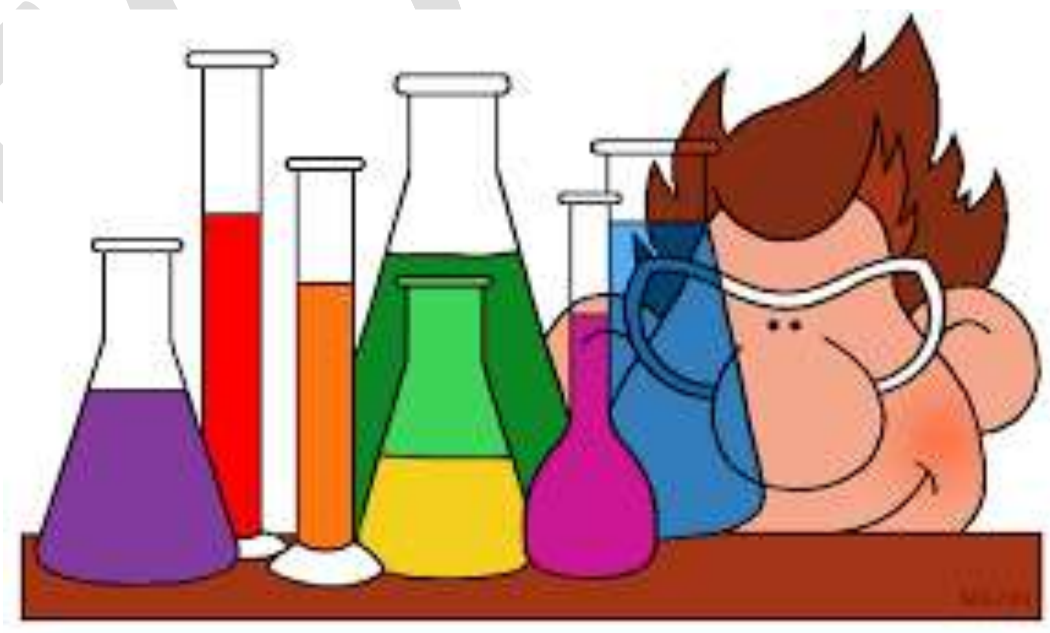
Edexcel O-L

Chemistry

(Code: 4CH1)

Unit 1
Principles of Chemistry

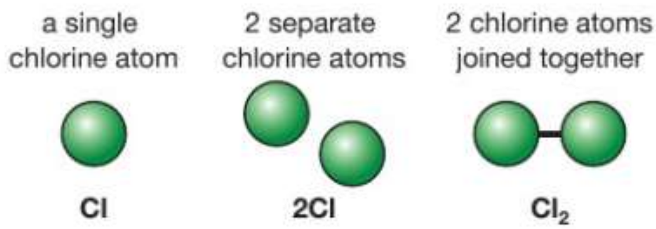


1.5 Chemical formulae, equations and calculations: part 1

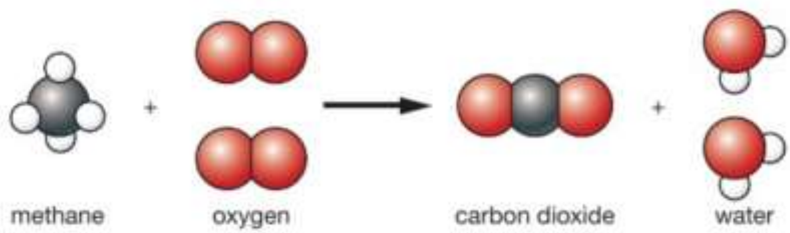
Writing equations

There are two types: word equations and symbol equations (chemical equations).

What all the numbers mean



Balancing equation



In the figure there are the same number of each type of atom before and after. This is called a *balanced equation*.

How to balance equations

EXAMPLE 1

Balance the equation for the reaction between zinc and hydrochloric acid:

$$Zn + HCl \rightarrow ZnCl_2 + H_2$$

Work from left to right. Count the zinc atoms: 1 on each side; no problem!

Count the hydrogen atoms: 1 on the left, 2 on the right. If you have 2 at the end, you must have started with 2. The only way of achieving this is to have 2HCl. (You must not change the formula to H_2Cl because this substance does not exist.)

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

Now count the chlorines: there are 2 on each side. Good! Finally check everything again to make sure and you've finished.

	$Zn + 2HCl$		$\rightarrow$	$ZnCl_2 + H_2$	
numbers of atoms	Zn	1		Zn	1
	H	2		H	2
	Cl	2		Cl	2

State symbols

The symbols are: (s) – Solid, (l) – Liquid, (g) – Gas, (aq) – Aqueous solution

HINT

Remember that water is a liquid (l), not an aqueous solution (aq). An aqueous solution is formed when something is *dissolved in water.*

So an equation may look like this:

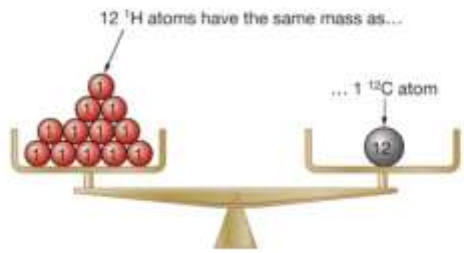
$$2K_{(s)} + 2H_2O_{(l)} \rightarrow 2KOH_{(aq)} + H_{2\,(g)}$$

This shows that *solid* potassium reacts with *liquid* water to make a *solution* of potassium hydroxide in water and hydrogen *gas*.

How much of each substance reacts in a chemical reaction?

You can get the right proportions only if you know about the masses of the individual atom that take part in the reaction.

Relative atomic mass (A_r)



The mass of hydrogen atom is about 1.67×10^{-24} g.
The mass of atoms (and molecules) are compared with the mass of an atom of the *carbon-12 isotope*. We call this the carbon-12 scale.

The basic unit on the carbon-12 scale is of the mass of a ^{12}C atom, which is approximately the mass of the most common hydrogen atom.

The relative atomic mass of an element is the weighted average mass of the isotopes of the element. It is measured on a scale on which a carbon-12 (^{12}C) atom has a mass of exactly 12.

Relative atomic mass has no units.

The relative mass of; Fluorine is 19, the most common isotope of Magnesium 24, Chlorine 35.45 (35.5), Lithium 6.94 (7), Sodium 22.99 (23).

Relative formulae mass(M_r)

Sometimes called relative molecular mass.

To find the relative formula mass (M_r) of magnesium carbonate, $MgCO_3$

Relative atomic masses: C = 12, O = 16, Mg = 24.

Add up the relative atomic masses to give the relative formula mass of the whole compound. In this case, you need to add up the masses of 1 × Mg, 1 × C and 3 × O.

$$M_r = \underset{Mg}{24} + \underset{C}{12} + \underset{(3 \times O)}{(3 \times 16)}$$

$$M_r = 84$$

To find the relative formula mass of calcium hydroxide, $Ca(OH)_2$

Relative atomic masses: H = 1, O = 16, Ca = 40.

$$M_r = \underset{Ca}{40} + (\underset{O}{16} + \underset{H}{1}) \times 2$$

$$M_r = 74$$

<u>Water of crystallization</u>

When some substances crystallize from solution, water becomes chemically bound up with the salt.

<u>Hydrated</u>

Salts containing water of crystallization.

Using relative formulae mass to find percentage composition

To find the percentage by mass of copper in copper(II) oxide, CuO

Relative atomic masses: O = 16, Cu = 63.5.

M_r of CuO = 63.5 + 16 = 79.5

Of this, 63.5 is copper.

$$\text{Percentage of copper} = \frac{63.5}{79.5} \times 100$$
$$= 79.9\%$$

To find the percentage by mass of oxygen in sodium carbonate, Na_2CO_3

Relative atomic masses: Na = 23, C = 12, O = 16.

$$M_r = (2 \times 23) + 12 + (3 \times 16)$$
$$= 106$$

48 of this 106 is due to the oxygen (remember that there are 3 oxygens in the formula, so the total mass of oxygen is 3 × 16 = 48).

$$\text{Percentage of oxygen} = \frac{3 \times 16}{106} \times 100$$
$$= 45.3\%$$

The mole

The unit of amount of substance. One mole of any substance as its own particular mass. The amount of moles and number of moles are considered to be two different things.

Calculating the masses of a mole of substance

1 mole of oxygen gas, O_2

Relative atomic mass: O = 16.

$$M_r \text{ of } O_2 = 2 \times 16$$
$$= 32$$

1 mole of oxygen, O_2, has a mass of 32 g.

1 mole of calcium chloride, $CaCl_2$

Relative atomic masses: Cl = 35.5, Ca = 40.

$$M_r \text{ of } CaCl_2 = 40 + (2 \times 35.5)$$
$$= 111$$

1 mole of calcium chloride has a mass of 111 g.

1 mole of iron(II) sulfate crystals, $FeSO_4{\cdot}7H_2O$

Relative atomic masses: H = 1, O = 16, S = 32, Fe = 56.

$$M_r \text{ of crystals} = 56 + 32 + (4 \times 16) + \{7 \times [(2 \times 1) + 16]\}$$

1 mole of iron(II) sulfate crystals has a mass of 278 g.

Simple calculations with moles

$$number\ of\ moles = \frac{mass(g)}{mass\ of\ one\ mole(g)}$$

Finding the mass of 0.2 mol of calcium carbonate, $CaCO_3$

Relative atomic masses: C = 12, O = 16, Ca = 40.

First find the relative formula mass of calcium carbonate.

$$M_r \text{ of } CaCO_3 = 40 + 12 + (3 \times 16) = 100$$

1 mol of $CaCO_3$ has a mass of 100 g.

$$\text{Mass (g)} = \text{number of moles} \times \text{mass of 1 mole (g)} = 0.2 \times 100 = 20\,g$$

0.2 mol of $CaCO_3$ has a mass of 20 g.

Finding the number of moles in 54 g of water, H_2O

Relative atomic masses: H = 1, O = 16.

1 mol of H_2O has a mass of 18 g.

$$\text{Number of moles} = \frac{\text{mass (g)}}{\text{mass of 1 mole (g)}} = \frac{54}{18} = 3\,\text{mol}$$

54 g of water is 3 mol.

HINT

We often use the term 'molar mass' instead of 'mass of 1 mole'.

Formulae

The *empirical formula* shows the simplest whole number ratio of the atoms present in a compound. The *molecular formula* shows the actual number of atoms of each element present in a molecule (covalent compound) or formula unit (ionic compound) of a compound.

KEY POINT

How do we know when the reaction has finished? We can generally see when there is no more reaction because nothing is happening in the crucible. We can also check by weighing the crucible and contents, then heating it and weighing again. If the mass is higher when we weigh it again then this indicates that a reaction has occurred. We keep heating and weighing until the mass stays constant.

Working out empirical formulae

EXAMPLE 3

A sample of a compound contains 1.27 g of Cu and 0.16 g of O. Calculate the empirical formula. (A_r of Cu = 63.5, A_r of O = 16)

It is easiest to do this in columns using a table:

	Cu	O
masses/g	1.27	0.16
find the number of moles of atoms by dividing the mass by the mass of 1 mole of atoms	1.27/63.5	0.16/16
number of moles of atoms	0.02	0.01
divide by the smaller number to find the ratio	0.02/0.01	0.01/0.01
ratio of moles	2	1
empirical formula	Cu_2O	

From calculating the number of moles of Cu and O we can see that there are twice as many moles of Cu and therefore there must be twice as many Cu atoms as O atoms in the compound.

It is important to remember in this calculation that we are working out how many *atoms* of copper combine with how many *atoms* of oxygen so we divide the mass of oxygen by 16 (the mass of 1 mole of oxygen atoms) rather than by 32 (the mass of 1 mole of oxygen molecules).

KEY POINT

Remember: Number of moles is mass in grams divided by the mass of 1 mole in grams.

The overall reaction that occurred was:

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

magnesium + oxygen → magnesium oxide

The type of reaction that occurred was combustion.

When we do this experiment in practice, it is not very often that the ratio comes out as 1:1. There are things that may go wrong with this experiment which could affect the results, such as:

- some of the magnesium oxide powder might escape when the lid is lifted
- not all the magnesium might have reacted
- the magnesium can also react with nitrogen in the air.

The formula for copper oxide

KEY POINT

We must always make sure that the tube is filled with hydrogen before lighting the stream of hydrogen because a mixture of hydrogen and oxygen is explosive! It is therefore important to let the stream of hydrogen gas flow through the tube for a little while (to flush out all the oxygen) before lighting it or lighting the Bunsen burner.

KEY POINT

The equation for the reaction that occurs in this experiment is:

$$CuO + H_2 \rightarrow Cu + H_2O$$

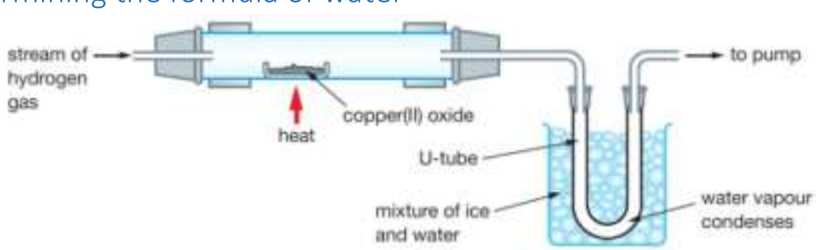
copper(II) oxide + hydrogen → copper + water

As well as a displacement reaction, this type of reaction can also be called a redox reaction. Redox reactions are discussed in Chapter 14.

KEY POINT

This is a displacement (or competition) reaction. Hydrogen is more reactive than copper and displaces it from copper oxide. Displacement reactions are discussed in Chapter 14.

Determining the formula of water



	H	O
masses/g	0.08	0.64
number of moles of atoms	0.08/1	0.64/16
number of moles of atoms	0.08	0.040
divide by the smaller number to find the ratio	0.08/0.040	0.040/0.040
ratio of moles	2	1
empirical formula	H_2O	

Working out formulae using percentage composition figure

EXAMPLE 5

Find the empirical formula of a compound containing 82.7% C and 17.3% H by mass (A_r of H = 1, A_r of C = 12).

The percentage figures apply to any amount of substance you choose, so let's choose 100 g. In this case the percentages convert simply into masses: 82.7% of 100 g is 82.7 g (Table 5.1).

Table 5.1 Finding the ratio from percentage by mass

	C	H
percentages/%	82.7	17.3
masses in 100 g/g	82.7	17.3
number of moles of atoms	82.7/12	17.3/1
number of moles	6.89	17.3
divide by smallest to get ratio	6.89/6.89	17.3/6.89
ratio of moles	1	2.5

From this we can see that there are 2.5 mol of H atoms for every mole of C atoms. The empirical formula, however, is the *whole number ratio* of the elements present in a compound. To obtain a whole number here we multiply both numbers by 2 to get C_2H_5, which is the empirical formula.

Converting empirical formulae into molecular formulae

EXAMPLE 6

A compound has the empirical formula CH_2. If the relative formula mass is 56, work out the molecular formula.

The relative formula mass of CH_2 is 12 + (2 × 1) = 14.

56/14 = 4

Therefore there must be 4 lots of CH_2 in the actual molecule and the molecular formula is C_4H_8.

Remember, the ratio of elements in the molecular formula must be the same as in the empirical formula: C_4H_8 cancels down to CH_2.

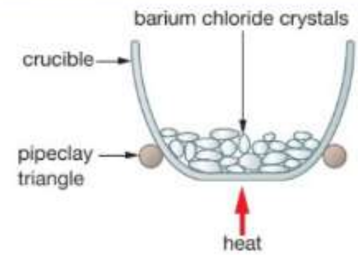
Empirical formula calculations involving water of crystallization

Finding the *n* in $BaCl_2.nH_2O$

	$BaCl_2$	H_2O
masses/g	2.08	0.36
divide by M_r to find the number of moles	2.08/208	0.36/18
number of moles	0.0100	0.020
ratio of moles	1	2
empirical formula	$BaCl_2{\cdot}2H_2O$	

REMINDER

Anhydrous means without water.



Calculations using moles, chemical equations and masses of substances

Interpreting symbols in equations in terms of moles

The chemical equation : $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

means : 1 mol CH_4 reacts with 2 mol O_2 to form 1 mol CO_2 and 2 mol H_2O

Problem involving heating calcium carbonate

When $CaCO_3$, is heated CaO is released. Calculating the mass of calcium oxide produces by heating 25g of calcium carbonate.

Calculation

The method we will use has three steps:

Step 1: Calculate the number of moles using the mass that you have been given.

Step 2: Use the chemical equation to work out the number of moles of the substance you are interested in.

Step 3: Convert the number of moles to a mass.

In this reaction we have 25 g of $CaCO_3$ and the number of moles can be calculated as:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}} = \frac{25}{100} = 0.25\,\text{mol}$$

The chemical equation for this reaction is

$$CaCO_3 \rightarrow CaO + CO_2$$

We can see from the equation that 1 mole of $CaCO_3$ will decompose to produce 1 mole of CaO. In other words, the equation is telling us that if we start with a certain number of moles of $CaCO_3$ we will obtain the same number of moles of CaO at the end.

Therefore, we can deduce that 0.25 mol $CaCO_3$ will decompose to produce 0.25 mol CaO. Now that we know the number of moles of CaO ($M_r = 56$) we can convert this to a mass:

$$\text{mass} = \text{number of moles} \times \text{mass of 1 mole}$$

$$\text{mass} = 0.25 \times 56 = 14\,\text{g}$$

Therefore, the reaction will produce 14 g of calcium oxide.

Alternative method

The calculation can also be done in a different way using ratios:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Interpret the equation in terms of moles:

1 mol $CaCO_3$ produces 1 mol CaO (and 1 mol CO_2)

Substitute masses where relevant:

100 g (1 mol) $CaCO_3$ produces 56 g (1 mol) CaO

Do a proportion calculation:

If 100 g of calcium carbonate gives 56 g of calcium oxide

1 g of calcium carbonate gives $\frac{56}{100}$ g of calcium oxide = 0.56 g

25 g of calcium carbonate gives 25 × 0.56 g of calcium oxide = 14 g of calcium oxide

A problem about extracting iron

The equation below is for a reaction that occurs in the extraction of iron:

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

Calculate the mass of iron which can be formed from 1000 g of iron oxide.

We are given the mass of Fe_2O_3 in the question and so we can calculate the number of moles of Fe_2O_3:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}$$

$$\text{number of moles} = \frac{1000}{160} = 6.25\,\text{mol}$$

From the chemical equation we can see that 1 mol Fe_2O_3 produces 2 mol Fe. We know this because of the big numbers in front of the formulae.

So we know that, if we start with a certain number of moles of Fe_2O_3, we will obtain twice as many moles of Fe.

So 6.25 mol Fe_2O_3 produces $2 \times 6.25 = 12.5$ mol Fe:

$$\text{mass} = \text{number of moles} \times \text{mass of 1 mole}$$

The mass of 12.5 mol Fe is $12.5 \times 56 = 700$ g.

Alternative method in terms of ratios

First interpret the equation in terms of moles:

1 mol Fe_2O_3 reacts with C to form 2 mol Fe (and CO_2).

We are only looking at how much iron is produced, so let's introduce just these masses.

160 g (1 mol) of Fe_2O_3 produces 2×56 g (2 mol) of Fe.

That is, 160 g of Fe_2O_3 produces 112 g of Fe.

From this we can calculate that 1 g of Fe_2O_3 will produce 112/160 g of Fe, and therefore 1000 g of Fe will produce $1000 \times 112/160 = 700$ g.

Problems involving the extraction of lead

Lead is extracted from galena, PbS. The ore is roasted in air to produce lead(II) oxide, PbO:

$$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$$

The lead(II) oxide is then converted to lead by heating it with carbon in a blast furnace:

$$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$$

The molten lead is tapped from the bottom of the furnace.

Calculate the mass of lead that would be produced from 1 tonne of galena.

(A_r: O = 16, S = 32, Pb = 207)

Number of moles = $\frac{\text{mass}}{\text{mass of 1 mole}}$

The number of moles of PbS = $\frac{1\,000\,000}{239}$ = 4184 mol.

From the first equation:

$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$

we can see that 2 mol PbS produce 2 mol PbO, therefore we can work out that 4184 mol PbS will produce 4184 mol PbO.

From the second equation:

$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$

we can see that 1 mol PbO produces 1 mol Pb. But we are starting this reaction with 4184 mol PbO and so we can work out that we will produce 4184 mol Pb.

The mass of 4184 mol Pb is 4184 × 207 = 866 000 g.

We can convert this to tonnes by dividing by 1 000 000:

866 000/1 000 000 = 0.866 tonne

Alternative method

$2PbS(s) + 3O_2(g) \rightarrow 2PbO(s) + 2SO_2(g)$

$PbO(s) + C(s) \rightarrow Pb(l) + CO(g)$

Interpret the equation in terms of moles and trace the lead through the equations:

2 mol PbS produces 2 mol PbO

If we double the second equation:

$2PbO(s) + 2C(s) \rightarrow 2Pb(l) + 2CO(g)$

we can see that 2 mol PbO produces 2 mol Pb.

In other words, every 2 mol of PbS produces 2 mol of Pb.

Substitute masses where relevant. In this case, the relevant masses are only the PbS and the Pb:

2 × 239 g PbS produces 2 × 207 g Pb

478 g PbS produces 414 g Pb

Now there seems to be a problem. The question is asking about tonnes and not grams. You could calculate how many grams there are in a tonne. However, it's much easier to think a bit, and realise that the ratio is always going to be the same, whatever the units, which means that:

478 tonnes PbS produces 414 tonnes Pb

Do the proportion calculation:

If 478 tonnes PbS produces 414 tonnes Pb

then 1 tonne PbS gives $\frac{417}{478}$ tonne Pb = 0.866 tonne

0.866 tonne of lead is produced from 1 tonne of galena.

KEY POINT

We've doubled the second equation so that we can trace what happens to all the 2PbO from the first one.
This could also be simplified to 1 mol PbS produces 1 mol Pb. This would save you doing some arithmetic as you would not have to multiply everything by 2. However, in the end it does not make any difference to the answer, so you do not have to simplify it if you don't want to.

calculating percentage yields

The percentage yield shows how much product is obtained compared to the maximum possible mass.

Calculating the percentage yield

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

EXAMPLE 7

A student reacted 2.40 g of copper(II) oxide (CuO) with hot sulfuric acid. She made 5.21 g of copper(II) sulfate crystals ($CuSO_4{\cdot}5H_2O$).

The equations for the reactions are:

$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$

$CuSO_4(aq) + 5H_2O(l) \rightarrow CuSO_4{\cdot}5H_2O(l)$

First we need to calculate the theoretical yield. This is done by carrying out a moles calculation. We have enough information to calculate the number of moles of copper(II) oxide:

$$\text{number of moles of CuO} = \frac{2.40}{79.5} = 0.0302\,\text{mol}$$

From the first equation we can deduce that 0.0302 mol CuO will produce 0.0302 mol $CuSO_4$.

From the second equation we can deduce that 0.0302 mol $CuSO_4$ will produce 0.0302 mol $CuSO_4{\cdot}5H_2O$.

Now we need to calculate the mass of $CuSO_4{\cdot}5H_2O$ by multiplying the number of moles by the mass of 1 mole:

$$\text{mass} = 0.0302 \times 249.5 = 7.53\,\text{g}$$

This is the theoretical yield of copper(II) sulfate crystals.

The actual yield is what was obtained in the experiment. This was 5.21 g.

The percentage yield = (actual yield/theoretical yield) × 100

The percentage yield is 5.21/7.53 × 100 = 69.2%.

Calculations in which you have to calculate which substance is in excess

EXAMPLE 8

Copper reacts with concentrated nitric acid according to the equation:

$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l) + 2NO_2(g)$

3.2 g of copper is reacted with 0.40 mol concentrated nitric acid. Work out which reagent is in excess.

We need to convert the mass of Cu to a number of moles:

$$\text{number of moles of Cu} = \frac{3.2}{63.5} = 0.050\,\text{mol}$$

Looking at the chemical equation we can see that 1 mol Cu reacts with 4 mol HNO_3, so we can calculate that:

0.050 mol Cu reacts with 4 × 0.050 = 0.20 mol HNO_3

We have more than 0.20 mol HNO_3 therefore HNO_3 is in excess.

If the question went further and asked us to calculate the amount in moles of NO_2 produced using these quantities we would have to use the number of moles of Cu because not all the HNO_3 is used up. We would obtain 2 × 0.050 = 0.10 mol NO_2.

KEY POINT

You always have to use the chemical equation to determine which substance is in excess. If we had been given 0.2 mol Mg and 0.3 mol HCl in this reaction it would look like the HCl is in excess because there is more of it. But if you look at the chemical equation you can see that 0.3 mol HCl only reacts with 0.15 mol Mg so there is still not enough HCl to react with all the Mg, so the Mg is still in excess.

KEY POINT

A reagent is just something that reacts.

FOCUS

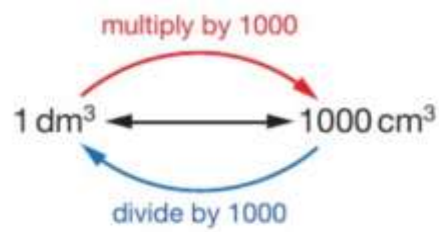
1.6 chemical formulae, equations and calculations: part 2

Calculations involving gas volume

Avogadro's law:
Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

Units of volume

- Cubic centimeters(cm^3)
- Cubic decimeters(dm^3)
- Liters(l)



$$1 \text{ liter} = 1 \text{ dm}^3 = 1000 \text{cm}^3$$

The volume occupied by one mole of gas

1 mole of any gas contains the same number of molecules and so occupies the same volume as 1 mole of any other gas at the same temperature and pressure. At room temperature and pressure, the volume occupied by 1 mole of any gas is approximately 24 dm^3 (24 000 cm^3). The volume occupied by 1 mole of a gas is often called the molar volume.

KEY POINT

1 mole of any gas occupies 24 dm^3 (24 000 cm^3) at rtp.
The abbreviation **rtp** is commonly used for 'room temperature and pressure'.
rtp is usually taken as 20–25 °C and 1 atmosphere pressure.

Calculations with molar volume

Calculating the volume of a certain number of moles

Calculate the volume in dm^3 of 0.20 mol CO_2 at rtp:

volume = number of moles × molar volume

Because we want the volume in dm^3 we use 24 dm^3 as the molar volume:

volume = 0.20 × 24 = 4.8 dm^3

Calculating the volume of a given mass of a gas

Calculate the volume (in cm^3) of 0.01 g of hydrogen at rtp (A_r: H = 1).

1 mol H_2 has a mass of 2 g:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mol}}$$

0.01 g of hydrogen is $\frac{0.01}{2}$ mol = 0.005 mol.

Because we want the volume in cm^3 we use the molar volume as 24 000 cm^3:

volume = number of moles × molar volume

0.005 mol of hydrogen occupies 0.005 × 24 000 = 120 cm^3

Calculating the number of moles from a volume

Calculate the amount of moles in 120 cm³ of carbon dioxide.

We can see that the volume is given in cm³ therefore we use 24 000 cm³ as the molar volume:

$$\text{number of moles} = \frac{\text{volume of gas}}{\text{molar volume}}$$

$$\text{number of moles} = \frac{120}{24\,000}$$

$$\text{number of moles} = 0.005\,\text{mol}$$

Using the molar volume in calculations with chemical equations

Calculate the volume of carbon dioxide produced at room temperature and pressure when an excess of dilute hydrochloric acid is added to 1.00 g of calcium carbonate. (A_r: C = 12, O = 16, Ca = 40; molar volume = 24 dm³ at rtp.)

The equation for the reaction is

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

We have been given the mass of calcium carbonate ($CaCO_3$) and so we can calculate the number of moles of calcium carbonate:

$$\text{number of moles} = \frac{\text{mass}}{\text{mass of 1 mole}}$$

The mass of 1 mole of $CaCO_3$ is the M_r in grams. The M_r is 40 + 12 + (3 × 16) = 100.

The mass of 1 mole of $CaCO_3$ is 100 g.

The number of moles in 1.00 g of $CaCO_3 = \frac{1.00}{100} = 0.0100\,\text{mol}$.

From the chemical equation we can see that 1 mol of $CaCO_3$ produces 1 mol of CO_2 so we know that 0.0100 mol of $CaCO_3$ will produce the same number of moles of CO_2, that is, 0.0100 mol.

We now need to work out the volume occupied by 0.0100 mol of CO_2:

$$\begin{aligned}\text{volume of gas} &= \text{number of moles} \times \text{molar volume}\\ &= 0.0100 \times 24 = 0.24\,\text{dm}^3\end{aligned}$$

Problem involving making Hydrogen

Aluminium reacts with dilute hydrochloric acid according to the equation

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

What mass of aluminium would you need to add to an excess of dilute hydrochloric acid so that you produced 100 cm^3 of hydrogen at room temperature and pressure? (A_r of Al = 27; molar volume = 24 000 cm^3 at rtp.)

We can only calculate the number of moles of the hydrogen gas. We do not know the mass of aluminium, that is what we are trying to find.

The volume of hydrogen gas is given in cm^3 so we use the molar volume of a gas as 24 000 cm^3:

$$\text{number of moles of gas} = \frac{\text{volume of gas}}{\text{molar volume}} = \frac{100}{24\,000} = 0.00417\,\text{mol}$$

We could also write this number in standard form as 4.17×10^{-3} mol.

Looking at the chemical equation we can see that 3 mol of H_2 is produced from 2 mol of Al (we are only looking at the big numbers in front of the formulae). So we can see that the number of moles of Al will be $\frac{2}{3}$ times the number of moles of H_2.

The number of moles of aluminium required to produce 0.00417 mol of H_2 is $0.00417 \times \frac{2}{3} = 0.00278$ mol.

We can convert this to a mass by using the equation:

$$\text{mass} = \text{number of moles} \times \text{mass of 1 mol}$$
$$= 0.00278 \times 27 = 0.075\,\text{g Al}$$

Working with solution concentration

Concentrations of solutions

- g/dm^3 or $g\ dm^{-3}$
- mol/dm^3 or $mol\ dm^{-3}$

EXAMPLE 1

What is the concentration of a 0.050 mol/dm^3 solution of sodium carbonate, Na_2CO_3, in g/dm^3? (A_r: C = 12, O = 16, Na = 23)

1 dm^3 of solution contains 0.050 mol Na_2CO_3.

1 mol Na_2CO_3 weighs 106 g.

0.050 mol weighs $0.050 \times 106 = 5.3$ g.

1 dm^3 of solution contains 5.3 g Na_2CO_3, therefore the concentration is 5.3 g/dm^3.

KEY POINT

You may also find the symbol M used. For example, dilute hydrochloric acid might have a concentration quoted as 2 M. M means 'mol/dm^3' and is described as the **molarity** of the solution. You can also read 2 M as '2 molar'.

EXAMPLE 2

What is the concentration in mol/dm^3 of a solution containing 2.1 g of sodium hydrogencarbonate, $NaHCO_3$, in 250 cm^3 of solution? (A_r: H = 1, C = 12, O = 16, Na = 23)

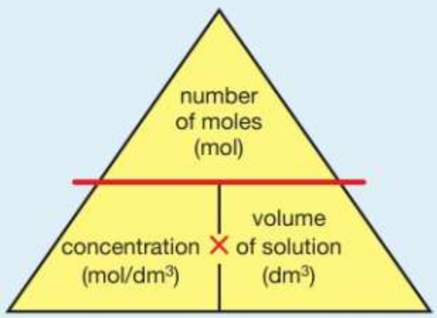
1 mol $NaHCO_3$ has a mass of 84 g.

$$2.1\text{ g is } \frac{2.1}{84} = 0.025\text{ mol}$$

This is in a volume of 250 cm^3 but we need to find out how much there is in 1 dm^3 (1000 cm^3). There are four lots of 250 cm^3 in 1000 cm^3. Each portion of 250 cm^3 contains 0.025 mol, therefore there must be $4 \times 0.025 = 0.10$ mol in 1000 cm^3.

The concentration is 0.10 mol/dm^3.

Another way of doing this is to use the triangle shown in Figure 6.8.



▲ Figure 6.8 We can use the triangle for calculations involving solutions.

Working out a number of moles from a volume and a concentration

$$\boxed{\text{number of moles} = \text{volume of solution} \times \text{concentration}}$$

EXAMPLE 3

Calculate the number of moles of NaOH in 50 cm^3 of 0.10 mol/dm^3 solution.

Converting the volume to dm^3 we get 50/1000 = 0.050 dm^3.

Number of moles = volume of solution (dm^3) × concentration (mol/dm^3).

Number of moles = 0.050 × 0.10 = 0.0050 mol.

Calculations with equations involving solutions

Limescale can be removed from, for example, electric kettles by reacting it with a dilute acid such as ethanoic acid, which is present in vinegar:

$$CaCO_3(s) + 2CH_3COOH(aq) \rightarrow (CH_3COO)_2Ca(aq) + CO_2(g) + H_2O(l)$$

What mass of calcium carbonate can be removed by 50 cm^3 of a solution of ethanoic acid that has a concentration of 2 mol/dm^3? (A_r: C = 12, O = 16, Ca = 40)

We know the volume and concentration of ethanoic acid and so we can calculate the number of moles of ethanoic acid:

number of moles = volume of solution (dm^3) × concentration (mol/dm^3)

We need to be careful to convert the volume of ethanoic acid to dm^3 by dividing by 1000:

$$\text{number of moles of ethanoic acid} = \frac{50}{1000} \times 2 = 0.1\,\text{mol}$$

From the chemical equation we can see that there is a 2 in front of the CH_3COOH (ethanoic acid) but no number (which means a 1) in front of the $CaCO_3$, so we can deduce that 2 mol of CH_3COOH react with 1 mol of $CaCO_3$, in other words the number of moles of $CaCO_3$ is half the number of moles of CH_3COOH. Therefore, we can deduce that 0.1 mol of ethanoic acid will react with 0.1/2 = 0.05 mol of $CaCO_3$.

We can now calculate the mass of $CaCO_3$ (M_r = 100):

mass = number of moles × mass of 1 mol = 0.05 × 100 = 5 g

Therefore the ethanoic acid reacts with 5 g of calcium carbonate.

Calculations from titrations

Acid-alkali titration

Titration is a technique that is used to find out how much of one solution reacts with a certain volume of another solution of known concentration. A solution of an *alkali* is measured into a conical flask using a pipette. An *acid* is run in from a burette, swirling the flask constantly. Towards the end, the acid is run in a drop at a time until the *indicator* just changes color.

The standard calculation

25.0 cm^3 of sodium hydroxide solution of unknown concentration was titrated with dilute sulfuric acid of concentration 0.050 mol/dm^3. 20.0 cm^3 of the acid was required to neutralise the alkali. Find the concentration of the sodium hydroxide solution in mol/dm^3.

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

This time, we know everything we need about the sulfuric acid and can calculate the number of moles.

The experiment used 20.0 cm^3 of 0.050 mol/dm^3 H_2SO_4:

$$\text{number of moles of sulfuric acid used} = \frac{20.0}{1000} \times 0.050$$
$$= 0.0010\,\text{mol}$$

The equation proportions aren't 1:1 this time. That's what makes the calculation slightly different from the last one. The equation shows that 1 mol of sulfuric acid reacts with 2 mol of sodium hydroxide. So the number of moles of sodium hydroxide is twice the number of moles of sulfuric acid:

number of moles of sodium hydroxide = 2 × 0.0010 = 0.0020 mol

That 0.0020 mol must have been in the 25.0 cm^3 (25/1000 = 0.025 dm^3) of sodium hydroxide solution:

$$\text{concentration (mol/dm}^3) = \frac{\text{number of moles (mol)}}{\text{volume (dm}^3)}$$

concentration = 0.0020/0.025 = 0.080 mol/dm^3

Therefore the concentration of the sodium hydroxide is 0.080 mol/dm^3.

Reverse the calculation

EXAMPLE 4

Calculate the volume of 0.100 mol/dm^3 sodium hydrogencarbonate ($NaHCO_3$) solution needed to neutralise 20.0 cm^3 of 0.125 mol/dm^3 hydrochloric acid (HCl).

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

We have been given the volume and concentration of the hydrochloric acid and so we can calculate the number of moles of this. We do not have enough information to calcuate the number of moles of anything else, so must start here:

$$\text{number of moles of HCl} = \frac{20.0}{1000} \times 0.125 = 0.00250\,\text{mol}$$

The equation shows that you will need the same number of moles of sodium hydrogencarbonate. Therefore we know that we need 0.00250 mol of $NaHCO_3$.

To calculate the volume this is in, we just need to rearrange our concentration equation to obtain:

$$\text{volume (dm}^3\text{)} = \frac{\text{number of moles (mol)}}{\text{concentration (mol/dm}^3\text{)}}$$

$$\text{volume} = \frac{0.00250}{0.100} = 0.0250\text{ dm}^3$$

The volume comes out in dm^3 because the concentration is in mol/dm^3. We can convert to cm^3 by multiplying by 1000:

$$\text{volume of } NaHCO_3 \text{ solution} = 25.0\text{ cm}^3$$

You will need 25.0 cm^3 of the sodium hydrogencarbonate solution to neutralise the hydrochloric acid.

EXCERSISE

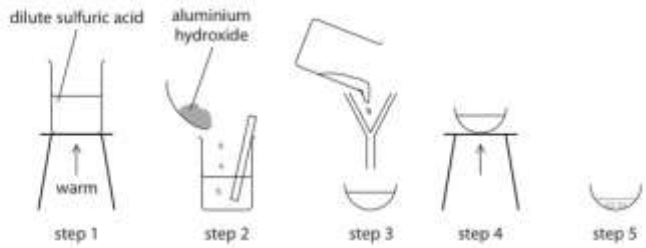
1.
   a. Many different salts can be prepared from acids. The table shows the reactants used in two salt preparations. Complete the table to show the name of the salt formed and the other product(s) in each case.

Reactants	Name of salt formed	Other product(s)
zinc + hydrochloric acid		
calcium carbonate + nitric acid		

b. A student uses the reaction between aluminium hydroxide and dilute sulfuric acid to prepar a pure, dry sample of aluminium sulfate crystals. The equation for the reaction used to prepare this salt is:

$2Al(OH)_3 + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6H_2O$

The diagram shows the steps in the student's method.



i. State two ways to make sure that all the acid is reacted in step 2.

ii. State the purpose of filtration in step 3.

iii. In step 5, the basin is left to cool to room temperature to allow crystals of aluminum sulfate to form. Sate one method of dying these crystals.

c. The student records this information about the reagents she uses in her preparation. Mass of aluminium hydroxide = 3.9 g amount of sulfuric acid = 0.090 mol. Determine which reagent is in excess, making use of this information and the equation in part (b).

d. Another student prepares 0.025 mol of aluminium sulfate. The formula of aluminium sulfate is $Al_2(SO_4)_3$. Calculate the mass of aluminium sulfate prepared.

e. The equation for another reaction used to prepare a sample of a salt is:

$PbO+2HNO_3 \rightarrow Pb(NO_3)_2 + H_2O$

In one experiment, the amount of lead(II) oxide used was 0.75 mol and the amount of nitric acid used was 1.5 mol. At the end of the experiment, the mass of lead(II) nitrate obtained was 209 g. Calculate the percentage yield of lead(II) nitrate in this experiment.

[M, of lead(II) nitrate = 331]

2.

a. This question is about three stages in the manufacture of sulfuric acid. In stage 1, sulfur is burned in oxygen to form sulfur dioxide gas.

$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$

i. State one environmental problem caused by the release of sulfur dioxide into the atmosphere. A mass of 6.4 tonnes of sulfur is burned to produce sulfur dioxide gas. Calculate the maximum volume, in dm3, of sulfur dioxide gas that can be produced at rtp.

[molar volume of sulfur dioxide gas at rtp = 24 dm3]

[1 tonne = 10° gl

Give your answer in standard form.

b. In stage 2, sulfur dioxide is reacted with oxygen to form sulfur trioxide gas.

$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$

The yield of sulfur trioxide is approximately 98%.

i. A catalyst is used in this reaction. Explain how a catalyst increases the rate of a reaction.

ii. The temperature is kept constant. Give a reason why increasing the pressure would increase the yield of sulfur trioxide.

iii. Suggest why it is not necessary to increase the pressure in stage 2.

c. In stage 3, the sulfur trioxide is reacted with concentrated sulfuric acid to form a liquid called oleum, $H_2S_2O_7$. The oleum is then added to water to form concentrated sulfuric acid. Complete the chemical equations for these two reactions.

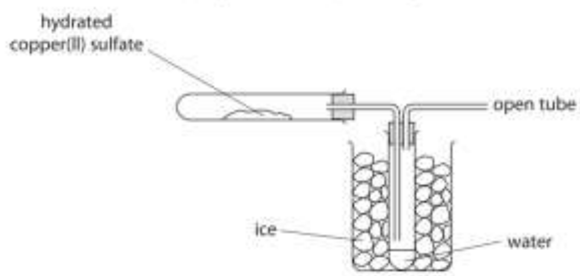
……………………..… + ……………………….. → $H_2S_2O_7$

$H_2S_2O_7$ + …………………………….. → ………………………….

d. Sulfuric acid reacts with ammonia to form ammonium sulfate, $(NH_4)_2SO_4$. Calculate the percentage by mass of nitrogen in ammonium sulfate.
[M, of $(NH_4)_2SO_4$ = 132]

3. The apparatus in the diagram is used to heat a sample of hydrated copper(II) sulfate crystals, $CuSO_4.5H_2O$. The equation for the reaction that takes place is

$CuSO_4.5H_2O_{(s)} \rightarrow CuSO_{4(s)} + 5H_2O_{(l)}$



a. Draw an arrow on the diagram to show where heat is applied.

b. What is the purpose of the ice?

c. Calculate the maximum mass of water that could be collected when a sample of hydrated copper(II) sulfate of mass 2.50 g is heated. [M_r of $CuSO_4.5H_2O$ is 250]

4. Potassium hydrogen carbonate ($KHCO_3$) decomposes on heating. Three possible equations for the decomposition are:

equation 1: $2KHCO_{3(s)} \rightarrow K_2O_{(s)} + 2CO_{2(g)} + H_2O_{(g)}$

equation 2: $KHCO_{3(s)} \rightarrow KOH_{(s)} + CO_{2(g)}$

equation 3: $2KHCO_{3(s)} \rightarrow K_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$

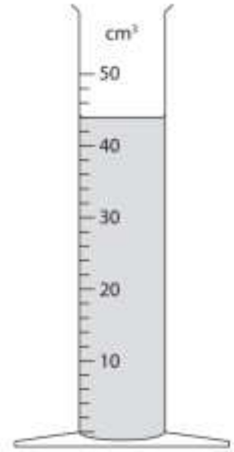
When 8.00 g of potassium hydrogen carbonate is heated until it is fully decomposed, 5.52 g of solid is formed.

a. Complete the table by calculating the amount, in moles, of each solid.

Solid	M_r of solid	Mass of solid in g	Amount of solid in mol
$KHCO_3$	100	8.00	
K_2O	94	5.52	
KOH	56	5.52	
K_2CO_3	138	5.52	

b. Use the information in the table to explain which equation, 1, 2 or 3, represents the decomposition of potassium hydrogen carbonate.

5. The piece of apparatus shown contains 0.010 mol/dm^3 hydrochloric acid.

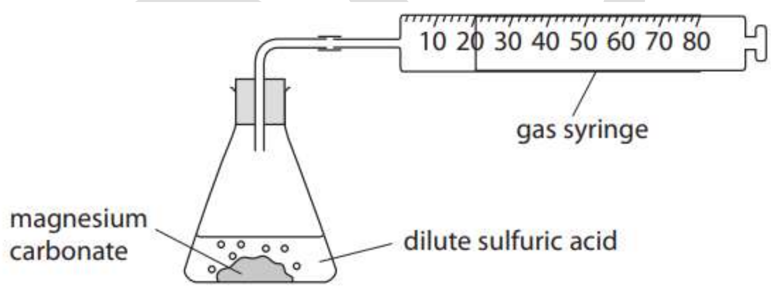


a.
   i. Give the name of this piece of apparatus.
   ii. What volume of hydrochloric acid is in the apparatus?
   iii. Use your answer in (a)(ii) to calculate the amount, in moles, of hydrochloric acid in the apparatus.

b. A student poured a solution containing 0.010 mol of hydrochloric acid into a beaker. He then added 0.0075 mol of zinc powder and collected the hydrogen given off in a gas syringe.
The equation for the reaction is
$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
Is the zinc or the hydrochloric acid in excess? Explain your answer.

c. The student repeated the experiment with 0.0075 mol of magnesium powder with the same total surface area as the zinc. The equation for the reaction is
$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
   i. What effect would this change have on the rate at which the hydrogen is given off?
   ii. What effect would this change have on the volume of hydrogen produced?

6. A student carries out an investigation into the reaction between magnesium carbonate and dilute sulfuric acid. He uses this apparatus.



The student carries out seven experiments. In each experiment he uses the same mass of magnesium carbonate but a different volume of acid. He measures the total volume of carbon dioxide collected in each experiment. The table shows his results.

Volume of sulfuric acid used in cm^3	0	5	15	20	25	30	35	40
Volume of carbon dioxide collected in cm^3	0	16	47	61	64	78	80	80

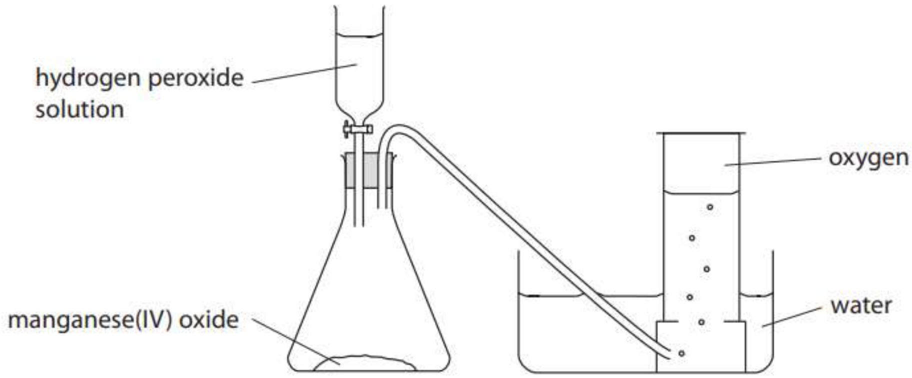
a. Plot the results on the grid and draw a curve of best fit.

b.
   i. Which volume of sulfuric acid produces an anomalous result?
   ii. Explain what the results with 35 cm3 and 40 cm3 of sulfuric acid indicate about the reaction.

iii. Use the graph to find the volume of carbon dioxide that would be collected if 10 cm^3 of acid were used.

iv. Use the graph to find the volume of sulfuric acid that would result in 55 cm^3 of carbon dioxide being collected.

7. In 1774, the scientist Joseph Priestley produced oxygen by heating mercury(II) oxide, (HgO). When heated, mercury(II) oxide breaks down into its elements.
   a.
      i. Write a chemical equation for the breakdown of mercury(II) oxide into its elements.
      ii. What name is given to this type of reaction?
   b. Priestley's method of producing oxygen is no longer used because of the high toxicity of mercury and mercury compounds. A student prepares oxygen by adding hydrogen peroxide solution to solid manganese(IV) oxide. The diagram shows the apparatus used.



   The equation for the reaction is: $2H_2O_{2(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)}$
      i. Give the name of the apparatus that contains the hydrogen peroxide solution.
      ii. Suggest how the first sample of gas collected may be different from the samples collected later.
   c. A catalyst increases the rate of decomposition of the hydrogen peroxide. Describe a method you could use to show that the manganese(IV) oxide is acting as a catalyst in this reaction.
   d. Sulfur burns in oxygen to produce sulfur dioxide (SO_2). Sulfur dioxide is very soluble in water.
      i. Write a chemical equation for the reaction that takes place when sulfur dioxide dissolves in water.
      ii. Universal indicator is added to the solution formed in (d)(i). Explain the effect that the solution has on the universal indicator

1.7 ionic bonding

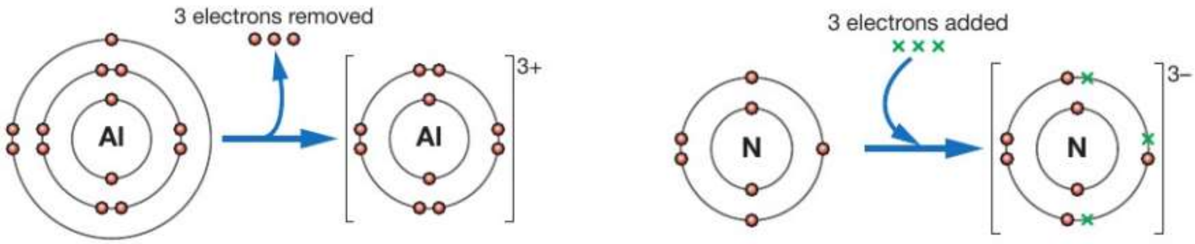
Ionic boning

Magnesium oxide, calcium fluoride, zinc bromide are a few examples. All these compounds contain a metal combined with a non-metal.

When an ionic compound is formed, electron(s) are transferred from a metal atom to a non-metal atom to form positive and negative ions. Ionic bonding is the strong electrostatic attraction between positive and negative ions. Ionic bonding is often shown using dot-and-cross diagrams.

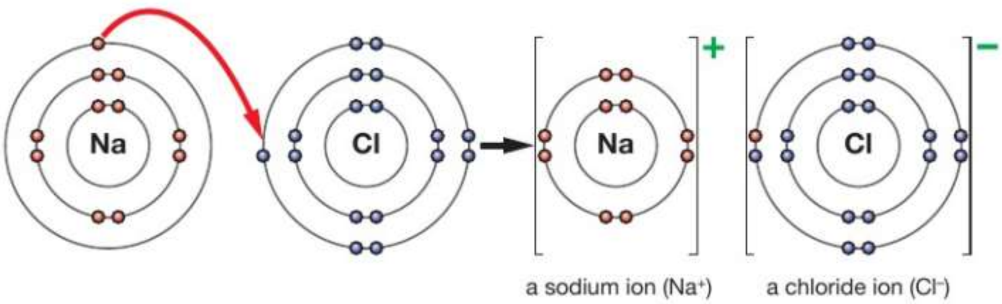
Charged particles are called ions. Whereas the positive ions are called cations and negative ions are called anions.

Significance of noble gas electronic configuration in ionic bonding

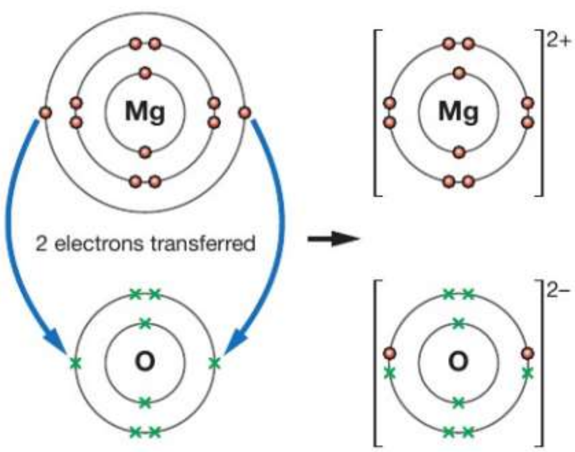


Examples for ionic bonds

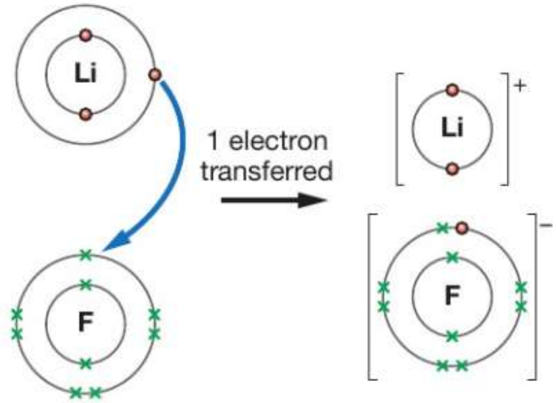
Sodium chloride



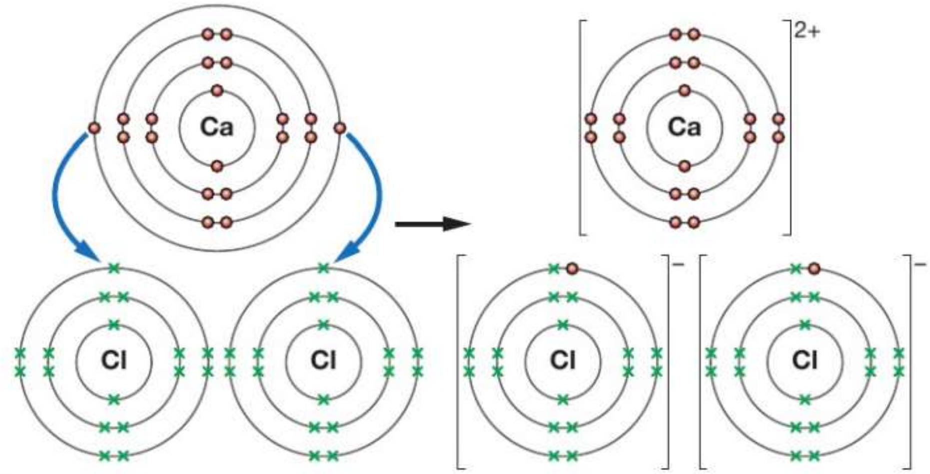
Magnesium oxide



Lithium flouride



Calcium chlirode



Formulae for ionic bonding

Shortcut methods:

1. The need for equal numbers or pluses and minuses
2. Workout the charge on an ion

Group in Periodic Table	Charge on ion	Example
1	1+	Na^+
2	2+	Mg^{2+}
3	3+	Al^{3+}
5	3–	N^{3-}
6	2–	O^{2-}
7	1–	Br^-

KEY POINT

Ammonium chloride (NH_4Cl) is an example of an ionic compound that does not contain a metal. There is ionic bonding between the NH_4^+ and Cl^- ions. There is, however, also covalent bonding (see Chapter 8) in this compound: the NH_4^+ ion is held together by covalent bonding.

3. Name tells you the charge

All metals form positive ions. But you cannot work out some of the ions(given in the table)

Charge	Substance	Ion
positive	zinc	Zn^{2+}
	silver	Ag^+
	hydrogen	H^+
	ammonium	NH_4^+

Charge	Substance	Ion
negative	nitrate	NO_3^-
	hydroxide	OH^-
	carbonate	CO_3^{2-}
	sulfate	SO_4^{2-}

Confusing endings

Any 'ide' endings mean that there isn't anything complicated. Whereas 'ate' ending, means that there is oxygen (and possibly any other things) there in it.

Deducing the formula for an ionic compound

EXAMPLE 1

To find the formula for sodium oxide

Sodium is in Group 1, so the ion is Na^+.

Oxygen is in Group 6, so the ion is O^{2-}.

To have equal numbers of positive and negative charges, you would need two sodium ions to provide the two positive charges to cancel the two negative charges on one oxide ion. In other words, you need:

$Na^+ \quad Na^+ \quad O^{2-}$

The formula is therefore $\mathbf{Na_2O}$.

EXAMPLE 2

To find the formula for barium nitrate

Barium is in Group 2, so the ion is Ba^{2+}.

Nitrate ions are NO_3^-. You will have to remember this.

To have equal numbers of positive and negative charges, you would need two nitrate ions for each barium ion.

The formula is $\mathbf{Ba(NO_3)_2}$.

Notice the brackets around the nitrate group. *Brackets must be written if you have more than one of these complex ions* (ions containing more than one atom). In any other situation, they are completely unnecessary.

EXAMPLE 3

To find the formula for iron(III) sulfate

Iron(III) tells you that the metal ion is Fe^{3+}.

Sulfate ions are SO_4^{2-}.

To have equal numbers of positive and negative charges, you would need two iron(III) ions for every three sulfate ions, giving 6+ and 6– in total.

The formula is $\mathbf{Fe_2(SO_4)_3}$.

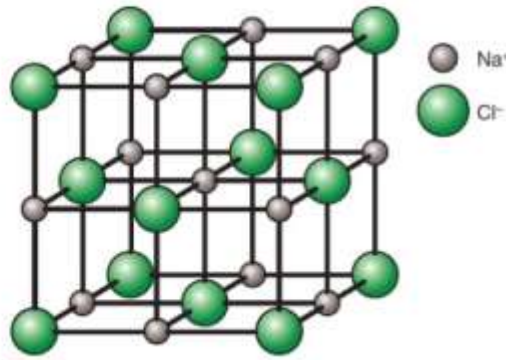
A shortcut to working out complicated formulae such as these is to just swap over the numbers in the charges. This is shown in Figure 7.13.

$$Fe^{3+} \; SO_4^{2-} \Rightarrow Fe_2(SO_4)_3$$

▲ Figure 7.13 If you cross over the numbers in the charges you will get the formula.

Giant ionic structures

A lattice is a regular array of particles. The lattice is held together by the strong electrostatic attractions between the positively and negatively charged ions.
The diagram shows an exploded view of sodium chloride. Magnesium oxide also has the same exact structure as sodium chloride but stronger.



Physical properties of ionic substances

- high melting points
- high boiling points
- tend to be crystalline
- brittle
- soluble in water
- insoluble in organic solvents
- does not conduct electricity in solid state. But a very good conductor when molten or dissolved in water

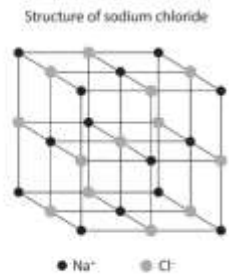
EXCERSISE

1. Hydrogen chloride is formed in the reaction between hydrogen and chlorine. The equation for the reaction is:

   $H_2 + Cl_2 \rightarrow 2HCl$

   a. Each molecule in this equation contains the same type of bonding. Name this type of bonding.
   b. The bonding in a hydrogen molecule is strong. Explain why the boiling point of hydrogen is low.
   c. Explain how the two atoms in a chlorine molecule are held together.
   d. Draw a dot and cross diagram to show the bonding in a hydrogen chloride molecule. Show only the outer electrons in each atom.
   e. Hydrogen chloride gas dissolves in water to form solution A. Hydrogen chloride gas dissolves in methylbenzene to form solution B. A teacher adds a piece of magnesium ribbon to each solution. Explain why she observes effervescence with solution A but not with solution B.

2. Sodium chloride (NaCl) and silicon dioxide (SiO_2) both have giant lattice structures. Sodium chloride is an ionic compound. Silicon dioxide is a covalent compound. The table shows some properties of each compound.



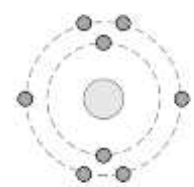
Sodium chloride	Silicon dioxide
melting point = 801 °C	melting point = 1610 °C
soluble in water	insoluble in water
conducts electricity when molten	does not conduct electricity when molten

a.
   i. Explain why silicon dioxide has a high melting point.
   ii. Suggest why the melting point of silicon dioxide is higher than the melting point of sodium chloride.

b. State why sodium chloride conducts electricity when molten.

c. Carbon dioxide is described as a simple molecular substance. State why carbon dioxide (CO_2) is a gas at room temperature.

3. The diagram shows how the electrons are arranged in an atom of oxygen. Oxygen atoms form both covalent and ionic bonds.



a. Water is formed when two atoms of hydrogen combine with one atom of oxygen.
   i. Draw a dot and cross diagram of a molecule of water. You need only show the electrons in the outer shells.
   ii. Explain how the covalent bonds in the water molecule hold the hydrogen and oxygen atoms together.

b. The electronic configuration of a sodium atom is 2.8.1 Sodium oxide, Na_2O, is an ionic compound formed when sodium reacts with oxygen.
   i. Describe, in terms of electrons, what happens when sodium oxide is formed in this reaction.
   ii. The reaction of sodium to form sodium oxide can be described as oxidation because it involves the addition of oxygen. State one other reason why this reaction can be described as oxidation.

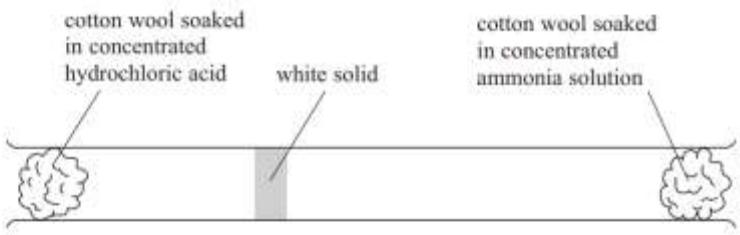
c. Explain why water has a much lower melting point than sodium oxide.

d. A teacher added sodium oxide to water in a beaker. The equation shows the reaction that occurred.

   $Na_2O(........) + H_2O(........) \rightarrow 2NaOH(........)$

   i. Insert the appropriate state symbols in this equation.
   ii. Some universal indicator was then added to the beaker. A colour change occurred. State the final colour of the universal indicator and identify the ion responsible for the colour change.

4. Ammonium chloride contains oppositely charged ions.
   a. State the formula of each ion.
   b.
      i. Describe a chemical test to show that a substance contains ammonium ions.
      ii. Describe a chemical test to show that a substance contains chloride ions.
   c. The reaction between ammonia and hydrogen chloride can be used to illustrate diffusion with the following apparatus.



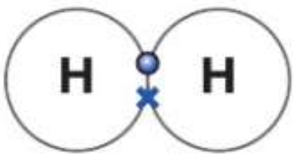
   After a few minutes, a white solid appears inside the tube.
      i. Identify the white solid
      ii. What does the diagram show about the speed of the ammonia molecules compared to the speed of the hydrogen chloride molecules?
   d. State the main hazard when using concentrated hydrochloric acid in the experiment in (d). Suggest one precaution you could use to minimize this hazard.

1.8 Covalent bonding

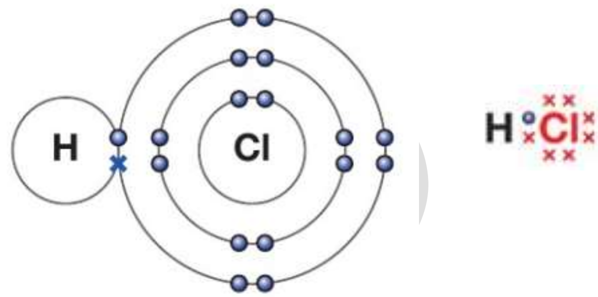
In a covalent bond, a pair of electrons is shared between two atoms. What holds the atoms together is the strong electrostatic attraction between the nuclei (positively charged) of the atoms that make up the bond, and the shared pair of electrons (negatively charged).

Covalent bonds are often shown using dot-and-cross diagrams. Molecules with just two atoms in known as *diatomic.*

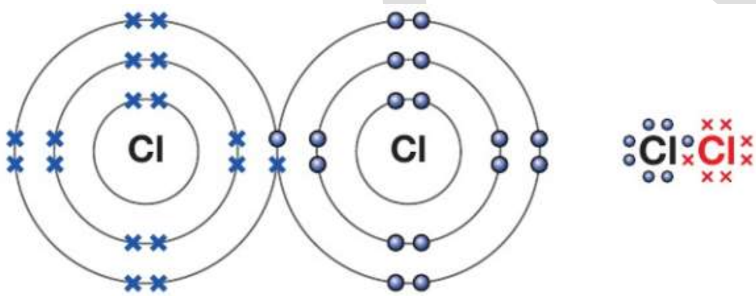
Hydrogen molecule (H_2)



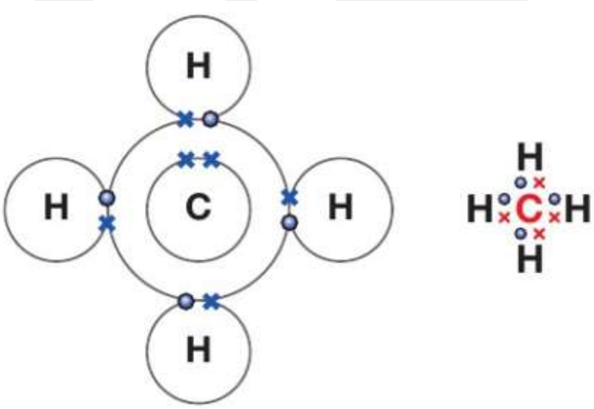
Hydrogen chloride molecule (HCl)



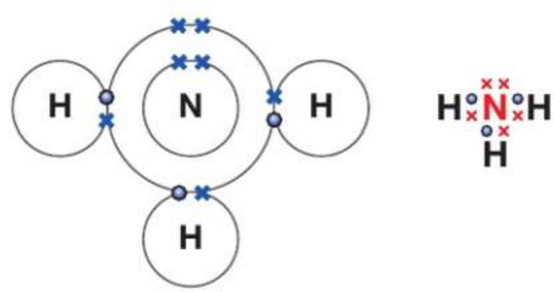
Chlorine molecule (Cl_2)



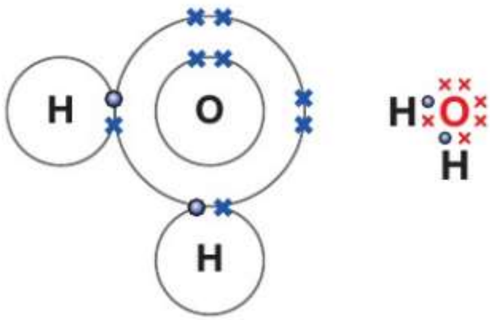
Methane molecule (CH_4)



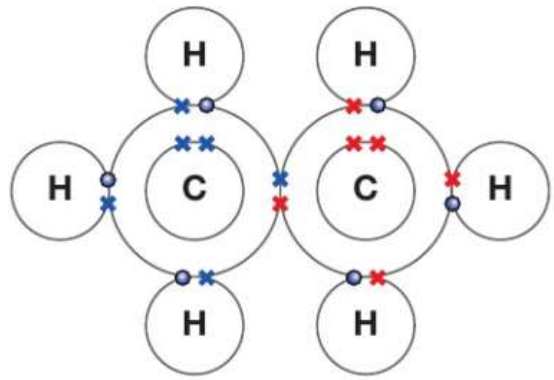
Ammonia molecule (NH_3)



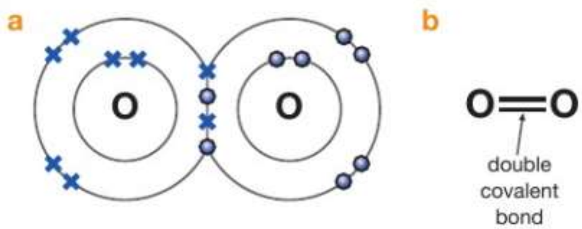
Water molecule (H_2O)



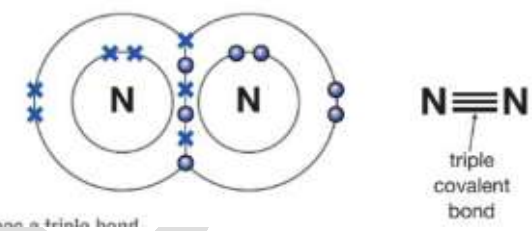
Ethane molecule (C_2H_6)



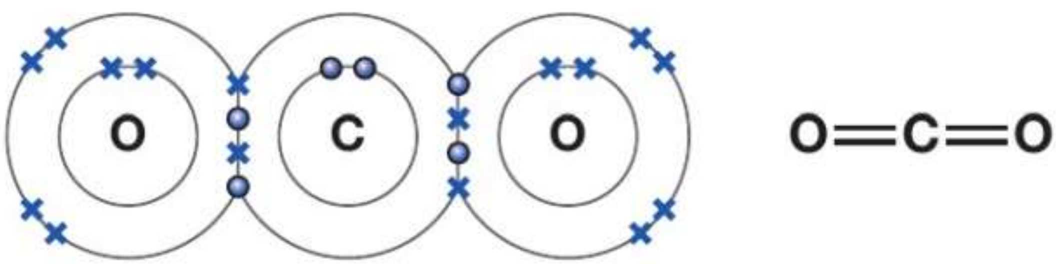
oxygen molecule (O_2)



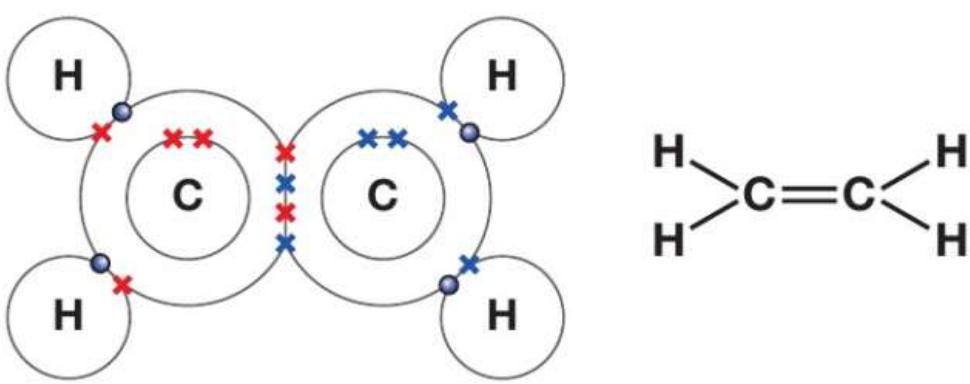
Nitrogen molecule(N_2)



Carbon dioxide (CO_2)

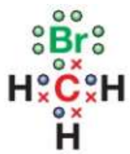


Ethane molecule (C_2H_4)

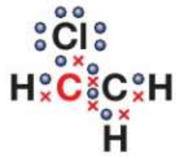


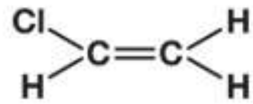
Organic molecules containing hydrogen atom

<u>CH_3Br</u>



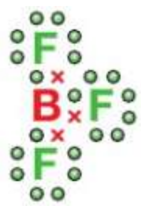
<u>Chloroethene</u>





molecules where the central atom does not have 8 electrons in its outer shell

<u>BF_3</u>



<u>SO_2</u>

Simple molecular structures

Intermolecular forces between water molecules which keep them in the liquid state. Intermolecular forces are much weaker than covalent bonds.

When a substance consists of molecules with intermolecular forces of attraction between them, we say that it has a simple molecular structure. Substances with simple molecular structures tend to be gases or liquids or solids with low melting points and boiling points. The reason for this is that not much energy is required to break the weak intermolecular forces of attractions between molecules. Examples of things that have simple molecular structures are H_2O, CO_2, CH_4, NH_3 and C_2H_4.

Melting and boiling points increase as relative molecular mass increases

Halogen	Formula	Relative molecular mass/M_r	Melting point/°C	Boiling point/°C
fluorine	F_2	38	–220	–188
chlorine	Cl_2	71	–101	–34
bromine	Br_2	160	–7	59
iodine	I_2	254	114	184

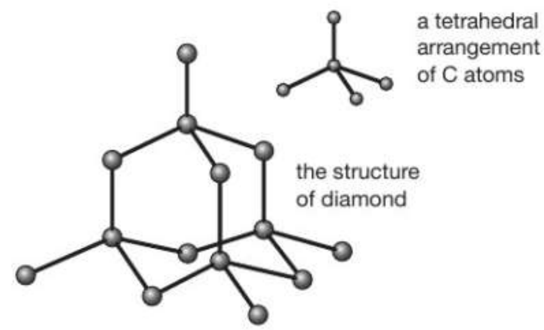
Some other physical properties of covalent bonds

- do not conduct electricity
- insoluble in water
- soluble in organic solvents

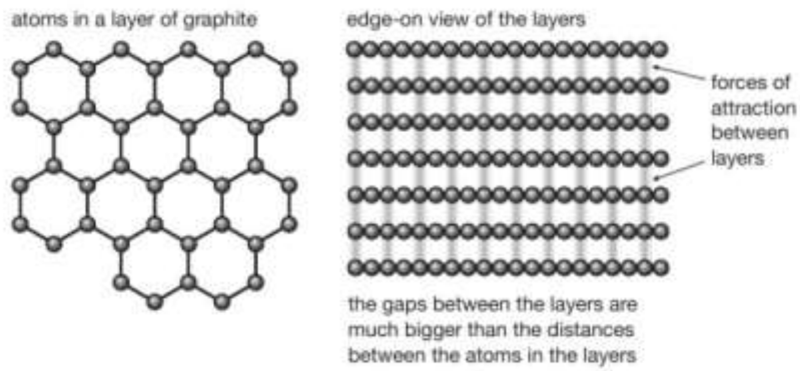
Giant organic structures

Diamond

Diamond is a form of pure carbon. Each carbon bond form four other strong bonds with other carbon atoms in a tetrahedral arrangement. It is not a molecule because the number of atoms joined up in a real diamond is completely variable and depends on the size of the crystal.
They have very high melding and boiling points. Diamonds are very hard. And they do not conduct electricity.



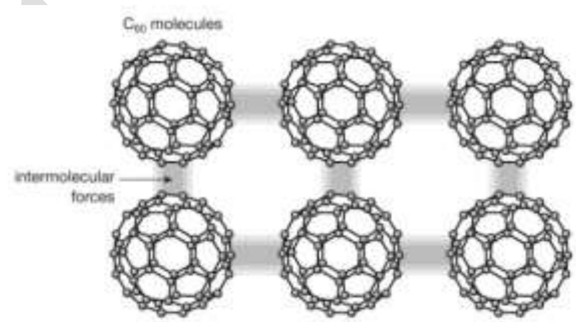
Graphite



A form of carbon. Graphite is a soft material. Used in pencils. Graphite also has high melting and boiling points. Also conducts electricity. The movement of delocalized electrons (the electrons that are free to move around) allow graphite to conduct electricity.

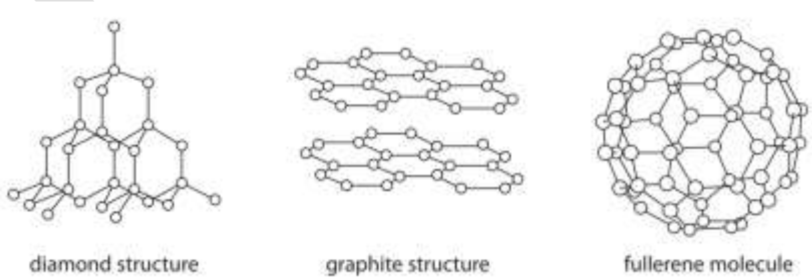
C_{60} fullerene

This another allotrope (different forms of the same element) of carbon. They have weak intermolecular forces between. Has a simple structure. C_{60} fullerene has lower melting and boiling point than diamond and graphite. Not hard as diamond and does not conduct electricity.



EXCERSISE

1. The diagram shows three different forms of carbon.



   a. Name the type of bond that exists between the carbon atoms in all three structures.

b.
   i. Explain why diamond has a very high melting point.
   ii. Fullerene has a simple molecular structure. Explain why it has a low melting point.
c. There are two theories used to explain why graphite can act as a solid lubricant.

   Theory A — The forces of attraction between the layers are weak, allowing the layers to slide over one another

   Theory B — Gas molecules are trapped between the layers allowing the layers to slide over one another.

   The table shows the ability of graphite to act as a lubricant in different locations.

Location	Ability to act as a lubricant
Earth's surface	good
high altitude	average
outer space	very poor

   Suggest which theory is supported by the evidence in the table. Give a reason for your choice.

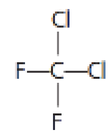
2. Ethene can be converted into many useful substances.
   a. Draw a dot and cross diagram to show the covalent bonding in a molecule of ethene. Only the outer electrons in each atom need to be shown.
   b. Compound X is made from ethene and is used in cars to prevent the engine coolant from freezing in cold weather.
      i. Compound X contains 38.7% carbon, 9.7% hydrogen and 51.6% oxygen by mass. Calculate the empirical formula of X.
      ii. The relative formula mass (M_r) of X is 62. What is the molecular formula of X?

3. Hydrogen chloride is formed in the reaction between hydrogen and chlorine. The equation for the reaction is

   $H_2 + Cl_2 \rightarrow 2HCl$

   a. Each molecule in this equation contains the same type of bonding. Name this type of bonding.
   b. The bonding in a hydrogen molecule is strong. Explain why the boiling point of hydrogen is low.
   c. Explain how the two atoms in a chlorine molecule are held together.
   d. Draw a dot and cross diagram to show the bonding in a hydrogen chloride molecule. Show only the outer electrons in each atom.
   e. Hydrogen chloride gas dissolves in water to form solution A. Hydrogen chloride gas dissolves in methylbenzene to form solution B. A teacher adds a piece of magnesium ribbon to each solution. Explain why she observes effervescence with solution A but not with solution B.

4. A sample of a chlorofluorocarbon (CFC) contains 0.24 g of carbon, 0.38 g of fluorine and 1.42 g of chlorine.
   a.
      i. Show, by calculation, that the empirical formula of the CFC is $CFCl_2$
      ii. The relative formula mass of the CFC is 204. Deduce the molecular formula of the CFC.
      iii. The displayed formula of another CFC is

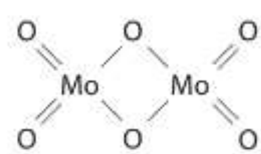


Draw a dot and cross diagram of this CFC.
Show only the outer electrons.

5. Molybdenum (Mo) is a metal. It is often used to make an alloy with iron. Like iron, it is extracted from its oxide. Unlike iron, it occurs mainly as its sulfide.
   a. Molybdenum sulfide is converted into molybdenum oxide by heating in air. The equation for this reaction is: $2MoS_2 + 7O_2 \rightarrow 2MoO_3 + 4SO_2$
      i. Why is molybdenum said to be oxidised in this reaction?
      ii. The sulfur dioxide formed in the reaction could form acid rain if it escaped into the atmosphere. Write a chemical equation for the formation of an acid from sulfur dioxide.
   b. The table shows the melting points of molybdenum oxide and sulfur dioxide.

	Melting point in °C
molybdenum oxide	800
sulfur dioxide	−75

The melting point indicates the type of bonding and structure in a compound.
      i. What is the type of bonding in a molecule of sulfur dioxide?
      ii. Explain why the melting point of sulfur dioxide is low.
      iii. The melting point of molybdenum oxide suggests that it has ionic bonding. However, it is often represented as a molecular structure. Deduce the molecular formula of molybdenum oxide as shown in this structure.



   c. The metallic structure of molybdenum gives it some typical properties.
      i. Describe the metallic structure of molybdenum.
      ii. Explain why molybdenum is a good conductor of electricity.
      iii. Explain why molybdenum is malleable.

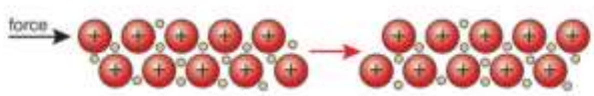
1.9 Metallic bonds

Metallic bonding

Metallic bonding is the electrostatic force of attraction between each positive ion and the delocalized electron. This holds the structure together.

Physical properties of metal

- high melting point
- conduct electricity
  this happens because the delocalized electron is free to move.
- Malleable



- Ductile

1.10 Electrolysis

Why things conduct electricity

- Metals
  Delocalized ions are free to move.
- Ionic compounds
  Conduct only when they are molten or aqueous. Ions are free to move in these states.
- Covalent compounds
  Generally, do not conduct electricity. Some exceptions are: ammonia.

Passing electricity through compounds: Electrolysis

<u>Electrolysis</u>
A chemical change caused by passing an electric current through a compound which is either molten or in solution.

<u>Electrolyte</u>
A liquid or a solution that undergoes electrolysis. Contain ions.

<u>Electrodes</u>
The electricity passes in and out through this.

<u>Inert</u>
Does not react with anything easily.

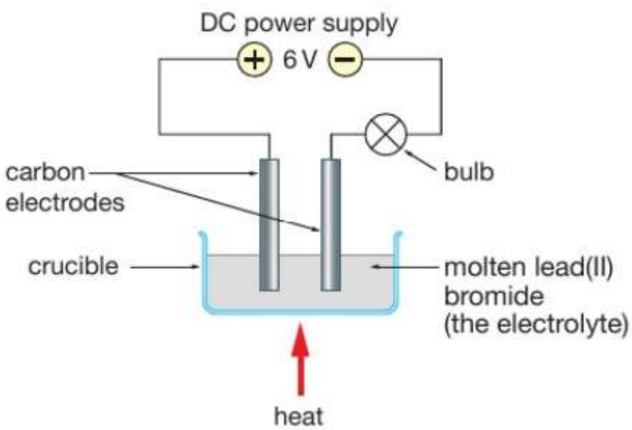
<u>Anode</u>
Positive electrode.

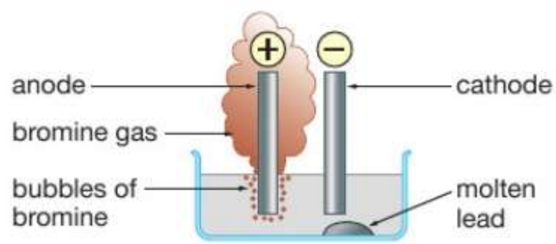
Cathode
Negative electrode.

Discharging an ion
Losing an it's charge.

Electrolysis of molten compounds

Electrolyzing molten lead bromide, $PbBr_2$





Explaining what's happening

- $Pb^{2+} + 2e^- \rightarrow Pb$
- $Br^- \rightarrow Br + e^-$
  $2Br \rightarrow Br_2$
  Which gives: $2Br^- \rightarrow Br_2 + 2e^-$

Electrolysis and redox

Oxidation happens when loosing and electron.
Reduction happens when gaining an electron

HINT
OILRIG:
Oxidation
Is
Loss of electrons
Reduction
Is
Gain of electrons

Electrolysis of other molten substances

Molten Sodium chloride
Reduction: $Na^+ + e^- \rightarrow Na$
Oxidation: $2Cl^- \rightarrow Cl_2 + 2e^-$

Molten Aluminum oxide
Reduction: $Al^{3+} + 3e^- \rightarrow Al$
Oxidation: $2O^{2-} \rightarrow O_2 + 4e$

<u>Molten Zinc chloride</u>

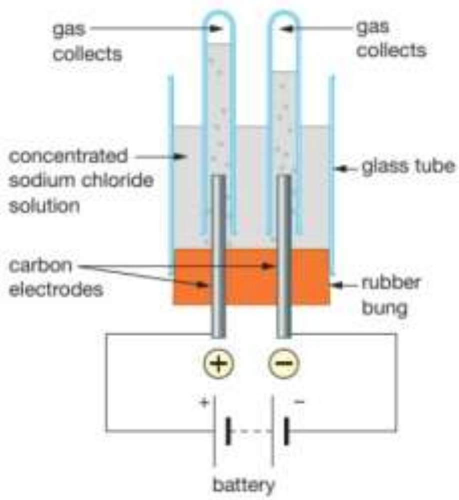
Reduction: $Zn^{2+} + 2e^- \rightarrow Zn$

Oxidation: $2Cl^- \rightarrow Cl_2 + 2e^-$

Generalizations that can be made:

- If you electrolyse a molten ionic compound only containing two elements, you will get the metal at the cathode and the non-metal at the anode.
- Reduction always occurs at the cathode and oxidation always occurs at the anode.

Electrolysis of aqueous solution



The procedure:

- Set up the apparatus as shown in Figure. The glass tube, rubber bung and electrodes together are sometimes called an electrolytic cell.
- Pour concentrated sodium chloride solution into the glass tube.
- Place a test-tube containing sodium chloride solution over each electrode. The test-tubes must not completely cover the electrodes or ions will be unable to flow and there will be no current.
- Connect the battery/powerpack to the electrodes.
- The experiment should be done in a fume cupboard (fume hood) or well-ventilated room because chlorine gas is poisonous.

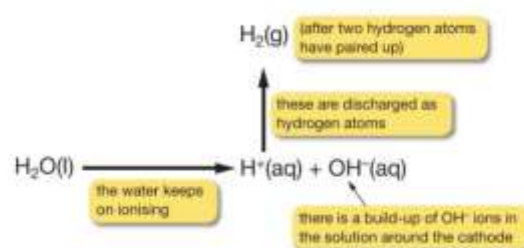
Electrolysis of sodium chloride

Anode: Chlorine

Cathode: Sodium

Water is called a weak electrolyte. It ionizes very slightly to give hydrogen ions and hydroxide ions:

$$H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$$



At the cathode

- $2H^+_{(aq)} + 2e^- \rightarrow H_{2\,(g)}$
- $2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

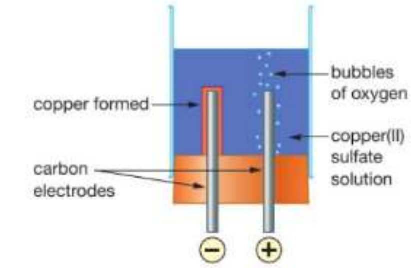
At the anode

- $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

The remaining solution

from NaCl:	Na^+	$\cancel{Cl^-}$
from H_2O:	$\cancel{H^+}$	OH^-

The electrolysis of copper sulfate solution using inert electrodes

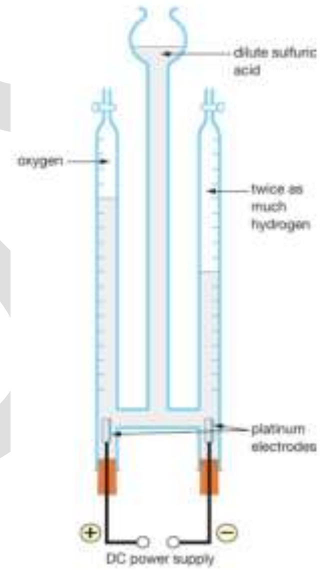


- Reduction: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$
- Oxidation: $4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + 4e^-$
- $2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$

from $CuSO_4$:	Cu^{2+}	SO_4^{2-}
from water:	H^+	OH^-

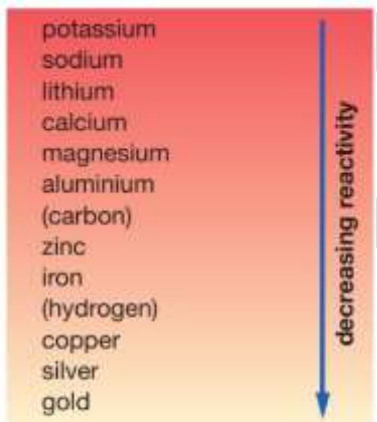
Electrolysis of dilute sulfuric acid using inert electrodes

- $2H^+_{(aq)} + 2e^- \rightarrow H_{2(g)}$
- $4OH^-_{(aq)} \rightarrow 2H_2O_{(l)} + O_{2(g)} + 4e^-$



Electrolysis of some other solutions using inert electrodes

How to work out what will happen



- If the metal is high in the reactivity series, you get hydrogen produced at the cathode instead of the metal.
- If the metal is below hydrogen in the reactivity series, you obtain the metal at the cathode.
- If you have solutions of halides (chlorides, bromides or iodides), you obtain the halogen (chlorine, bromine or iodine) at the anode. With other common negative ions (sulfate, nitrate, hydroxide), you obtain oxygen at the anode.

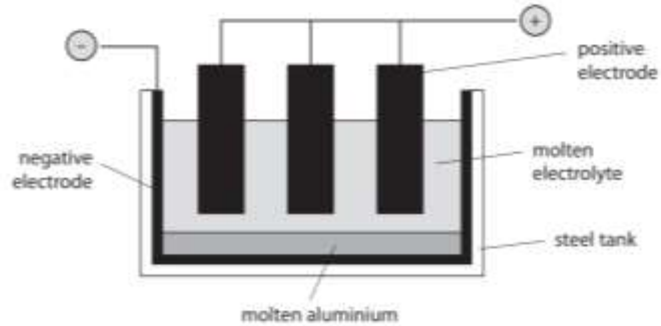
Cathode (–)			Anode (+)	
Solution	Product	Half-equation	Product	Half-equation
$KI(aq)$	hydrogen	$2H^+(aq) + 2e^- \rightarrow H_2(g)$	iodine	$2I^-(aq) \rightarrow I_2(aq) + 2e^-$
$MgBr_2(aq)$	hydrogen	$2H^+(aq) + 2e^- \rightarrow H_2(g)$	bromine	$2Br^-(aq) \rightarrow Br_2(aq) + 2e^-$
$H_2SO_4(aq)$	hydrogen	$2H^+(aq) + 2e^- \rightarrow H_2(g)$	oxygen	$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$
$CuSO_4(aq)$	copper	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	oxygen	$4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$

What would happen with non-electrolytes

For electrolysis to happen, there have to be ions present. The current in the external circuit can flow only if there are ions which can move and discharge.

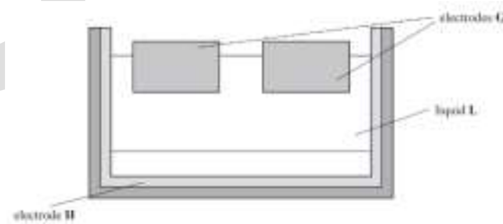
EXCERSISE

1. This question is about the extraction and uses of aluminium.



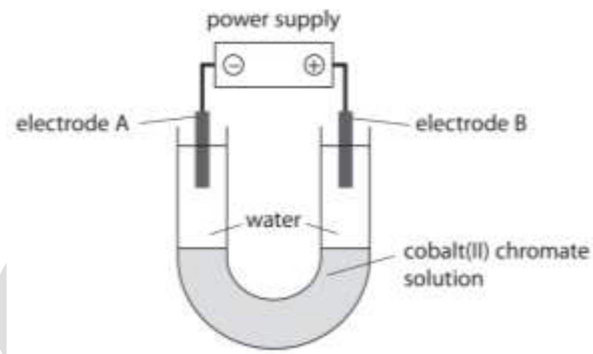
   a. Aluminium is extracted from aluminium oxide by electrolysis. What are the electrodes made of?
   b.
      i. Explain why the operating temperature would need to be very high if pure aluminium oxide were used as the electrolyte.
      ii. Describe how the operating temperature is kept low.
   c. The ionic half-equation for the reaction at the negative electrode is $Al^{3+} + 3e^{-} \rightarrow Al$
      What type of reaction is occurring at the negative electrode? Explain your answer.
   d. The waste gases escaping from the electrolysis cell contain carbon dioxide. Describe how the carbon dioxide is formed.
   e. Aluminium is used to make cans for food and drinks. State two properties of aluminium that make it suitable for this use. You should not refer to cost in your answers.

2. The diagram shows how aluminium is extracted in industry.



   a.
      i. Name the process used to extract aluminium.
      ii. Identify the element used to make the electrodes labelled G.
      iii. State whether electrode H is positive or negative.
      iv. Liquid L contains aluminium oxide and one other substance. Name this other substance and give one reason for its use in the extraction of aluminium.
   b. The product formed at electrode G reacts with the electrode to form carbon monoxide and carbon dioxide.
      i. Identify this product.
      ii. State why carbon monoxide is poisonous.
      iii. Describe a simple chemical test, and its result, for carbon dioxide.
   c. The uses of aluminium depend on its structure and physical properties.
      i. The strength of solid aluminium depends on the electrostatic force of attraction between two types of particle in its structure. Name these two types of particle.
      ii. Aluminium is described as ductile because it can easily be pulled into a wire. Explain, in terms of its structure, why it is ductile.
      iii. Explain, in terms of its structure, why aluminium is a good conductor of electricity
      iv. State a property that makes aluminium suitable for manufacturing aircraft bodies.

3. The apparatus shown in the diagram can be used to investigate the colours of the cobalt(II) ion (Co^{2+}) and the chromate ion (CrO_4^{2-}) in cobalt(II) chromate.
These are the results of the experiment.
   - a pink colour moves towards electrode A
   - a yellow colour moves towards electrode B



a. Explain how the results show that the chromate ion is yellow

b. Chromate ions in aqueous solution can be converted into dichromate ions ($Cr_2O_7^{2-}$) by the addition of hydrogen ions.
Balance the equation that represents this reaction.
$$\ldots\ldots CrO_4^{2-}(aq) + \ldots\ldots H^+(aq) \rightarrow \ldots\ldots Cr_2O_7^{2-}(aq) + \ldots\ldots H_2O(l)$$

c. When aqueous potassium chromate is added to aqueous lead(II) nitrate, a bright yellow precipitate is formed.
   i. Complete the equation for the reaction by inserting the missing state symbols.
   K_2CrO_4(...............) + $Pb(NO_3)_2$(...............) → 2KNO3(aq) + $PbCrO_4$(...............)
   ii. Describe how you could obtain a pure, dry sample of the insoluble solid from the final reaction mixture.