B 96 g
C 122.5 g
D 149 g

Iron is produced by reducing iron(III) oxide with carbon monoxide.

$$Fe_2O_3 \quad + \quad 3CO \longrightarrow 2Fe \quad + \quad 3CO_2$$

What mass of iron is formed when 80 g of iron(III) oxide react completely?
(Relative atomic masses: O = 16; Fe = 56)

A 160 g
B 112 g
C 56 g
D 28 g