5.2 Equations for chemical reactions

Equations for two sample reactions

1. The reaction between carbon and oxygen  When carbon is heated in oxygen, they react together to form carbon dioxide. Carbon and oxygen are the reactants. Carbon dioxide is the product. You could show the reaction using a diagram, like this:

   ![Diagram of carbon and oxygen reaction]

   or by a word equation, like this:

   carbon + oxygen → carbon dioxide

   or by a symbol equation, which gives symbols and formulae:

   \[ C + O_2 \rightarrow CO_2 \]

2. The reaction between hydrogen and oxygen  Hydrogen and oxygen react together to give water. The diagram is:

   ![Diagram of hydrogen and oxygen reaction]

   and you can use it to write the symbol equation:

   \[ 2H_2 + O_2 \rightarrow 2H_2O \]

Symbol equations must be balanced

Now look at the number of atoms on each side of this equation:

\[ 2H_2 + O_2 \rightarrow 2H_2O \]

<table>
<thead>
<tr>
<th>On the left:</th>
<th>On the right:</th>
</tr>
</thead>
<tbody>
<tr>
<td>4 hydrogen atoms</td>
<td>4 hydrogen atoms</td>
</tr>
<tr>
<td>2 oxygen atoms</td>
<td>2 oxygen atoms</td>
</tr>
</tbody>
</table>

The number of each type of atoms is the same on both sides of the arrow. This is because atoms do not disappear during a reaction – they are just rearranged, as shown in the diagram of the molecules, in 2 above.

When the number of each type of atom is the same on both sides, the symbol equation is balanced. If it is not balanced, it is not correct.

Adding state symbols

Reactants and products may be solids, liquids, gases, or in solution. You can show their states by adding state symbols to the equations:

(s) for solid  (l) for liquid
(g) for gas  (aq) for aqueous solution (solution in water)

For the two reactions above, the equations with state symbols are:

\[ C (s) + O_2 (g) \rightarrow CO_2 (g) \]
\[ 2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l) \]

Note
Reactants are sometimes called reagents.
How to write the equation for a reaction

These are the steps to follow, when writing an equation:

1. Write the equation in words.
2. Now write it using symbols. Make sure all the formulae are correct.
3. Check that the equation is balanced, for each type of atom in turn. **Make sure you do not change any formulae.**
4. Add the state symbols.

**Example 1** Calcium burns in chlorine to form calcium chloride, a solid. Write an equation for the reaction, using the steps above.

1. calcium + chlorine → calcium chloride
2. Ca + Cl₂ → CaCl₂
3. Ca: 1 atom on the left and 1 atom on the right.
   Cl: 2 atoms on the left and 2 atoms on the right.
   The equation is balanced.
4. Ca (s) + Cl₂ (g) → CaCl₂ (s)

**Example 2** Hydrogen chloride is formed by burning hydrogen in chlorine. Write an equation for the reaction.

1. hydrogen + chlorine → hydrogen chloride
2. H₂ + Cl₂ → HCl
3. H: 2 atoms on the left and 1 atom on the right.
   Cl: 2 atoms on the left and 1 atom on the right.
   The equation is not balanced. It needs another molecule of hydrogen chloride on the right. So a 2 is put in front of the HCl.
   H₂ + Cl₂ → 2HCl
   The equation is now balanced. Do you agree?
4. H₂ (g) + Cl₂ (g) → 2HCl (g)

**Example 3** Magnesium burns in oxygen to form magnesium oxide, a white solid. Write an equation for the reaction.

1. magnesium + oxygen → magnesium oxide
2. Mg + O₂ → MgO
3. Mg: 1 atom on the left and 1 atom on the right.
   O: 2 atoms on the left and 1 atom on the right.
   The equation is not balanced. Try this:
   Mg + O₂ → 2MgO (The 2 goes in front of the MgO.)
   Another magnesium atom is now needed on the left:
   2Mg + O₂ → 2MgO
   The equation is balanced.
4. 2Mg (s) + O₂ (g) → 2MgO (s)

Q

1. What do + and → mean, in an equation?
2. Balance the following equations:
   a. Na (s) + Cl₂ (g) → NaCl (s)
   b. H₂ (g) + I₂ (g) → HI (g)
   c. Na (s) + H₂O (l) → NaOH (aq) + H₂ (g)
   d. NH₃ (g) → N₂ (g) + H₂ (g)
   e. C (s) + CO₂ (g) → CO (g)
   f. Al (s) + O₂ (g) → Al₂O₃ (s)
3. Aluminium burns in chlorine to form aluminium chloride, AlCl₃, a solid. Write a balanced equation for the reaction.
5.3 The masses of atoms, molecules, and ions

Relative atomic mass

A single atom weighs hardly anything. You can’t use scales to weigh it. But scientists do need a way to compare the masses of atoms. So this is what they did.

First, they chose an atom of carbon-12 to be the standard atom. They fixed its mass as exactly 12 atomic mass units. (It has 6 protons and 6 neutrons, as shown on the right. They ignored the electrons.) Then they compared all the other atoms with this standard atom, in a machine called a mass spectrometer, and found values for their masses. Like this:

This is the standard atom, $^{12}\text{C}$ or carbon-12. Its mass is taken as exactly 12.

This magnesium atom is twice as heavy as the carbon-12 atom. So its mass is 24.

This hydrogen atom has only one-twelfth the mass of the carbon-12 atom. So its mass is 1.

The mass of an atom compared with the carbon-12 atom is called its relative atomic mass, or $A_r$.

The small r stands for relative to the mass of a carbon-12 atom.

So the $A_r$ of hydrogen is 1, and the $A_r$ of magnesium is 24.

$A_r$ and isotopes

As you saw on page 34, the atoms of an element are not always identical. Some may have extra neutrons. Different atoms of the same element are called isotopes. Chlorine has two isotopes:

<table>
<thead>
<tr>
<th>Name</th>
<th>Protons</th>
<th>Neutrons</th>
<th>Nucleon number</th>
<th>% of chlorine atoms like this</th>
</tr>
</thead>
<tbody>
<tr>
<td>chlorine-35</td>
<td>17</td>
<td>18</td>
<td>35</td>
<td>75%</td>
</tr>
<tr>
<td>chlorine-37</td>
<td>17</td>
<td>20</td>
<td>37</td>
<td>25%</td>
</tr>
</tbody>
</table>

We need to take all the natural isotopes of an element into account, to work out the relative atomic mass. This is the formula to use:

**relative atomic mass ($A_r$) of an element** =

\[
\frac{\text{mass}}{12} = \left( \frac{\% \times \text{nucleon number for the first isotope}}{100} \right) + \left( \frac{\% \times \text{nucleon number for the second isotope}}{100} \right) + \ldots \text{and so on, for all its natural isotopes}
\]

The calculation for chlorine is given on the right. It shows that the relative atomic mass of chlorine is 35.5.

The relative atomic mass ($A_r$) of an element is the average mass of its isotopes compared to an atom of carbon-12.

For most elements, $A_r$ is very close to a whole number. It is usually rounded off to a whole number, to make calculations easier.
Finding the masses of molecules and ions

Using \( A_r \) values, it is easy to work out the mass of any molecule or group of ions. Read the blue panel on the right above, then look at these examples:

**Finding the mass of an ion**

mass of sodium atom = 23, so 
mass of sodium ion = 23 
since a sodium ion is just a sodium atom minus an electron (which has negligible mass).

An ion has the same mass as the atom from which it is made.

**Finding the mass of a molecule**

Hydrogen gas is made of molecules. Each molecule contains 2 hydrogen atoms, so its mass is 2. \((2 \times 1 = 2)\)

The formula for water is \( H_2O \). Each water molecule contains 2 hydrogen atoms and 1 oxygen atom, so its mass is 18. \((2 \times 1 + 16 = 18)\)

If the substance is made of molecules, its mass found in this way is called the **relative molecular mass**, or \( M_r \). So the \( M_r \) for hydrogen is 2, and for water is 18.

But if the substance is made of **ions**, its mass is called the **relative formula mass**, which is also \( M_r \) for short. So the \( M_r \) for NaCl is 58.5.

This table gives two more examples of how to calculate \( M_r \) values.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>Atoms in formula</th>
<th>( A_r ) of atoms</th>
<th>( M_r )</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonia</td>
<td>( \text{NH}_3 )</td>
<td>( 1 \text{N} \quad 3 \text{H} )</td>
<td>( N = 14 \quad H = 1 )</td>
<td>( 1 \times 14 = 14 \frac{3 \times 1}{3} \frac{1}{2} ) ( \text{Total} = 17 )</td>
</tr>
<tr>
<td>magnesium nitrate</td>
<td>( \text{Mg(NO}_3\text{)}_2 )</td>
<td>( 1 \text{Mg} \quad 2 \text{N} \quad 6 \text{O} )</td>
<td>( \text{Mg} = 24 \quad \text{N} = 14 \quad \text{O} = 16 )</td>
<td>( 1 \times 24 = 24 \frac{2 \times 14}{2} \frac{6 \times 16}{6} ) ( \text{Total} = 148 )</td>
</tr>
</tbody>
</table>

**Questions**

1. **a** What does relative atomic mass mean?
   **b** Why does it have the word relative?
2. What is the \( A_r \) of the iodide ion, \( \text{I}^- \)?
3. The relative molecular mass and formula mass are both called \( M_r \) for short. What is the difference between them?
4. Work out the \( M_r \) for each of these, and say whether it is the relative molecular mass or the relative formula mass:
   **a** oxygen, \( \text{O}_2 \)  
   **b** iodine, \( \text{I}_2 \)  
   **c** methane, \( \text{CH}_4 \)  
   **d** chlorine, \( \text{Cl}_2 \)  
   **e** butane, \( \text{C}_4\text{H}_{10} \)  
   **f** ethanol, \( \text{C}_2\text{H}_5\text{OH} \)  
   **g** ammonium sulfate (\( \text{NH}_4\text{SO}_4 \))
5.4 Some calculations about masses and %

Two laws of chemistry

If you know the actual amounts of two substances that react, you can:

- predict other amounts that will react
- say how much product will form.

You just need to remember these two laws of chemistry:

1. **Elements always react in the same ratio, to form a given compound.**
   - For example, when carbon burns in oxygen to form carbon dioxide:
     - 6 g of carbon combines with 16 g of oxygen, so
     - 12 g of carbon will combine with 32 g of oxygen, and so on.

2. **The total mass does not change, during a chemical reaction.**
   - So total mass of reactants = total mass of products.
   - So 6 g of carbon and 16 g of oxygen give 22 g of carbon dioxide.
   - 12 g of carbon and 32 g of oxygen give 44 g of carbon dioxide.

Calculating quantities

Calculating quantities is quite easy, using the laws above.

**Example**  64 g of copper reacts with 16 g of oxygen to give the black compound copper(II) oxide.

a. What mass of copper will react with 32 g of oxygen?
   - 64 g of copper reacts with 16 g of oxygen, so
   - 2 \times 64 g or 128 g of copper will react with 32 g of oxygen.

b. What mass of oxygen will react with 32 g of copper?
   - 16 g of oxygen reacts with 64 g of copper, so
   - \( \frac{16}{2} \) or 8 g of oxygen will react with 32 g of copper.

c. What mass of copper(II) oxide will be formed, in b?
   - 40 g of copper(II) oxide will be formed. (32 + 8 = 40)

d. How much copper and oxygen will give 8 g of copper(II) oxide?
   - 64 g of copper and 16 g of oxygen give 80 g of copper(II) oxide, so
   - \( \frac{64}{10} \) of copper and \( \frac{16}{10} \) g of oxygen will give 8 g of copper(II) oxide, so
   - 6.4 g of copper and 1.6 g of oxygen are needed.

Percentages: a reminder

Calculations in chemistry often involve percentages. Remember:

- The full amount of anything is 100%.
- To change a fraction to a %, just multiply it by 100.

**Example 1**  Change the fractions \( \frac{1}{2} \) and \( \frac{18}{25} \) to percentages.

\[ \frac{1}{2} \times 100 = 50\% \quad \frac{18}{25} \times 100 = 72\% \]

**Example 2**  Give 19 % as a fraction.

\[ 19\% = \frac{19}{100} \]
Calculating the percentage composition of a compound

The **percentage composition** of a compound tells you how much of each element it contains, *as a percentage of the total mass*. This is how to work it out:

1. Write down the formula of the compound.
2. Using \( A_r \) values, work out its molecular or formula mass (\( M_r \)).
3. Write the mass of the element as a fraction of the \( M_r \).
4. Multiply the fraction by 100, to give a percentage.

**Example**  Calculate the percentage of oxygen in sulfur dioxide.

1. The formula of sulfur dioxide is \( \text{SO}_2 \).
2. The \( M_r \) of the compound is 64, as shown on the right.
3. Mass of oxygen as a fraction of the total \( \frac{32}{64} \)
4. Mass of oxygen as a percentage of the total \( \frac{32}{64} \times 100 = 50\% \)
   
   So the compound is **50\% oxygen**.
   
   This means it is also 50\% sulfur (100\% – 50\% = 50\%).

Calculating \% purity

A **pure** substance has nothing else mixed with it.

But substances often contain unwanted substances, or **impurities**.

Purity is usually given as a percentage. This is how to work it out:

\[
\text{\% purity of a substance} = \frac{\text{mass of pure substance in it}}{\text{total mass}} \times 100\%
\]

**Example**  Impure copper is refined (purified), to obtain pure copper for use in computers. 20 tonnes of copper gave 18 tonnes of pure copper, on refining.

a  What was the \% purity of the copper before refining?

\[
\text{\% purity of the copper} = \frac{18 \text{ tonnes}}{20 \text{ tonnes}} \times 100\% = 90\%
\]

So the copper was **90\% pure**.

b  How much pure copper will 50 tonnes of the impure copper give?

The impure copper is 90\% pure.

\[
90\% = \frac{90}{100}
\]

So 50 tonnes of it will give \( \frac{90}{100} \times 50 \text{ tonnes} \) or **45 tonnes** of pure copper.

---

1. Magnesium burns in chlorine to give magnesium chloride, \( \text{MgCl}_2 \). In an experiment, 24 g of magnesium was found to react with 71 g of chlorine.
   a  How much magnesium chloride was obtained in the experiment?
   b  How much chlorine will react with 12 g of magnesium?
   c  How much magnesium chloride will form, in b?

2. Methane has the formula \( \text{CH}_4 \). Work out the \% of carbon and hydrogen in it. \( A_r: \text{C} = 12, \text{H} = 1 \)

3. In an experiment, a sample of lead(II) bromide was made. It weighed 15 g. But the sample was found to be impure. In fact it contained only 13.5 g of lead(II) bromide.
   a  Calculate the \% purity of the sample.
   b  What mass of impurity was present in the sample?
Checkup on Chapter 5

Revision checklist

Core curriculum

Make sure you can …

☐ name a simple compound, when you are given the names of the two elements that form it

☐ work out the formula of a compound from a drawing of its structure

☐ work out the formula of a simple compound when you know the two elements in it, by balancing their valencies

☐ work out the formula of an ionic compound by balancing the charges of the ions, so that the total charge is zero

☐ write the equation for a reaction:
  – as a word equation
  – as a symbol equation

☐ balance a symbol equation

☐ say what the state symbols mean: (s), (l), (g), (aq)

☐ define aqueous

☐ explain that the carbon-12 atom is taken as the standard, for working out masses of atoms

☐ say what these two symbols mean: \( \text{A}_r \), \( \text{M}_r \)

☐ work out \( \text{M}_r \) values, given the \( \text{A}_r \) values

☐ explain the difference between relative formula mass and relative molecular mass (both known as \( M_r \) for short)

☐ explain that the \( \text{A}_r \) value of an element is the average value for all its isotopes

☐ predict other amounts of reactants that will react, when you are given some actual amounts

☐ calculate:
  – the mass of a product, when you are given the masses of the reactants that combine to form it
  – the percentage of an element in a compound, using the formula and \( \text{A}_r \) values
  – the percentage purity of a substance, when you are given the total mass of the impure substance, and the amount of the pure substance in it

Questions

Core curriculum

If you are not sure about symbols for the elements, you can check the Periodic Table on page 314.

1. Write the formulae for these compounds:
   a. water
   b. carbon monoxide
   c. carbon dioxide
   d. sulfur dioxide
   e. sulfur trioxide
   f. sodium chloride
   g. magnesium chloride
   h. hydrogen chloride
   i. methane
   j. ammonia

2. You can work out the formula of a compound from the ratio of the different atoms in it. Sodium carbonate has the formula \( \text{Na}_2\text{CO}_3 \) because it contains 2 atoms of sodium for every 1 atom of carbon and 3 atoms of oxygen. Deduce the formula for each compound a to h:

<table>
<thead>
<tr>
<th>Compound</th>
<th>Ratio in which the atoms are combined in it</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. lead oxide</td>
<td>1 of lead, 2 of oxygen</td>
</tr>
<tr>
<td>b. lead oxide</td>
<td>3 of lead, 4 of oxygen</td>
</tr>
<tr>
<td>c. potassium nitrate</td>
<td>1 of potassium, 1 of nitrogen, 3 of oxygen</td>
</tr>
<tr>
<td>d. nitrogen oxide</td>
<td>2 of nitrogen, 1 of oxygen</td>
</tr>
<tr>
<td>e. nitrogen oxide</td>
<td>2 of nitrogen, 4 of oxygen</td>
</tr>
<tr>
<td>f. sodium hydrogen carbonate</td>
<td>1 of sodium, 1 of hydrogen, 1 of carbon, 3 of oxygen</td>
</tr>
<tr>
<td>g. sodium sulfate</td>
<td>2 of sodium, 1 of sulfur, 4 of oxygen</td>
</tr>
<tr>
<td>h. sodium thiosulfate</td>
<td>2 of sodium, 2 of sulfur, 3 of oxygen</td>
</tr>
</tbody>
</table>

3. For each compound, write down the ratio of atoms present:
   a. copper(II) oxide, \( \text{CuO} \)
   b. copper(I) oxide, \( \text{Cu}_2\text{O} \)
   c. aluminium chloride, \( \text{AlCl}_3 \)
   d. nitric acid, \( \text{HNO}_3 \)
   e. calcium hydroxide, \( \text{Ca(OH)}_2 \)
   f. ethanoic acid, \( \text{CH}_3\text{COOH} \)
   g. ammonium nitrate, \( \text{NH}_4\text{NO}_3 \)
   h. ammonium sulfate, \( \text{(NH}_4\text{)}_2\text{SO}_4 \)
   i. sodium phosphate, \( \text{Na}_3\text{(PO}_4\text{)}_2 \)
   j. hydrated iron(II) sulfate, \( \text{FeSO}_4\cdot7\text{H}_2\text{O} \)
   k. hydrated cobalt(II) chloride, \( \text{CoCl}_2\cdot6\text{H}_2\text{O} \)
4 Write the chemical formulae for the compounds with the structures shown below:

a. $H - Br$

b. $Cl - P - Cl$

c. $H - O - O - H$

d. $H - C \equiv C - H$

e. $H - N \equiv N - H$

f. $H - O - S - O - H$

g. $Br - Cl - Br$

h. $O - S - O - H$

5 This shows the structure of an ionic compound.

a. Name the compound.

b. What is the simplest formula for it?

6 This shows the structure of a molecular compound.

a. Name the compound.

b. What is the simplest formula for it?

7 Write these as word equations:

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

b. $Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + CO_2 + H_2O$

c. $2Mg + CO_2 \rightarrow 2MgO + C$

d. $ZnO + C \rightarrow Zn + CO$

e. $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$

f. $CuO + 2HNO_3 \rightarrow Cu(NO_3)_2 + H_2O$

8 Balance these equations:

a. $N_2 + ... O_2 \rightarrow ... NO_2$

b. $K_2CO_3 + ... HCl \rightarrow ... KCl + CO_2 + H_2O$

c. $C_2H_6 + ... O_2 \rightarrow ... CO_2 + 4H_2O$

d. $Fe_2O_3 + ... CO \rightarrow ... Fe + ... CO_2$

e. $Ca(OH)_2 + ... HCl \rightarrow CaCl_2 + ... H_2O$

f. $2Al + ... HCl \rightarrow 2AlCl_3 + ... H_2$

9 Copy and complete these equations:

a. $MgSO_4 + ... \rightarrow MgSO_4.6H_2O$

b. $... C + \rightarrow ... 2CO$

c. $2CuO + C \rightarrow 2Cu + ...$

d. $C_2H_6 \rightarrow ... + H_2$

e. $ZnO + C \rightarrow Zn + ...$

f. $NiCO_3 \rightarrow NiO + ...$

g. $CO_2 + ... \rightarrow CH_4 + O_2$

h. $NaOH + HNO_3 \rightarrow NaNO_3 + ...$

i. $C_2H_6 \rightarrow C_2H_4 + ...$

10 Calculate $M_r$ for these compounds. ($A_r$ values are given at the top of page 315.)

a. water, $H_2O$

b. ammonia, $NH_3$

c. ethanol, $CH_3CH_2OH$

d. sulfur trioxide, $SO_3$

e. sulfuric acid, $H_2SO_4$

f. hydrogen chloride, $HCl$

g. phosphorus(V) oxide, $P_2O_5$

11 Calculate the relative formula mass for these ionic compounds. ($A_r$ values are given on page 315.)

a. magnesium oxide, $MgO$

b. lead sulfide, $PbS$

c. calcium fluoride, $CaF_2$

d. sodium chloride, $NaCl$

e. silver nitrate, $AgNO_3$

f. ammonium sulfate, $(NH_4)_2SO_4$

g. potassium carbonate, $K_2CO_3$

h. hydrated iron(II) sulfate, $FeSO_4.7H_2O$

12 Iron reacts with excess sulfuric acid to give iron(II) sulfate. The equation for the reaction is: $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$

5 g of iron gives 13.6 g of iron(II) sulfate.

a. Using excess acid, how much iron(II) sulfate can be obtained from:

i. 10 g of iron?

ii. 1 g of iron?

b. How much iron will be needed to make 136 g of iron(II) sulfate?

c. A 10 g sample of impure iron(II) sulfate contains 8 g of iron(II) sulfate. Calculate the percentage purity of the iron(II) sulfate.

13 Aluminium is extracted from the ore bauxite, which is impure aluminium oxide.

1 tonne (1000 kg) of the ore was found to have this composition:

- aluminium oxide 825 kg
- iron(III) oxide 100 kg
- sand 75 kg

a. What percentage of this ore is impurities?

b. 1 tonne of the ore gives 437 kg of aluminium.

i. How much aluminium will be obtained from 5 tonnes of the ore?

ii. What mass of sand is in this 5 tonnes?

C. What will the percentage of aluminium oxide in the ore be, if all the iron(III) oxide is removed, leaving only the aluminium oxide and sand?
What is a mole?
As you saw on page 70, the masses of atoms are found by comparing them with the carbon-12 atom:

This is an atom of carbon-12. It is chosen as the standard atom. Its $A_r$ is taken as 12. Then other atoms are compared with it.

This is a magnesium atom. It is twice as heavy as a carbon-12 atom, so its $A_r$ is 24. So it follows that ...

... 24 g of magnesium contains the same number of atoms as 12 g of carbon-12. 24 g of magnesium is called a mole of magnesium atoms.

A mole of a substance is the amount that contains the same number of units as the number of carbon atoms in 12 grams of carbon-12. These units can be atoms, or molecules, or ions, as you will see.

The Avogadro constant
Thanks to the work of the Italian scientist Avogadro, we know that 12 g of carbon-12 contains $602,000,000,000,000,000,000$ carbon atoms!

This huge number is called the Avogadro constant. It is written in a short way as $6.02 \times 10^{23}$. (The $10^{23}$ tells you to move the decimal point 23 places to the right, to get the full number.)

So 1 mole of magnesium atoms contains $6.02 \times 10^{23}$ magnesium atoms.

More examples of moles

<table>
<thead>
<tr>
<th>Substance</th>
<th>Formula</th>
<th>$M_r$</th>
<th>$A_r$</th>
<th>Mass</th>
<th>Mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>Na</td>
<td></td>
<td>23</td>
<td>23 g</td>
<td>1 mole of sodium atoms</td>
</tr>
<tr>
<td>Iodine</td>
<td>I$_2$</td>
<td>254</td>
<td></td>
<td>254 g</td>
<td>1 mole of iodine molecules</td>
</tr>
<tr>
<td>Water</td>
<td>H$_2$O</td>
<td>18</td>
<td></td>
<td>18 g</td>
<td>1 mole of water molecules</td>
</tr>
</tbody>
</table>

So you can see that:
One mole of a substance is obtained by weighing out the $A_r$ or $M_r$ of the substance, in grams.
Finding the mass of a mole

You can find the mass of one mole of any substance by these steps:

1. Write down the symbol or formula of the substance.
2. Find its \( A_r \) or \( M_r \).
3. Express that mass in grams (g).

This table shows three more examples:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Symbol or formula</th>
<th>( A_r )</th>
<th>( M_r )</th>
<th>Mass of 1 mole</th>
</tr>
</thead>
<tbody>
<tr>
<td>helium</td>
<td>He</td>
<td>He = 4</td>
<td>exists as single atoms</td>
<td>4 grams</td>
</tr>
<tr>
<td>oxygen</td>
<td>O₂</td>
<td>O = 16</td>
<td>2 × 16 = 32</td>
<td>32 grams</td>
</tr>
<tr>
<td>ethanol</td>
<td>( \text{C}_2\text{H}_5\text{OH} )</td>
<td>C = 12, H = 1, O = 16</td>
<td>( \frac{2 × 12}{16} = 2 \frac{1}{2}, \frac{6 × 1}{1} = 6, \frac{1 × 16}{16} = 1 \frac{1}{2} )</td>
<td>46 grams</td>
</tr>
</tbody>
</table>

Some calculations on the mole

These equations will help you:

\[
\text{Mass of a given number of moles} = \text{mass of 1 mole} \times \text{number of moles}
\]

\[
\text{Number of moles in a given mass} = \frac{\text{mass}}{\text{mass of 1 mole}}
\]

Example 1  Calculate the mass of 0.5 moles of bromine atoms.
The \( A_r \) of bromine is 80, so 1 mole of bromine atoms has a mass of 80 g.
So 0.5 moles of bromine atoms has a mass of \( 0.5 \times 80 \) g, or 40 g.

Example 2  Calculate the mass of 0.5 moles of bromine molecules.
A bromine molecule contains 2 atoms, so its \( M_r \) is 160.
So 0.5 moles of bromine molecules has a mass of \( 0.5 \times 160 \) g, or 80 g.

Example 3  How many moles of oxygen molecules are in 64 g of oxygen?
The \( M_r \) of oxygen is 32, so 32 g of it is 1 mole.
Therefore 64 g is \( \frac{64}{32} \) moles, or 2 moles of oxygen molecules.

Q

1. How many atoms are in 1 mole of atoms?
2. How many molecules are in 1 mole of molecules?
3. What name is given to the number \( 6.02 \times 10^{23} \)?
4. Find the mass of 1 mole of:
   a. hydrogen atoms
   b. iodine atoms
   c. chlorine atoms
   d. chlorine molecules
5. Find the mass of 2 moles of:
   a. oxygen atoms
   b. oxygen molecules
6. Find the mass of 3 moles of ethanol, \( \text{C}_2\text{H}_5\text{OH} \).
7. How many moles of molecules are there in:
   a. 18 grams of hydrogen, \( \text{H}_2 \)?
   b. 54 grams of water?
8. Sodium chloride is made up of \( \text{Na}^+ \) and \( \text{Cl}^- \) ions.
   a. How many sodium ions are there in 58.5 g of sodium chloride? \( (A_r: \text{Na} = 23; \text{Cl} = 35.5) \)
   b. What is the mass of 1 mole of chloride ions?
Calculations from equations, using the mole

What an equation tells you
When carbon burns in oxygen, the reaction can be shown as:

\[ \text{C (s) + O}_2 (g) \rightarrow \text{CO}_2 (g) \]

or in a short way, using the symbol equation:

\[ \text{C} (s) + \text{O}_2 (g) \rightarrow \text{CO}_2 (g) \]

This equation tells you that:

- 1 carbon atom reacts with 1 molecule of oxygen to give 1 molecule of carbon dioxide

Now suppose there is 1 mole of carbon atoms. Then we can say that:

- 1 mole of carbon atoms reacts with 1 mole of oxygen molecules to give 1 mole of carbon dioxide molecules

So from the equation, we can tell how many moles react.

But moles can be changed to grams, using \( A_r \) and \( M_r \).

The \( A_r \) values are: \( \text{C} = 12, \text{O} = 16 \).

So the \( M_r \) values are: \( \text{O}_2 = 32, \text{CO}_2 = (12 + 32) = 44 \), and we can write:

\[ 12 \text{ g of carbon} \text{ reacts with } 32 \text{ g of oxygen} \rightarrow 44 \text{ g of carbon dioxide} \]

Since substances always react in the same ratio, this also means that:

\[ 6 \text{ g of carbon} \text{ reacts with } 16 \text{ g of oxygen} \rightarrow 22 \text{ g of carbon dioxide} \]

and so on.

So we have gained a great deal of information from the equation. In fact you can obtain the same information from any equation.

From the equation for a reaction you can tell:

- how many moles of each substance take part
- how many grams of each substance take part.

Reminder: the total mass does not change
Look what happens to the total mass, during the reaction above:

- mass of carbon and oxygen at the start: \( 12 \text{ g} + 32 \text{ g} = 44 \text{ g} \)
- mass of carbon dioxide at the end: \( 44 \text{ g} \)

The total mass has not changed, during the reaction. This is because no atoms have disappeared. They have just been rearranged.

That is one of the two laws of chemistry that you met on page 72:
**The total mass does not change, during a chemical reaction.**
Calculating masses from equations

These are the steps to follow:

1. Write the balanced equation for the reaction. (It gives moles.)
2. Write down the $A_r$ or $M_r$ for each substance that takes part.
3. Using $A_r$ or $M_r$, change the moles in the equation to grams.
4. Once you know the theoretical masses from the equation, you can then find any actual mass.

**Example** Hydrogen burns in oxygen to form water. What mass of oxygen is needed for 1 g of hydrogen, and what mass of water is obtained?

1. The equation for the reaction is: $2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l)$
3. So, for the equation, the amounts in grams are:

   $2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l)$
   $2 \times 2 \, g \quad 32 \, g \quad 2 \times 18 \, g \quad \text{or}$
   $4 \, g \quad 32 \, g \quad 36 \, g$

4. But you start with only 1 g of hydrogen, so the actual masses are:

   $1 \, g \quad 32/4 \, g \quad 36/4 \, g \quad \text{or}$
   $1 \, g \quad 8 \, g \quad 9 \, g$

So 1 g of hydrogen needs 8 g of oxygen to burn, and gives 9 g of water.

**Working out equations, from masses**

If you know the actual masses that react, you can work out the equation for the reaction. Just change the masses to moles.

**Example** Iron reacts with a solution of copper(II) sulfate ($CuSO_4$) to give copper and a solution of iron sulfate. The formula for the iron sulfate could be either $FeSO_4$ or $Fe_2(SO_4)_3$. 1.4 g of iron gave 1.6 g of copper. Write the correct equation for the reaction.

1. $A_r$: Fe = 56, Cu = 64.
2. Change the masses to moles of atoms:

   $\frac{1.4}{56}$ moles of iron atoms gave $\frac{1.6}{64}$ moles of copper atoms, or
   0.025 moles of iron atoms gave 0.025 moles of copper atoms, so
   1 mole of iron atoms gave 1 mole of copper atoms.
3. So the equation for the reaction must be:

   $Fe + CuSO_4 \rightarrow Cu + FeSO_4$
4. Add the state symbols to complete it:

   $Fe (s) + CuSO_4 (aq) \rightarrow Cu (s) + FeSO_4 (aq)$

**Q**

1. The reaction between magnesium and oxygen is:

   $2Mg (s) + O_2 (g) \rightarrow 2MgO (s)$
   a. Write a word equation for the reaction.
   b. How many moles of magnesium atoms react with 1 mole of oxygen molecules?
   c. The $A_r$ values are: Mg = 24, O = 16. How many grams of oxygen react with:
      i. 48 g of magnesium? ii. 12 g of magnesium?

2. Copper(II) carbonate breaks down on heating, like this:

   $CuCO_3 (s) \rightarrow CuO (s) + CO_2 (g)$
   a. Write a word equation for the reaction.
   b. Find the mass of 1 mole of each substance taking part in the reaction. ($A_r$: Cu = 64, C = 12, O = 16.)
   c. When 31 g of copper(II) carbonate is used:
      i. how many grams of carbon dioxide form?
      ii. what mass of solid remains after heating?
6.3 Reactions involving gases

A closer look at some gases

Imagine five very large flasks, each with a volume of 24 dm$^3$. Each is filled with a different gas. Each gas is at room temperature and pressure, or rtp. (We take room temperature and pressure as the standard conditions for comparing gases; rtp is 20 °C and 1 atmosphere.)

If you weighed the gas in the five flasks, you would discover something amazing. There is exactly 1 mole of each gas!

So we can conclude that:

1 mole of every gas occupies the same volume, at the same temperature and pressure. At room temperature and pressure, this volume is 24 dm$^3$.

This was discovered by Avogadro, in 1811. So it is often called Avogadro’s Law. It does not matter whether a gas exists as atoms or molecules, or whether its atoms are large or small. The law still holds.

The volume occupied by 1 mole of a gas is called its molar volume.

The molar volume of a gas is 24 dm$^3$ at rtp.

Another way to look at it

Look at these two gas jars. A is full of nitrogen dioxide, NO$_2$. B is full of oxygen, O$_2$.

The two gas jars have identical volumes, and the gases are at the same temperature and pressure.

You cannot see the gas molecules – let alone count them. But, from Avogadro’s Law, you can say that the two jars contain the same number of molecules.

Remember

24 dm$^3$ = 24 litres
= 24 000 cm$^3$

Imagine a ball about 36 cm in diameter. Its volume is about 24 dm$^3$. 

www.aswarphysics.weebly.com
Calculating gas volumes from moles and grams

Avogadro’s Law makes it easy to work out the volumes of gases.

**Example 1** What volume does 0.25 moles of a gas occupy at rtp?
1 mole occupies 24 dm$^3$ so
0.25 moles occupies $0.25 \times 24 \text{ dm}^3 = 6 \text{ dm}^3$
so 0.25 moles of any gas occupies 6 dm$^3$ (or 6000 cm$^3$) at rtp.

**Example 2** What volume does 22 g of carbon dioxide occupy at rtp?
$M_r$ of carbon dioxide = 44, so
44 g = 1 mole, so
22 g = 0.5 mole
so the volume occupied = $0.5 \times 24 \text{ dm}^3 = 12 \text{ dm}^3$.

### Calculating gas volumes from equations

From the equation for a reaction, you can tell how many moles of a gas take part. Then you can use Avogadro’s Law to work out its volume.

In these examples, all volumes are measured at rtp.

**Example 1** What volume of hydrogen will react with 24 dm$^3$ of oxygen to form water?
1. The equation for the reaction is: $2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O} (l)$
2. So 2 volumes of hydrogen react with 1 of oxygen, or
   $2 \times 24 \text{ dm}^3$ react with 24 dm$^3$, so
   48 dm$^3$ of hydrogen will react.

**Example 2** When sulfur burns in air it forms sulfur dioxide. What volume of this gas is produced when 1 g of sulfur burns? ($A_r$: S = 32.)
1. The equation for the reaction is: $\text{S} (s) + \text{O}_2 (g) \rightarrow \text{SO}_2 (g)$
2. 32 g of sulfur atoms = 1 mole of sulfur atoms, so
   $1 \text{ g} = \frac{1}{32}$ mole or 0.03125 moles of sulfur atoms.
3. 1 mole of sulfur atoms gives 1 mole of sulfur dioxide molecules so
   0.03125 moles give 0.03125 moles.
4. 1 mole of sulfur dioxide molecules has a volume of 24 dm$^3$ at rtp so
   0.03125 moles has a volume of $0.03125 \times 24 \text{ dm}^3$ at rtp, or 0.75 dm$^3$. 
   So 0.75 dm$^3$ (or 750 cm$^3$) of sulfur dioxide are produced.

---

**Q**

(A: O = 16, N = 14, H = 1, C = 12.)

1. What does *rtp* mean? What values does it have?
2. What does *molar volume* mean, for a gas?
3. What is the molar volume of neon gas at rtp?
4. For any gas, calculate the volume at rtp of:
   a. 7 moles  
   b. 0.5 moles  
   c. 0.001 moles
5. Calculate the volume at rtp of:
   a. 16 g of oxygen ($\text{O}_2$)  
   b. 1.7 g of ammonia (NH$_3$)
6. You burn 6 grams of carbon in plenty of air:
   $\text{C} (s) + \text{O}_2 (g) \rightarrow \text{CO}_2 (g)$
   a. What volume of gas will form (at rtp)?
   b. What volume of oxygen will be used up?
7. If you burn the carbon in limited air, the reaction is different: $2 \text{C} (s) + \text{O}_2 (g) \rightarrow 2 \text{CO} (g)$
   a. What volume of gas will form this time?
   b. What volume of oxygen will be used up?
The concentration of a solution is the amount of solute, in grams or moles, that is dissolved in 1 dm³ of solution.

Finding the concentration in moles

Example  Find the concentrations of A and C above, in moles per dm³.

First, change the mass of the solute to moles.

The formula mass of copper(II) sulfate is 250, as shown on the right. So 1 mole of the compound has a mass of 250 g.

Solution A has 2.5 g of the compound in 1 dm³ of solution.

2.5 g moles 0.01 moles
so its concentration is 0.01 mol/dm³.

Note the unit of concentration: mol/dm³. This is often shortened to M, so the concentration of solution A can be written as 0.01 M.

Solution C has 250 g of the compound in 1 dm³ of solution.

250 g 1 mole
so its concentration is 1 mol/dm³, or 1 M for short.

A solution that contains 1 mole of solute per dm³ of solution is often called a molar solution. So C is a molar solution.

In general, to find the concentration of a solution in moles per dm³:

\[ \text{concentration (mol/dm³)} = \frac{\text{amount of solute (mol)}}{\text{volume of solution (dm³)}} \]

Use the equation above to check that the last column in this table is correct:

<table>
<thead>
<tr>
<th>Amount of solute (mol)</th>
<th>Volume of solution (dm³)</th>
<th>Concentration of solution (mol/dm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0</td>
<td>1.0</td>
<td>1.0</td>
</tr>
<tr>
<td>0.2</td>
<td>0.1</td>
<td>2.0</td>
</tr>
<tr>
<td>0.5</td>
<td>0.2</td>
<td>2.5</td>
</tr>
<tr>
<td>1.5</td>
<td>0.3</td>
<td>5.0</td>
</tr>
</tbody>
</table>

Remember

- 1 dm³ = 1 litre = 1000 cm³ = 1000 ml
- All these mean the same thing: moles per dm³, mol/dm³, mol dm⁻³, moles per litre
Finding the amount of solute in a solution
If you know the concentration of a solution, and its volume:
- you can work out how much solute it contains, in moles.
  Just rearrange the equation from the last page:
  \[ \text{amount of solute (mol)} = \text{concentration (mol/dm³)} \times \text{volume (dm³)} \]
- you can then convert moles to grams, by multiplying the number of moles by \( M_r \).

Sample calculations
The table shows four solutions, with different volumes and concentrations. Check that you understand the calculations that give the masses of solute in the bottom row.

<table>
<thead>
<tr>
<th>solution</th>
<th>sodium hydroxide NaOH</th>
<th>sodium thiosulfate Na₂S₂O₃</th>
<th>lead nitrate Pb(NO₃)₂</th>
<th>silver nitrate AgNO₃</th>
</tr>
</thead>
<tbody>
<tr>
<td>concentration (mol/dm³)</td>
<td>1</td>
<td>2</td>
<td>0.1</td>
<td>0.05</td>
</tr>
<tr>
<td>amount of solute present (moles)</td>
<td>1 × 2 = 2</td>
<td>2 × ( \frac{250}{1000} ) = 0.5</td>
<td>0.1 × ( \frac{100}{1000} ) = 0.01</td>
<td>0.05 × ( \frac{25}{1000} ) = 0.00125</td>
</tr>
<tr>
<td>( M_r )</td>
<td>40</td>
<td>158</td>
<td>331</td>
<td>170</td>
</tr>
<tr>
<td>mass of solute present (g)</td>
<td>80</td>
<td>79</td>
<td>3.31</td>
<td>0.2125</td>
</tr>
</tbody>
</table>

Use the calculation triangle

\[ \text{conc (mol/dm³) \times vol (dm³)} \]

▲ Cover the one you want to find—and you will see how to calculate it.
To draw this triangle, remember that alligators chew visitors!

| Q |  
|---|---
| **1** | How many moles of solute are in:  
| a | 500 cm³ of solution, of concentration 2 mol/dm³?  
| b | 2 litres of solution, of concentration 0.5 mol/dm³?  
| **2** | What is the concentration of a solution containing:  
| a | 4 moles in 2 dm³ of solution?  
| b | 0.3 moles in 200 cm³ of solution?  
| **3** | Different solutions of salt \( \chi \) are made up. What volume of:  
| a | a 4 mol/dm³ solution contains 2 moles of \( \chi \)?  
| b | a 6 mol/dm³ solution contains 0.03 moles of \( \chi \)?  
| **4** | The \( M_r \) of sodium hydroxide is 40. How many grams of sodium hydroxide are there in:  
| a | 500 cm³ of a molar solution?  
| b | 25 cm³ of a 0.5 M solution?  
| **5** | What is the concentration in moles per litre of:  
| a | a sodium carbonate solution containing 53 g of the salt (Na₂CO₃) in 1 litre?  
| b | a copper(II) sulfate solution containing 62.5 g of the salt (CuSO₄·5H₂O) in 1 litre?  

---

83
6.5 Finding the empirical formula

What a formula tells you about moles and masses
The formula of carbon dioxide is $\text{CO}_2$. Some molecules of it are shown on the right. You can see that:

- 1 carbon atom combines with 2 oxygen atoms
- 1 mole of carbon atoms combines with 2 moles of oxygen atoms

Moles can be changed to grams, using $A_r$ and $M_r$. So we can write:

- 12 g of carbon combines with 32 g of oxygen

In the same way:
- 6 g of carbon combines with 16 g of oxygen
- 24 kg of carbon combines with 64 kg of oxygen, and so on.

The masses of substances that combine are always in the same ratio.

Therefore, from the formula of a compound, you can tell:
- how many moles of the different atoms combine
- how many grams of the different elements combine.

Finding the empirical formula
From the formula of a compound you can tell what masses of the elements combine. But you can also do things the other way round.

If you know what masses combine, you can work out the formula.
These are the steps:

1. Find the masses that combine (in grams) by experiment.
2. Change grams to moles of atoms.
3. This tells you the ratio in which atoms combine.
4. So you can write a formula.

A formula found in this way is called the empirical formula.
The empirical formula shows the simplest ratio in which atoms combine.

Example 1 32 grams of sulfur combine with 32 grams of oxygen to form an oxide of sulfur. What is its empirical formula?

Draw up a table like this:

<table>
<thead>
<tr>
<th>Elements that combine</th>
<th>sulfur</th>
<th>oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses that combine</td>
<td>32 g</td>
<td>32 g</td>
</tr>
<tr>
<td>Relative atomic masses ($A_r$)</td>
<td>32</td>
<td>16</td>
</tr>
<tr>
<td>Moles of atoms that combine</td>
<td>$32/32$</td>
<td>$32/16$</td>
</tr>
<tr>
<td>Ratio in which atoms combine</td>
<td>1:2</td>
<td></td>
</tr>
<tr>
<td>Empirical formula</td>
<td>$\text{SO}_2$</td>
<td></td>
</tr>
</tbody>
</table>

So the empirical formula of the oxide that forms is $\text{SO}_2$. 

△ Sulfur combines with oxygen when it burns.
Example 2: An experiment shows that compound Y is 80% carbon and 20% hydrogen. What is its empirical formula?

Y is 80% carbon and 20% hydrogen. So 100 g of Y contains 80 g of carbon and 20 g of hydrogen. Draw up a table like this:

<table>
<thead>
<tr>
<th>Elements that combine</th>
<th>carbon</th>
<th>hydrogen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses that combine</td>
<td>80 g</td>
<td>20 g</td>
</tr>
<tr>
<td>Relative atomic masses ($A_i$)</td>
<td>12</td>
<td>1</td>
</tr>
<tr>
<td>Moles of atoms that combine</td>
<td>80/12</td>
<td>20/1</td>
</tr>
<tr>
<td>Ratio in which atoms combine</td>
<td>6.67 : 20 or 1:3 in its simplest form</td>
<td></td>
</tr>
<tr>
<td><strong>Empirical formula</strong></td>
<td>$\text{CH}_3$</td>
<td></td>
</tr>
</tbody>
</table>

So the empirical formula of Y is $\text{CH}_3$.

But we can tell right away that the molecular formula for Y must be different. (A carbon atom does not bond to only 3 hydrogen atoms.) You will learn how to find the molecular formula from the empirical formula in the next unit.

An experiment to find the empirical formula

To work out the empirical formula, you need to know the masses of elements that combine. The only way to do this is by experiment.

For example, magnesium combines with oxygen to form magnesium oxide. The masses that combine can be found like this:

1. Weigh a crucible and lid, empty. Then add a coil of magnesium ribbon and weigh it again, to find the mass of the magnesium.
2. Heat the crucible. Raise the lid carefully at intervals to let oxygen in. The magnesium burns brightly.
3. When burning is complete, let the crucible cool (still with its lid on). Then weigh it again. The increase in mass is due to oxygen.

The results showed that 2.4 g of magnesium combined with 1.6 g of oxygen. Draw up a table again:

<table>
<thead>
<tr>
<th>Elements that combine</th>
<th>magnesium</th>
<th>oxygen</th>
</tr>
</thead>
<tbody>
<tr>
<td>Masses that combine</td>
<td>2.4 g</td>
<td>1.6 g</td>
</tr>
<tr>
<td>Relative atomic masses ($A_i$)</td>
<td>24</td>
<td>16</td>
</tr>
<tr>
<td>Moles of atoms that combine</td>
<td>2.4/24</td>
<td>1.6/16</td>
</tr>
<tr>
<td>Ratio in which atoms combine</td>
<td>1:1</td>
<td></td>
</tr>
<tr>
<td><strong>Empirical formula</strong></td>
<td>$\text{MgO}$</td>
<td></td>
</tr>
</tbody>
</table>

So the empirical formula for the oxide is $\text{MgO}$.

1. **a** How many atoms of hydrogen combine with one carbon atom to form methane, $\text{CH}_4$?
   **b** How many grams of hydrogen combine with 12 grams of carbon to form methane?

2. What does the word *empirical* mean? (Check the glossary?)

3. 56 g of iron combine with 32 g of sulfur to form iron sulfide. Find the empirical formula for iron sulfide. ($A_i$: Fe = 56, S = 32.)

4. An oxide of sulfur is 40% sulfur and 60% oxygen. What is its empirical formula?
The formula of an ionic compound
You saw in the last unit that the empirical formula shows the simplest ratio in which atoms combine.

The diagram on the right shows the structure of sodium chloride. The sodium and chlorine atoms are in the ratio 1:1 in this compound. So its empirical formula is NaCl.

The formula of an ionic compound is the same as its empirical formula.
In the experiment on page 85, the empirical formula for magnesium oxide was found to be MgO. So the formula for magnesium oxide is also MgO.

The formula of a molecular compound
The gas ethane is one of the alkane family of compounds. An ethane molecule is drawn on the right. It contains only hydrogen and carbon atoms, so ethane is a hydrocarbon.

From the drawing you can see that the ratio of carbon to hydrogen atoms in ethane is 2:6. The simplest ratio is therefore 1:3.
So the empirical formula of ethane is CH₃. (It is compound Y on page 85.) But its molecular formula is C₂H₆.

The molecular formula shows the actual numbers of atoms that combine to form a molecule.
The molecular formula is more useful than the empirical formula, because it gives you more information.

For some molecular compounds, both formulae are the same. For others they are different. Compare them for the alkanes in the table on the right. What do you notice?

How to find the molecular formula
To find the molecular formula for an unknown compound, you need to know these:
- the relative molecular mass of the compound (Mᵣ). This can be found using a mass spectrometer.
- its empirical formula. This is found by experiment, as on page 85.
- its empirical mass. This is the mass calculated using the empirical formula and Aᵣ values.

Once you know those, you can work out the molecular formula by following these steps:

To find the molecular formula:

i Calculate \( \frac{Mᵣ}{\text{empirical mass}} \) for the compound. This gives a number, \( n \).

ii Multiply the numbers in the empirical formula by \( n \).

Let’s look at two examples.
Calculating the molecular formula

Example 1 A molecular compound has the empirical formula HO. Its relative molecular mass is 34. What is its molecular formula? (A.; H = 1, O = 16.)

For the empirical formula HO, the empirical mass = 17. But \( M_r = 34 \).

So \( \frac{M_r}{\text{empirical mass}} = \frac{34}{17} = 2 \)

So the molecular formula is \( 2 \times \text{HO} \), or \( \text{H}_2\text{O}_2 \).

So the compound is hydrogen peroxide.

Note how you write the 2 after the symbols, when you multiply.

Example 2 Octane is a hydrocarbon – it contains only carbon and hydrogen. It is 84.2% carbon and 15.8% hydrogen by mass. Its \( M_r \) is 114. What is its molecular formula?

1 First find the empirical formula for the compound.

From the %, we can say that in 100 g of octane, 84.2 g is carbon and 15.8 g is hydrogen.

So 84.2 g of carbon combines with 15.8 g of hydrogen.

Changing masses to moles:

\[
\frac{84.2}{12} \text{ moles of carbon atoms combine with } \frac{15.8}{1} \text{ moles of hydrogen atoms, or }
\]

7.02 moles of carbon atoms combine with 15.8 moles of hydrogen atoms, so

1 mole of carbon atoms combines with \( \frac{15.8}{7.02} \) or 2.25 moles of hydrogen atoms.

So the atoms combine in the ratio of 1: 2.25 or 4:9.

(Give the ratio as whole numbers, since only whole atoms combine.)

The empirical formula of octane is therefore \( \text{C}_4\text{H}_9 \).

2 Then use \( M_r \) to find the molecular formula.

For the empirical formula (\( \text{C}_4\text{H}_9 \)), the empirical mass = 57.

But \( M_r = 114 \).

So \( \frac{M_r}{\text{empirical mass}} = \frac{114}{57} = 2 \)

So the molecular formula of octane is \( 2 \times \text{C}_4\text{H}_9 \), or \( \text{C}_8\text{H}_{18} \).

1 In the ionic compound magnesium chloride, magnesium and chlorine atoms combine in the ratio 1:2. What is the formula of magnesium chloride?

2 In the ionic compound aluminium fluoride, aluminium and fluorine atoms combine in the ratio 1:3. What is the formula of aluminium fluoride?

3 What is the difference between an empirical formula and a molecular formula? Can they ever be the same?

4 What is the empirical formula of benzene, \( \text{C}_6\text{H}_6 \)?

5 A compound has the empirical formula \( \text{CH}_2 \). Its \( M_r \) is 28. What is its molecular formula?

6 A hydrocarbon is 84% carbon, by mass. Its relative molecular mass is 100. Find:

   a its empirical formula  
   b its molecular formula

7 An oxide of phosphorus has an \( M_r \) value of 220. It is 56.4% phosphorus. Find its molecular formula.
Yield and purity
The **yield** is the amount of product you obtain from a reaction. Suppose you own a factory that makes paint or fertilisers. You will want the highest yield possible, for the lowest cost!

Now imagine your factory makes medical drugs, or flavouring for foods. The yield will still be important – but the **purity** of the product may be even more important. Impurities could harm people.

In this unit you’ll learn how to calculate the % yield from a reaction, and remind yourself how to calculate the % purity of the product obtained.

**Finding the % yield**
You can work out % yield like this:

\[ \% \text{ yield} = \frac{\text{actual mass obtained}}{\text{calculated mass}} \times 100\% \]

**Example** The medical drug aspirin is made from salicylic acid. 1 mole of salicylic acid gives 1 mole of aspirin:

\[
\text{C}_7\text{H}_6\text{O}_3 \xrightarrow{\text{chemicals}} \text{C}_9\text{H}_8\text{O}_4
\]

salicylic acid \quad \text{aspirin}

In a trial, 100.0 grams of salicylic acid gave 121.2 grams of aspirin. What was the % yield?
1. \(A_c: C = 12, H = 1, O = 16.\)
   So \(M_r: \text{salicylic acid} = 138, \text{aspirin} = 180.\)
2. 138 g of salicylic acid = 1 mole
   so \(100 \text{ g} = \frac{100}{138} \text{ mole} = 0.725 \text{ moles}\)
3. 1 mole of salicylic acid gives 1 mole of aspirin
   so 0.725 moles give 0.725 moles of aspirin
   or \(0.725 \times 180 \text{ g} = 130.5 \text{ g}\)

   So 130.5 g is the **calculated mass** for the reaction.
4. But the **actual mass** obtained in the trial was 121.2 g.
   So \(\% \text{ yield} = \frac{121.2 \text{ g}}{130.5 \text{ g}} \times 100 = 92.9\% \)

This is a high yield – so it is worth continuing with those trials.

**Finding the % purity**
When you make something in a chemical reaction, and separate it from the final mixture, it will not be pure. It will have impurities mixed with it – for example small amounts of unreacted substances, or another product.

You can work out the % purity of the product you obtained like this:

\[ \% \text{ purity of a product} = \frac{\text{mass of the pure product}}{\text{mass of the impure product obtained}} \times 100\% \]
Below are examples of how to work out the % purity.

**Example 1**  Aspirin is itself an acid. (Its full name is acetylsalicylic acid.) It is neutralised by sodium hydroxide in this reaction:

$$\text{C}_9\text{H}_8\text{O}_4 (aq) + \text{NaOH} (aq) \rightarrow \text{C}_9\text{H}_7\text{O}_4\text{Na} (aq) + \text{H}_2\text{O} (l)$$

Some aspirin was prepared in the lab. Through titration, it was found that 4.00 g of the aspirin were neutralised by 17.5 cm$^3$ of 1M sodium hydroxide solution. How pure was the aspirin sample?

1. $M_r$ of C$_9$H$_8$O$_4$ = 180  \( (A_r: C = 12, H = 1, O = 16) \)
2. 17.5 cm$^3$ of 1M sodium hydroxide contain 17.5 $\times$ 1000 moles or 0.0175 moles of NaOH
3. 1 mole of NaOH reacts with 1 mole of C$_9$H$_8$O$_4$ so 0.0175 moles react with 0.0175 moles.
4. 0.0175 moles of C$_9$H$_8$O$_4$ = 0.0175 $\times$ 180 g or 3.15 g of aspirin.
5. But the mass of the aspirin sample was 4 g.

So % purity of the aspirin = $\frac{3.15}{4} \times 100\%$ or 78.75%.

This is far from acceptable for medical use. The aspirin could be purified by crystallisation. Repeated crystallisation might be needed.

**Example 2**  Chalk is almost pure calcium carbonate.

10 g of chalk was reacted with an excess of dilute hydrochloric acid. 2280 cm$^3$ of carbon dioxide gas was collected at room temperature and pressure (rtp). What was the purity of the sample?

You can work out its purity from the volume of carbon dioxide given off. The equation for the reaction is:

$$\text{CaCO}_3 (s) + 2\text{HCl} (aq) \rightarrow \text{CaCl}_2 (aq) + \text{H}_2\text{O} (l) + \text{CO}_2 (g)$$

1. $M_r$ of CaCO$_3$ = 100  \( (A_r: Ca = 40, C = 12, O = 16) \)
2. 1 mole of CaCO$_3$ gives 1 mole of CO$_2$ and
   1 mole of gas has a volume of 24 000 cm$^3$ at rtp.
3. So 24 000 cm$^3$ of gas is produced by 100 g of calcium carbonate and
   2280 cm$^3$ is produced by $\frac{2280}{24000} \times 100$ g or 9.5 g.

So there is 9.5 g of calcium carbonate in the 10 g of chalk.

So the % purity of the chalk = $\frac{9.5}{10} \times 100 = 95\%$.

**Q**

1. Define the term:  a % yield  b % purity
2. 100 g of aspirin was obtained from 100 g of salicylic acid. What was the % yield?
3. 17 kg of aluminium was produced from 51 kg of aluminium oxide (Al$_2$O$_3$) by electrolysis. What was the percentage yield?  \( (A_r: Al = 27, O = 16) \)
4. Some seawater is evaporated. The sea salt obtained is found to be 86% sodium chloride. How much sodium chloride could be obtained from 200 g of this salt?
5. A 5.0 g sample of dry ice (solid carbon dioxide) turned into 2400 cm$^3$ of carbon dioxide gas at rtp. What was the percentage purity of the dry ice?  \( (M_r \text{ of CO}_2 = 44) \)
Questions

Extended curriculum

1. Iron is obtained by reducing iron(III) oxide using the gas carbon monoxide. The reaction is:
   \[ \text{Fe}_2\text{O}_3 (s) + 3\text{CO (g)} \rightarrow 2\text{Fe (s)} + 3\text{CO}_2 (g) \]
   a. Write a word equation for the reaction.
   b. What is the formula mass of iron(III) oxide? (\(A_r: \text{Fe} = 56, \text{O} = 16\).)
   c. How many moles of \(\text{Fe}_2\text{O}_3\) are there in 320 kg of iron(III) oxide? (1 kg = 1000 g.)
   d. How many moles of \(\text{Fe}\) are obtained from 1 mole of \(\text{Fe}_2\text{O}_3\)?
   e. From c and d, find how many moles of iron atoms are obtained from 320 kg of iron(III) oxide.
   f. How much iron (in kg) is obtained from 320 kg of iron(III) oxide?

2. With strong heating, calcium carbonate undergoes thermal decomposition:
   \[ \text{CaCO}_3 (s) \rightarrow \text{CaO (s)} + \text{CO}_2 (g) \]
   a. Write a word equation for the change.
   b. How many moles of \(\text{CaCO}_3\) are in 50 g of calcium carbonate? (\(A_r: \text{Ca} = 40, \text{C} = 12, \text{O} = 16\).)
   c. i. What mass of calcium oxide is obtained from the thermal decomposition of 50 g of calcium carbonate, assuming a 40% yield?
      ii. What mass of carbon dioxide will be given off at the same time?
      iii. What volume will this gas occupy at rtp?

3. Nitroglycerine is used as an explosive.
   The equation for the explosion reaction is:
   \[ 4\text{C}_3\text{H}_5(\text{NO}_3)_3 (l) \rightarrow 12\text{CO}_2 (g) + 10\text{H}_2\text{O (l)} + 6\text{N}_2 (g) + \text{O}_2 (g) \]
   a. How many moles does the equation show for:
      i. nitroglycerine?
      ii. gas molecules produced?
   b. How many moles of gas molecules are obtained from 1 mole of nitroglycerine?
   c. What is the total volume of gas (at rtp) obtained from 1 mole of nitroglycerine?
   d. What is the mass of 1 mole of nitroglycerine? (\(A_r: \text{H} = 1, \text{C} = 12, \text{N} = 14, \text{O} = 16\).)
   e. What will be the total volume of gas (at rtp) from exploding 1 kg of nitroglycerine?
   f. Using your answers above, try to explain why nitroglycerine is used as an explosive.
4 Nitrogen monoxide reacts with oxygen like this: 
\[ 2\text{NO} (g) + \text{O}_2 (g) \rightarrow 2\text{NO}_2 (g) \]
- a. How many moles of oxygen molecules react with 1 mole of nitrogen monoxide molecules?
- b. What volume of oxygen will react with 50 cm\(^3\) of nitrogen monoxide?
- c. Using the volumes in b, what is:
  i. the total volume of the two reactants?
  ii. the volume of nitrogen dioxide formed?

5 2 g (an excess) of iron is added to 50 cm\(^3\) of 0.5 M sulfuric acid. When the reaction is over, the reaction mixture is filtered. The mass of the unreacted iron is found to be 0.6 g. (\(A_r\): Fe = 56.)
- a. What mass of iron took part in the reaction?
- b. How many moles of iron atoms took part?
- c. How many moles of sulfuric acid reacted?
- d. Write the equation for the reaction, and deduce the charge on the iron ion that formed.
- e. What volume of hydrogen (calculated at rtp) bubbled off during the reaction?

6 27 g of aluminium burns in chlorine to form 133.5 g of aluminium chloride. (\(A_r\): Al = 27, Cl = 35.5.)
- a. What mass of chlorine is present in 133.5 g of aluminium chloride?
- b. How many moles of chlorine atoms is this?
- c. How many moles of aluminium atoms are present in 27 g of aluminium?
- d. Use your answers for parts b and c to find the simplest formula of aluminium chloride.
- e. 1 dm\(^3\) of an aqueous solution is made using 13.35 g of aluminium chloride. What is its concentration in moles per dm\(^3\)?

7 You have to prepare some 2 M solutions, with 10 g of solute in each. What volume of solution will you prepare, for each solute below? (\(A_r\): H = 1, Li = 7, N = 14, O = 16, Mg = 24, S = 32.)
- a. lithium sulfate, Li\(_2\)SO\(_4\)
- b. magnesium sulfate, MgSO\(_4\)
- c. ammonium nitrate, NH\(_4\)NO\(_3\)

8 Phosphorus forms two oxides, which have the empirical formulae P\(_2\)O\(_3\) and P\(_2\)O\(_5\).
- a. Which oxide contains the higher percentage of phosphorus? (\(A_r\): P = 31, O = 16.)
- b. What mass of phosphorus will combine with 1 mole of oxygen molecules (O\(_2\)) to form P\(_2\)O\(_3\)?
- c. What is the molecular formula of the oxide that has a formula mass of 284?
- d. Suggest a molecular formula for the other oxide.

9 Zinc and phosphorus react to give zinc phosphide. 9.75 g of zinc combines with 3.1 g of phosphorus.
- a. Find the empirical formula for the compound. (\(A_r\): Zn = 65, P = 31.)
- b. Calculate the percentage of phosphorus in it.

10 110 g of manganese was extracted from 174 g of manganese oxide. (\(A_r\): Mn = 55, O = 16.)
- a. What mass of oxygen is there in 174 g of manganese oxide?
- b. How many moles of oxygen atoms is this?
- c. How many moles of manganese atoms are there in 110 g of manganese?
- d. Give the empirical formula of manganese oxide.
- e. What mass of manganese can obtained from 1000 g of manganese oxide?

11 Find the molecular formulae for these compounds. (\(A_r\): H = 1, C = 12, N = 14, O = 16.)

<table>
<thead>
<tr>
<th>Compound</th>
<th>(M_r)</th>
<th>Empirical formula</th>
<th>Molecular formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>a hydrazine</td>
<td>32</td>
<td>NH(_2)</td>
<td></td>
</tr>
<tr>
<td>b cyanogen</td>
<td>52</td>
<td>CN</td>
<td></td>
</tr>
<tr>
<td>c nitrogen oxide</td>
<td>92</td>
<td>NO(_2)</td>
<td></td>
</tr>
<tr>
<td>d glucose</td>
<td>180</td>
<td>CH(_2)O</td>
<td></td>
</tr>
</tbody>
</table>

12 Hydrocarbons A and B both contain 85.7% carbon. Their molar masses are 42 and 84 g respectively.
- a. Which elements does a hydrocarbon contain?
- b. Calculate the empirical formulae of A and B.
- c. Calculate the molecular formulae of A and B.

13 Mercury(II) oxide breaks down on heating: 
\[ 2\text{HgO} (s) \rightarrow 2\text{Hg} (l) + \text{O}_2 (g) \]
- a. Calculate the mass of 1 mole of mercury(II) oxide. (\(A_r\): O = 16, Hg = 201)
- b. How much mercury and oxygen could be obtained from 21.7 g of mercury(II) oxide?
- c. Only 19.0 g of mercury was collected. Calculate the % yield of mercury for this experiment.

14 A 5-g sample of impure magnesium carbonate is reacted with an excess of hydrochloric acid: 
\[ \text{MgCO}_3 (s) + 2\text{HCl} (aq) \rightarrow \text{MgCl}_2 (aq) + \text{H}_2\text{O} (l) + \text{CO}_2 (g) \]
- a. How many moles of CO\(_2\) are produced?
- b. What mass of pure magnesium carbonate would give this volume of carbon dioxide? (\(A_r\): C = 12, O = 16, Mg = 24.)
- c. Calculate the % purity of the 5-g sample.
Different groups of reactions
Thousands of different reactions go on around us, in labs, and factories, and homes. We can divide them into different groups. For example two of the groups are neutralisation reactions and precipitation reactions.
One big group is the redox reactions, in which oxidation and reduction occur. We focus on those in this chapter.

Oxidation: oxygen is gained
Magnesium burns in air with a dazzling white flame. A white ash is formed. The reaction is:

\[
2\text{Mg (s)} + \text{O}_2 (g) \rightarrow 2\text{MgO (s)}
\]

The magnesium has gained oxygen. We say it has been oxidised.

A gain of oxygen is called oxidation. The substance has been oxidised.

Reduction: oxygen is lost
Now look what happens when hydrogen is passed over heated copper(II) oxide. The black compound turns pink:

\[
\text{copper(II) oxide + hydrogen } \rightarrow \text{ copper + water}
\]

\[
\text{CuO (s) + H}_2 (g) \rightarrow \text{Cu (s) + H}_2\text{O (l)}
\]

This time the heated substance is losing oxygen. It is being reduced.

A loss of oxygen is called reduction. The substance is reduced.

Iron occurs naturally in the earth as iron(III) oxide, Fe$_2$O$_3$.
This is reduced to iron in the blast furnace. Here, molten iron runs out from the bottom of the furnace.
And here, iron is being oxidised to iron(III) oxide again! We call this process rusting. It is ruining the bikes. The formula for rust is Fe$_2$O$_3$.2H$_2$O.
**Oxidation and reduction take place together**

Look again at the reaction between copper(II) oxide and hydrogen. Copper(II) oxide loses oxygen, and hydrogen gains oxygen:

$$\text{CuO (s)} + \text{H}_2(g) \rightarrow \text{Cu (s)} + \text{H}_2\text{O (l)}$$

So the copper(II) oxide is reduced, and the hydrogen is oxidised.

**Oxidation and reduction always take place together. So the reaction is called a redox reaction.**

**Two more examples of redox reactions**

**The reaction between calcium and oxygen** Calcium burns in air with a red flame, to form the white compound calcium oxide. It is easy to see that calcium has been oxidised. But oxidation and reduction always take place together, which means oxygen has been reduced:

$$2\text{Ca (s)} + \text{O}_2(g) \rightarrow 2\text{CaO (s)}$$

**The reaction between hydrogen and oxygen** Hydrogen reacts explosively with oxygen, to form water. Hydrogen is oxidised, and oxygen is reduced:

$$2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O (l)}$$

**Those burning reactions**
- Another name for burning is combustion.
- Combustion is a redox reaction.
- For example, when an element burns in oxygen, it is oxidised to its oxide.

**Q**

1. Copy and complete the statements:
   - a. Oxidation means ...
   - b. Reduction means ...
   - c. Oxidation and reduction always ...

2. Magnesium reacts with sulfur dioxide like this:
   $$2\text{Mg (s)} + \text{SO}_2(g) \rightarrow 2\text{MgO (s)} + \text{S (s)}$$
   Copy the equation, and use labelled arrows to show which substance is oxidised, and which is reduced.

3. Explain where the term redox comes from.

4. Many people cook with natural gas, which is mainly methane, CH₄. The equation for its combustion is:
   $$\text{CH}_4 (g) + 2\text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2\text{H}_2\text{O (l)}$$
   Show that this is a redox reaction.

5. Write down the equation for the reaction between magnesium and oxygen. Use labelled arrows to show which element is oxidised, and which is reduced.
Another definition for oxidation and reduction
When magnesium burns in oxygen, magnesium oxide is formed:

\[ 2\text{Mg (s)} + \text{O}_2 (g) \rightarrow 2\text{MgO (s)} \]

The magnesium has clearly been oxidised. Oxidation and reduction always take place together, so the oxygen must have been reduced. But how?
Let's see what is happening to the electrons:

During the reaction, each magnesium atom loses two electrons and each oxygen atom gains two. This leads us to a new definition:

**If a substance loses electrons during a reaction, it has been oxidised.**

**If it gains electrons, it has been reduced.**

The reaction is a redox reaction.

Writing half-equations to show the electron transfer
You can use **half-equations** to show the electron transfer in a reaction. One half-equation shows electron loss, and the other shows electron gain.

This is how to write the half-equations for the reaction above:

1. **Write down each reactant, with the electrons it gains or loses.**
   - magnesium: \( \text{Mg} \rightarrow \text{Mg}^{2+} + 2e^- \)
   - oxygen: \( \text{O} + 2e^- \rightarrow \text{O}^{2-} \)

2. **Check that each substance is in its correct form (ion, atom or molecule) on each side of the arrow. If it is not, correct it.**
   Oxygen is not in its correct form on the left above. It exists as molecules, so you must change \( \text{O} \rightarrow \text{O}_2 \). That means you must also double the number of electrons and oxide ions:
   - oxygen: \( \text{O}_2 + 4e^- \rightarrow 2\text{O}^{2-} \)

3. **The number of electrons must be the same in both equations. If it is not, multiply one (or both) equations by a number, to balance them.**
   - So we must multiply the magnesium half-equation by 2.
   - magnesium: \( 2\text{Mg} \rightarrow 2\text{Mg}^{2+} + 4e^- \)
   - oxygen: \( \text{O}_2 + 4e^- \rightarrow 2\text{O}^{2-} \)
   The equations are now balanced, each with 4 electrons.
Redox without oxygen

Our definition of redox reactions is now much broader:

**Any reaction in which electron transfer takes place is a redox reaction.**

So the reaction does not have to include oxygen! Look at these examples:

1. **The reaction between sodium and chlorine**
   The equation is:
   \[ 2Na (s) + Cl_2 (g) \rightarrow 2NaCl (s) \]
   The sodium atoms give electrons to the chlorine atoms, forming ions as shown on the right. So sodium is oxidised, and chlorine is reduced.

   So the reaction is a redox reaction. Look at the half-equations:
   - sodium: \( 2Na \rightarrow 2Na^+ + 2e^- \) (oxidation)
   - chlorine: \( Cl_2 + 2e^- \rightarrow 2Cl^- \) (reduction)

2. **The reaction between chlorine and potassium bromide**
   When chlorine gas is bubbled through a colourless solution of potassium bromide, the solution goes orange due to this reaction:
   \[ Cl_2 (g) + 2KBr (aq) \rightarrow 2KCl (aq) + Br_2 (aq) \]
   colourless orange

   Bromine has been displaced. The half-equations for the reaction are:
   - chlorine: \( Cl_2 + 2e^- \rightarrow 2Cl^- \) (reduction)
   - bromide ion: \( 2Br^- \rightarrow Br_2 + 2e^- \) (oxidation)

**From half-equations to the ionic equation**

Adding the balanced half-equations gives the ionic equation for the reaction.

**An ionic equation shows the ions that take part in the reaction.**

For example, for the reaction between chlorine and potassium bromide:

\[
\begin{align*}
Cl_2 + 2e^- & \rightarrow 2Cl^- \\
2Br^- & \rightarrow Br_2 + 2e^- \\
\hline
Cl_2 + 2e^- + 2Br^- & \rightarrow 2Cl^- + Br_2 + 2e^- 
\end{align*}
\]

The electrons cancel, giving the ionic equation for the reaction:

\[ Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2 \]

**Redox: a summary**

**Oxidation** is gain of oxygen, or loss of electrons.

**Reduction** is loss of oxygen, or gain of electrons.

Oxidation and reduction always take place together, in a **redox reaction**.

---

**Q**

1. Give a full definition for: **a** oxidation  **b** reduction

2. What does a half-equation show?

3. Potassium and chlorine react to form potassium chloride.
   **a** It is a redox reaction. Explain why.
   **b** See if you can write the balanced half-equations for it.

   **a** Write the balanced half-equations for this reaction.
   **b** Add the half-equations, to give the ionic equation for the reaction.
Redox and changes in oxidation state

What does oxidation state mean?

Oxidation state tells you how many electrons each atom of an element has gained, lost, or shared, in forming a compound.

As you will see, oxidation states can help you to identify redox reactions.

The rules for oxidation states

1. Each atom in a formula has an oxidation state.
2. The oxidation state is usually given as a Roman numeral.

<table>
<thead>
<tr>
<th>Number</th>
<th>Roman numeral</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>1</td>
<td>I</td>
</tr>
<tr>
<td>2</td>
<td>II</td>
</tr>
<tr>
<td>3</td>
<td>III</td>
</tr>
<tr>
<td>4</td>
<td>IV</td>
</tr>
<tr>
<td>5</td>
<td>V</td>
</tr>
<tr>
<td>6</td>
<td>VI</td>
</tr>
<tr>
<td>7</td>
<td>VII</td>
</tr>
</tbody>
</table>
3. Where an element is not combined with other elements, its atoms are in oxidation state 0.
4. Many elements have the same oxidation state in most or all their compounds. Look at these:

<table>
<thead>
<tr>
<th>Element</th>
<th>Usual oxidation state in compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen</td>
<td>+1</td>
</tr>
<tr>
<td>sodium and the other Group I metals</td>
<td>+1</td>
</tr>
<tr>
<td>calcium and the other Group II metals</td>
<td>+2</td>
</tr>
<tr>
<td>aluminium</td>
<td>+3</td>
</tr>
<tr>
<td>chlorine and the other Group VII non-metals, in compounds without oxygen</td>
<td>-1</td>
</tr>
<tr>
<td>oxygen (except in peroxides)</td>
<td>-2</td>
</tr>
</tbody>
</table>

5. But atoms of transition elements can have variable oxidation states in their compounds. Look at these:

<table>
<thead>
<tr>
<th>Element</th>
<th>Common oxidation states in compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>iron</td>
<td>+II and +III</td>
</tr>
<tr>
<td>copper</td>
<td>+I and +II</td>
</tr>
<tr>
<td>manganese</td>
<td>+II, +IV, and +VII</td>
</tr>
<tr>
<td>chromium</td>
<td>+III and +VI</td>
</tr>
</tbody>
</table>

So for these elements, the oxidation state is included in the compound’s name. For example iron(III) chloride, copper(II) oxide.

6. Note that in any formula, the oxidation states must add up to zero. Look at the formula for magnesium chloride, for example:

\[ \text{MgCl}_2 \]

\[ +II \quad 2 \times -I \quad \text{Total} = \text{zero} \]

So you could use oxidation states to check that formulae are correct.
**Oxidation states change during redox reactions**

Look at the equation for the reaction between sodium and chlorine:

\[ 2\text{Na (s)} + \text{Cl}_2 (g) \rightarrow 2\text{NaCl (s)} \]

- IV 0 0 +I −I

The oxidation states are also shown, using the rules on page 96. Notice how they have changed during the reaction.

Each sodium atom loses an electron during the reaction, to form an \( \text{Na}^+ \) ion. So sodium is oxidised, and its oxidation state rises from 0 to +I.

Each chlorine atom gains an electron, to form a \( \text{Cl}^- \) ion. So chlorine is reduced, and its oxidation state falls from 0 to −I.

**If oxidation states change during a reaction, it is a redox reaction.**

- A rise in oxidation number means oxidation has occurred.
- A fall in oxidation number means reduction has occurred.

**Using oxidation states to identify redox reactions**

**Example 1** Iron reacts with sulfur to form iron(II) sulfide:

\[ \text{Fe (s)} + \text{S (s)} \rightarrow \text{FeS (s)} \]

0 0 +II −II

The oxidation states are shown, using the rules on page 96. There has been a change in oxidation states. So this is a redox reaction.

**Example 2** When chlorine is bubbled through a solution of iron(II) chloride, iron(III) chloride is formed. The equation and oxidation states are:

\[ 2\text{FeCl}_2 (aq) + \text{Cl}_2 (aq) \rightarrow 2\text{FeCl}_3 (aq) \]

+II −I 0 +III −I

There has been a change in oxidation states. So this is a redox reaction.

**Example 3** When ammonia and hydrogen chloride gases mix, they react to form ammonium chloride. The equation and oxidation states are:

\[ \text{NH}_3 (g) + \text{HCl (g)} \rightarrow \text{NH}_4\text{Cl (s)} \]

−III +I +I −I −III −I

There has been no change in oxidation states. So this is not a redox reaction.

---

**Q**

1. **a** Write a word equation for this reaction:

   \[ 2\text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{H}_2\text{O (l)} \]

   **b** Now copy out the chemical equation from **a**. Below each symbol write the oxidation state of the atoms.

   **c** Is the reaction a redox reaction? Give evidence.

   **d** Say which substance is oxidised, and which reduced.

2. Repeat the steps in question 1 for each of these equations:

   **i** \[ 2\text{KBr (s)} \rightarrow 2\text{K(s)} + \text{Br}_2 (l) \]

   **ii** \[ 2\text{KI (aq)} + \text{Cl}_2 (g) \rightarrow 2\text{KCl (aq)} + \text{I}_2 (aq) \]

3. **a** Read point 6 on page 96.

   **b** Using the idea in point 6, work out the oxidation state of the carbon atoms in carbon dioxide, \( \text{CO}_2 \).

   **c** Carbon burns in oxygen to form carbon dioxide. Write a chemical equation for the reaction.

   **d** Now using oxidation states, show that this is a redox reaction, and say which substance is oxidised, and which is reduced.

4. *Every reaction between two elements is a redox reaction.* Do you agree with this statement? Explain.
7.4 Oxidising and reducing agents

What are oxidising and reducing agents?

When hydrogen reacts with heated copper(II) oxide, the reaction is:

\[
\text{copper(II) oxide} + \text{hydrogen} \rightarrow \text{copper} + \text{water}
\]

\[
\text{CuO (s)} + \text{H}_2 (g) \rightarrow \text{Cu (s)} + \text{H}_2\text{O (l)}
\]

The copper(II) oxide is **reduced** to copper by reaction with hydrogen. So hydrogen acts as a **reducing agent**.

The hydrogen is itself **oxidised** to water, in the reaction. So copper(II) oxide acts as an **oxidising agent**.

An oxidising agent oxidises another substance – and is itself reduced. A reducing agent reduces another substance – and is itself oxidised.

Oxidising and reducing agents in the lab

Some substances have a strong drive to gain electrons. So they are strong oxidising agents. They readily oxidise other substances by taking electrons from them. Examples are oxygen and chlorine.

Some substances are strong reducing agents, readily giving up electrons to other substances. Examples are hydrogen, carbon monoxide, and reactive metals like sodium.

Some oxidising and reducing agents show a colour change when they react. This makes them useful in lab tests. Look at these three examples.

1. **Potassium manganate(VII): an oxidising agent**

Manganese is a transition element. Like other transition elements, it can exist in different oxidation states. (Look back at point 5 on page 96.)

Potassium manganate(VII) is a purple compound. Its formula is KMnO₄. In it, the manganese is in oxidation state +VII. But it is much more stable in oxidation state +II. So it is strongly driven to reduce its oxidation state to +II, by gaining electrons.

That is why potassium manganate(VII) acts as a powerful oxidising agent. It takes electrons from other substances, in the presence of a little acid.

It is itself reduced in the reaction – with a colour change:

\[
\text{MnO}_4^- (aq) \rightarrow \text{Mn}^{2+} (aq)
\]

manganate(VII) ion (purple)  \hspace{1cm} \text{manganese(II) ion} (colourless)

This colour change means that potassium manganate(VII) can be used to test for the presence of a reducing agent. If a reducing agent is present, the purple colour will fade.

Remember OILRIG!

- **Oxidation** is Loss of electrons.
- **Reduction** is Gain of electrons.

Oxidants and reductants

- Oxidising agents are also called **oxidants**.
- Reducing agents are called **reductants**.

▲ Adding potassium manganate(VII) to an unknown liquid. The purple colour is fading, so the liquid must contain a reducing agent.
2 Potassium dichromate(VI): an oxidising agent

Chromium is also a transition element, and can exist in different oxidation states. In potassium dichromate(VI) it is in oxidation state +VI. But oxidation state +III is the most stable.

So potassium dichromate(VI) is a strong oxidising agent, in the presence of acid. It reacts to gain electrons and reduce the oxidation state to +III.

Once again there is a colour change on reduction:

\[
\text{Cr}_2\text{O}_7^{2-} (aq) \rightarrow 2\text{Cr}^{3+} (aq)
\]

dichromate(VI) ion (orange) \quad \text{chromium(III) ion} (green)

This colour change means that potassium dichromate(VI) can be used to test for the presence of reducing agents.

Outside the lab, it is used to test for alcohol (ethanol) on a driver’s breath, in the breathalyser test. It oxidises ethanol to ethanal:

\[
\text{C}_2\text{H}_5\text{OH} \overset{\text{K}_2\text{Cr}_2\text{O}_7}{\longrightarrow} \text{CH}_3\text{CHO}
\]

ethanol \quad \text{ethanal}

So a colour change proves that the driver had been drinking.

3 Potassium iodide: a reducing agent

When potassium iodide solution is added to hydrogen peroxide, in the presence of sulfuric acid, this redox reaction takes place:

\[
\text{H}_2\text{O}_2 (aq) + 2\text{KI} (aq) + \text{H}_2\text{SO}_4 (aq) \rightarrow \text{I}_2 (aq) + \text{K}_2\text{SO}_4 (aq) + 2\text{H}_2\text{O} (l)
\]

hydrogen \quad \text{potassium} \quad \text{iodine} \quad \text{potassium} \quad \text{sulfate}

peroxide \quad \text{iodide} \quad \text{peroxide} \quad \text{iodide}

You can see that the hydrogen peroxide loses oxygen: it is reduced. The potassium iodide acts as a reducing agent. At the same time the potassium iodide is oxidised to iodine. This causes a colour change:

\[
2\text{I}^- (aq) \rightarrow \text{I}_2 (aq)
\]

colourless \quad \text{red-brown}

So potassium iodide is used to test for the presence of an oxidising agent.

Q 1 What is:
   a an oxidising agent?  b a reducing agent?

2 Identify the oxidising and reducing agents in these reactions, by looking at the gain and loss of oxygen:
   a \[
   2\text{Mg} (s) + \text{O}_2 (g) \rightarrow 2\text{MgO} (s)
   \]
   b \[
   \text{Fe}_2\text{O}_3 (s) + 3\text{CO} (g) \rightarrow 2\text{Fe} (l) + 3\text{CO}_2 (g)
   \]

3 Now identify the oxidising and reducing agents in these:
   a \[
   2\text{Fe} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3
   \]
   b \[
   \text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}
   \]

4 Explain why:
   a potassium manganate(VII) is a powerful oxidising agent
   b potassium iodide is used to test for oxidising agents
Checkup on Chapter 7

Revision checklist

Core curriculum
Make sure you can ...

- define **oxidation** as a gain of oxygen
- define **reduction** as a loss of oxygen
- explain that oxidation and reduction always occur together, and give an example
- explain what a **redox reaction** is
- define these terms: **oxidising agent** **reducing agent**
- identify the oxidising and reducing agents, in reactions involving oxygen

Extended curriculum
Make sure you can also ...

- define oxidation and reduction in terms of electron transfer
- explain these terms: **half-equation** **ionic equation**
- write balanced half-equations for a redox reaction, to show the electron transfer
- give the ionic equation for a reaction, by adding the balanced half-equations
- explain the term **oxidation state**
- give the usual oxidation state for these elements, in their compounds:
  - hydrogen, oxygen, aluminium
  - sodium and other Group I metals
  - calcium and other Group II metals
  - chlorine and other Group VII non-metals
- tell the oxidation state from a compound’s name, for elements with variable oxidation states
- work out the oxidation state for each element in a compound (they must add up to zero)
- give the oxidation state for each element present, in the equation for a reaction
- identify a redox reaction from changes in oxidation states, in the equation
- explain why some substances are: **strong oxidising agents** **strong reducing agents**
  - and give examples
- explain why potassium manganate(VII) and potassium dichromate(VI) are used in the lab to test for the presence of reducing agents
- explain why potassium iodide is used in the lab to test for the presence of oxidising agents

Questions

Core curriculum

1. If a substance gains oxygen in a reaction, it has been oxidised. If it loses oxygen, it has been reduced. Oxidation and reduction always take place together, so if one substance is oxidised, another is reduced.
   
   - a. First, see if you can write a word equation for each redox reaction A to F below.
   - b. Then, using the ideas above, say which substance is being oxidised, and which is being reduced, in each reaction.

   A. \( \text{Ca (s)} + \text{O}_2 (g) \rightarrow \text{2 CaO (s)} \)
   
   B. \( \text{2 CO (g)} + \text{O}_2 (g) \rightarrow \text{2 CO}_2 (g) \)
   
   C. \( \text{CH}_4 (g) + \text{2 O}_2 (g) \rightarrow \text{CO}_2 (g) + \text{2 H}_2\text{O (l)} \)
   
   D. \( \text{2 CuO (s)} + \text{C (s)} \rightarrow \text{2 Cu (s)} + \text{CO}_2 (g) \)
   
   E. \( \text{2 Fe (s)} + \text{3 O}_2 (g) \rightarrow \text{2 Fe}_2\text{O}_3 (s) \)
   
   F. \( \text{Fe}_2\text{O}_3 (s) + \text{3 CO (g)} \rightarrow \text{2 Fe (s)} + \text{3 CO}_2 (g) \)

2. a. Is this a redox reaction? Give your evidence.
   
   A. \( \text{2 Mg (s)} + \text{CO}_2 (g) \rightarrow \text{2 MgO (s)} + \text{C (s)} \)
   
   B. \( \text{SiO}_2 (s) + \text{C (s)} \rightarrow \text{Si (s)} + \text{CO}_2 (g) \)
   
   C. \( \text{NaOH (aq)} + \text{HCl (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)} \)
   
   D. \( \text{Fe (s)} + \text{CuO (s)} \rightarrow \text{FeO (s)} + \text{Cu (s)} \)
   
   E. \( \text{C (s)} + \text{PbO (s)} \rightarrow \text{CO (g)} + \text{Pb (s)} \)

   b. For each redox reaction you identify, name:
   - i. the oxidising agent
   - ii. the reducing agent.

Extended curriculum

3. All reactions in which electron transfer take place are redox reactions. This diagram shows the electron transfer during one redox reaction.

   a. What is the product of this reaction?
   
   b. Write a balanced equation for the full reaction.
   
   c. i. Which element is being oxidised?
   - ii. Write a half-equation for the oxidation.
   
   d. i. Which element is being reduced?
   - ii. Write a half-equation for the reduction of this element.
4 Redox reactions involve electron transfer.
   a Fluorine, from Group VII, reacts with lithium, from Group I, to form a poisonous white compound. What is its name?
   b Write a balanced equation for the reaction.
   c Draw a diagram to show the electron transfer that takes place during the reaction.
   d i Which element is oxidised in the reaction?
      ii Write a half-equation for this oxidation.
   e Write a half-equation for the reduction of the other element.

5 Chlorine gas is bubbled into a solution containing sodium bromide. The equation for the reaction is:
   \[ \text{Cl}_2 (g) + 2\text{NaBr (aq)} \rightarrow \text{Br}_2 (aq) + 2\text{NaCl (aq)} \]
   a Chlorine takes the place of bromine, in the metal compound. What is this type of reaction called?
   b The compounds of Group I metals are white, and give colourless solutions. What would you see as the above reaction proceeds?
   c i Write a half-equation for the reaction of the chlorine.
      ii Is the chlorine oxidised, or reduced, in this reaction? Explain.
   d Write a half-equation for the reaction of the bromide ion.
   e Reactive elements have a strong tendency to exist as ions. Which is more reactive, chlorine or bromine? Explain why you think so.
   f i Which halide ion could be used to convert bromine back to the bromide ion?
      ii Write the ionic equation for this reaction.

6 Iodine is extracted from seaweed using acidified hydrogen peroxide, in a redox reaction. The ionic equation for the reaction is:
   \[ 2\text{I}^- (aq) + \text{H}_2\text{O}_2 (aq) + 2\text{H}^+ (aq) \rightarrow \text{I}_2 (aq) + 2\text{H}_2\text{O (l)} \]
   a In which oxidation state is the iodine in seaweed?
   b There is a colour change in this reaction. Why?
   c i Is the iodide ion oxidised, or reduced?
      ii Write the half-equation for this change.
   d In hydrogen peroxide, the oxidation state of the hydrogen is +I.
   i What is the oxidation state of the oxygen in hydrogen peroxide?
   ii How does the oxidation state of oxygen change during the reaction?
   iii Copy and complete this half-equation for hydrogen peroxide:
      \[ \text{H}_2\text{O}_2 (aq) + 2\text{H}^+ (aq) + \ldots \ldots \ldots \ldots \rightarrow 2\text{H}_2\text{O (l)} \]

7 The oxidation states in a formula add up to zero.
   a Give the oxidation state of the underlined atom in each formula below:
      i aluminium oxide, \( \text{Al}_2\text{O}_3 \)
      ii ammonia, \( \text{NH}_3 \)
      iii \( \text{H}_2\text{CO}_3 \) (aq), carbonic acid
      iv phosphorus trichloride, \( \text{PCl}_3 \)
      v copper(I) chloride, \( \text{CuCl}_2 \)
      vi copper(II) chloride, \( \text{CuCl}_2 \)
   b Now comment on the compounds in v and vi.

8 The oxidising agent potassium manganate(VII) can be used to analyse the % of iron(II) present in iron tablets. Below is an ionic equation, showing the ions that take part in the reaction:
   \[ \text{MnO}_4^- (aq) + 8\text{H}^+ (aq) + 5\text{Fe}^{2+} (aq) \rightarrow \text{Mn}^{2+} (aq) + 5\text{Fe}^{3+} (aq) + 4\text{H}_2\text{O (l)} \]
   a What does the \( \text{H}^+ \) in the equation tell you about this reaction? (Hint: check page 150.)
   b Describe the colour change.
   c Which is the reducing reagent in this reaction?
   d How could you tell when all the iron(II) had reacted?
   e Write the half-equation for the iron(II) ions.

9 Potassium chromate(VI) is yellow. In acid it forms orange potassium dichromate(VI). These are the ions that give those colours:
   \[ \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}_2\text{O}_4^{2-} + \text{H}_2\text{O} \]
   a What is the oxidation state of chromium in:
      i the yellow compound?
      ii the orange compound?
   b This reaction of chromium ions is not a redox reaction. Explain why.

10 When solutions of silver nitrate and potassium chloride are mixed, a white precipitate forms. The ionic equation for the reaction is:
    \[ \text{Ag}^+ (aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl (s)} \]
    a i What is the name of the white precipitate?
       ii Is it a soluble or insoluble compound?
    b This is the precipitation of silver chloride a redox reaction or not? Explain your answer.
    c When left in light, silver chloride decomposes to form silver and chlorine gas. Write an equation for the reaction and show clearly that this is a redox reaction.
8.1 Conductors and insulators

Batteries and electric current

The photograph above shows a battery, a bulb and a rod of graphite joined or connected to each other by copper wires. (Graphite is a form of carbon.) This arrangement is called an electric circuit.

The bulb is lit: this shows that electricity must be flowing in the circuit. Electricity is a stream of electrons.

The diagram shows how the electrons move through the circuit. The battery acts like an electron pump. Electrons leave it through the negative terminal. They travel through the wire, bulb, and rod, and enter the battery again through the positive terminal.

When the electrons stream through the fine wire in the bulb, they cause it to heat up. It gets white-hot and gives out light.

Conductors and insulators

In the circuit above, the graphite and copper wire allow electricity to pass through. So they are called conductors.

But if you connect a piece of plastic or ceramic into the circuit, the bulb will not light. Plastic and ceramic do not let electricity pass through them. They are non-conductors or insulators.

Some uses for conductors and insulators

The cables that carry electricity around the country are made of aluminium and steel. Both are conductors. (Aluminium is a better conductor than steel.)

At pylons, ceramic discs support the bare cables. Since it is an insulator, the ceramic prevents the current from running down the pylon. (Dangerous!)

Copper is used for wiring, at home. It is a very good conductor. But the wires are sheathed in plastic, and plug cases are made of plastic (an insulator), for safety.
Testing substances to see if they conduct
You can test any substance to see if it conducts, by connecting it into a circuit like the one on page 102. For example:

**Tin.** A strip of tin is connected into the circuit, in place of the graphite rod. The bulb lights, so tin must be a conductor.

**Ethanol.** The liquid is connected into the circuit by placing graphite rods in it. The bulb does not light, so ethanol is a non-conductor.

**Lead bromide.** It does not conduct when solid. But if you melt it, it conducts, and gives off a choking brown vapour.

The results  These are the results from a range of tests:

1 **The only solids that conduct are the metals and graphite.** These conduct because of their free electrons (pages 61 and 62). The electrons get pumped out of one end of the solid by the battery, while more electrons flow in the other end. For the same reason, *molten* metals conduct. (It is hard to test molten graphite, because at room pressure graphite goes from solid to gas.)

2 **Molecular substances are non-conductors.** This is because they contain no free electrons, or other charged particles, that can flow through them. Ethanol (above) is made of molecules. So is petrol, paraffin, sulfur, sugar, and plastic. These never conduct, whether solid or molten.

3 **Ionic substances do not conduct when solid. But they do conduct when melted or dissolved in water – and they decompose at the same time.** An ionic substance contains no free electrons. But it does contain ions, which have a charge. The ions become free to move when the substance is melted or dissolved, and it is they that conduct the electricity. Lead bromide is ionic. It does not conduct when solid, but conducts when it melts. The brown vapour that forms is bromine. Electricity has caused the lead bromide to decompose.

**Decomposition brought about by electricity is called electrolysis.** A liquid that contains ions, and therefore conducts electricity, is called an electrolyte.

So molten lead bromide is an electrolyte. Ethanol is a non-electrolyte.

**Q**

1 What is a conductor of electricity?
2 Draw a circuit to show how you would test whether mercury conducts.
3 Explain why metals are able to conduct electricity.
4 Naphthalene is a molecular substance. Do you think it will conduct electricity when molten? Explain.
5 What is: a an electrolyte? b a non-electrolyte? Give three examples of each.
8.2 The principles of electrolysis

Electrolysis: breaking down by electricity

Any liquid that contains ions will conduct electricity. This is because the ions are free to move. But at the same time, decomposition takes place. So you can use electricity to break down a substance. The process is called electrolysis.

The electrolysis of molten lead bromide

The diagram on the right shows the apparatus.

- The graphite rods are called electrodes.
- The electrode attached to the positive terminal of the battery is also positive. It is called the anode.
- The negative electrode is called the cathode.

The molten lead bromide contains lead ions \( \text{Pb}^{2+} \) and bromide ions \( \text{Br}^- \). This shows what happens when the switch is closed:

1. Electrons flow from the negative terminal of the battery to the cathode.
2. In the liquid, the ions move to the electrode of opposite charge.
3. At the cathode (-), the \( \text{Pb}^{2+} \) ions accept electrons. Lead begins to appear below the cathode.
4. At the anode (+), the \( \text{Br}^- \) ions give up electrons. Red-brown bromine vapour bubbles off.
5. Electrons flow from the anode to the positive terminal of the battery.

The result is that the lead bromide has decomposed:

\[
\text{lead bromide} \rightarrow \text{lead} + \text{bromine} \\
\text{PbBr}_2 (l) \rightarrow \text{Pb} (l) + \text{Br}_2 (g)
\]

Note that:

- Electrons carry the current through the wires and electrodes. But the ions carry it through the liquid.
- The graphite electrodes are inert. They carry the current into the liquid, but remain unchanged. (Electrodes made of platinum are also inert.)

The electrolysis of other molten compounds

The pattern is the same for all molten ionic compounds of two elements:

**Electrolysis breaks the molten ionic compound down to its elements, giving the metal at the cathode, and the non-metal at the anode.**

So it is a very important process. We depend on it to obtain reactive metals such as lithium, sodium, potassium, magnesium, and aluminium, from compounds dug from the Earth.

Which electrode is positive?
Remember **PA**!
**P**ositive **A**node.

Obtaining aluminium
Find out how electrolysis is used to extract aluminium, on page 201.
The electrolysis of aqueous solutions

Electrolysis can also be carried out on solutions of ionic compounds in water, because the ions in solutions are free to move. But the result may be different than for the molten compound. Compare these:

| Electrolyte                        | At the cathode (−) you get ... | At the anode (+) you get ...
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>molten sodium chloride</td>
<td>sodium</td>
<td>chlorine</td>
</tr>
<tr>
<td>a concentrated solution of sodium chloride</td>
<td>hydrogen</td>
<td>chlorine</td>
</tr>
</tbody>
</table>

Why the difference? Because the water itself produces ions. Although water is molecular, a tiny % of its molecules is split up into ions:

\[ H_2O (l) \rightarrow H^+ (aq) + OH^- (aq) \]

These ions also take part in the electrolysis, so the products may change.

The rules for the electrolysis of a solution

At the cathode (−), either a metal or hydrogen forms.

1. The more reactive an element, the more it ‘likes’ to exist as ions. So if a metal is more reactive than hydrogen, its ions stay in solution and hydrogen bubbles off. (Look at the list on the right.)
2. But if the metal is less reactive than hydrogen, the metal forms.

At the anode (+), a non-metal other than hydrogen forms.

1. If it is a concentrated solution of a halide (a compound containing Cl\(^-\), Br\(^-\) or I\(^-\) ions), then chlorine, bromine, or iodine form.
2. But if the halide solution is dilute, or there is no halide, oxygen forms.

Look at these examples. Do they follow the rules?

<table>
<thead>
<tr>
<th>Electrolyte</th>
<th>At the cathode (−) you get...</th>
<th>At the anode (+) you get ...</th>
</tr>
</thead>
<tbody>
<tr>
<td>a concentrated solution of potassium bromide, KBr</td>
<td>hydrogen</td>
<td>bromine</td>
</tr>
<tr>
<td>a concentrated solution of silver nitrate, AgNO(_3)</td>
<td>silver</td>
<td>oxygen</td>
</tr>
<tr>
<td>concentrated hydrochloric acid, HCl</td>
<td>hydrogen (H(^+) is the only positive ion present)</td>
<td>chlorine</td>
</tr>
<tr>
<td>a dilute solution of sodium chloride, NaCl</td>
<td>hydrogen</td>
<td>oxygen</td>
</tr>
</tbody>
</table>

Notice that, in the last example, the water has been decomposed!

\[ H_2O (l) \rightarrow H^+ (aq) + OH^- (aq) \]

Q

1. a Which type of compounds can be electrolysed? Why?
   b What form must they be in?
2. What does electrolysis of these molten compounds give?
   a sodium chloride, NaCl  
   b aluminium oxide, Al\(_2\)O\(_3\)  
   c calcium fluoride, CaF\(_2\)  
   d lead sulfide, PbS
3. Name the products at each electrode, when these aqueous solutions are electrolysed using inert electrodes:
   a a concentrated solution of magnesium chloride, MgCl\(_2\)  
   b concentrated hydrochloric acid, HCl  
   c a dilute solution of copper(II) sulfate, CuSO\(_4\)
What happens to ions in the molten lead bromide?
In molten lead bromide, the ions are free to move. This shows what happens to them, when the switch in the circuit is closed:

First, the ions move.
Opposite charges attract.
So the positive lead ions (Pb\(^{2+}\)) move to the cathode (\(-\)).
The negative bromide ions (Br\(^{-}\)) move to the anode (\(+\)).
The moving ions carry the current.

At the cathode (\(-\)):
the lead ions each receive two electrons and become lead atoms.
The half-equation is:
Pb\(^{2+}\) (\(l\)) + 2e\(^{-}\) \(\rightarrow\) Pb (\(l\))
Lead collects on the electrode and eventually drops off it.

At the anode (\(+\)):
the bromide ions each give up an electron, and become atoms. These then pair up to form molecules.
The half-equation is:
2Br\(^{-}\) (\(l\)) \(\rightarrow\) Br\(_2\) (\(g\)) + 2e\(^{-}\)
The bromine gas bubbles off.

Remember OILRIG:
Oxidation Is Loss of electrons,
Reduction Is Gain of electrons.

Overall, electrolysis is a redox reaction.
Reduction takes place at the cathode and oxidation at the anode.

The free ions move.
Ions gain electrons: reduction.
Ions lose electrons: oxidation.

The reactions for other molten compounds follow the same pattern.

For a concentrated solution of sodium chloride
This time, ions from water are also present:

The solution contains Na\(^{+}\) ions and Cl\(^{-}\) ions from the salt, and H\(^{+}\) and OH\(^{-}\) ions from water.
The positive ions go to the cathode and the negative ions to the anode.

At the cathode, the H\(^{+}\) ions accept electrons, since hydrogen is less reactive than sodium:
2H\(^{+}\) (\(aq\)) + 2e\(^{-}\) \(\rightarrow\) H\(_2\) (\(g\))
The hydrogen gas bubbles off.

At the anode, the Cl\(^{-}\) ions give up electrons more readily than the OH\(^{-}\) ions do.
2Cl\(^{-}\) (\(aq\)) \(\rightarrow\) Cl\(_2\) (\(aq\)) + 2e\(^{-}\)
The chlorine gas bubbles off.

When the hydrogen and chlorine bubble off, Na\(^{+}\) and OH\(^{-}\) ions are left behind – so a solution of sodium hydroxide is formed.
For a dilute solution of sodium chloride

The same ions are present as before. But now the proportion of Na$^+$ and Cl$^-$ ions is lower, since this is a dilute solution. So the result will be different.

At the cathode, hydrogen ‘wins’ as before, and bubbles off:

$$4H^+ (aq) + 4e^- \rightarrow 2H_2 (g)$$

(4 electrons are shown, to balance the half-equation at the anode.)

At the anode, OH$^-$ ions give up electrons, since not many Cl$^-$ ions are present. Oxygen bubbles off:

$$4OH^- (aq) \rightarrow O_2 (g) + 2H_2O (l) + 4e^-$$

When the hydrogen and oxygen bubble off, the Na$^+$ and Cl$^-$ ions are left behind. So we still have a solution of sodium chloride! The overall result is that water has been decomposed.

Writing the half-equations for electrode reactions

You may be asked to write half-equations for the reactions at electrodes. This table shows the steps.

<table>
<thead>
<tr>
<th>The steps</th>
<th>Example: the electrolysis of molten magnesium chloride</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 First, name the ions present, and the products.</td>
<td>Magnesium ions and chloride ions are present. Magnesium and chlorine form.</td>
</tr>
<tr>
<td>2 Write each half-equation correctly.</td>
<td>Ions: Mg$^{2+}$ and Cl$^-$</td>
</tr>
<tr>
<td>- Give the ion its correct charge.</td>
<td><strong>At the cathode:</strong></td>
</tr>
<tr>
<td>- Remember, positive ions go to the cathode, and negative ions to the anode.</td>
<td>Mg$^{2+}$ + 2e$^-$ $\rightarrow$ Mg</td>
</tr>
<tr>
<td>- Write the correct symbol for the element that forms.</td>
<td><strong>At the anode:</strong></td>
</tr>
<tr>
<td>- For example, Cl$_2$ for chlorine (not Cl).</td>
<td>2Cl$^-$ $\rightarrow$ Cl$_2$ + 2e$^-$ (two Cl$^-$ ions, so a total charge of 2$^-$) Note that it is also correct to write the anode reaction as: 2Cl$^-$ $-$ 2e$^-$ $\rightarrow$ Cl$_2$</td>
</tr>
<tr>
<td>- The number of electrons in the equation should be the same as the total charge on the ion(s) in it.</td>
<td></td>
</tr>
<tr>
<td>3 You could then add the state symbols.</td>
<td>Mg$^{2+}$ (l) + 2e$^-$ $\rightarrow$ Mg (l)</td>
</tr>
<tr>
<td></td>
<td>2Cl$^-$ (l) $\rightarrow$ Cl$_2$ (g) + 2e$^-$</td>
</tr>
</tbody>
</table>

Q1 At which electrode does reduction always take place?
Q2 Give the half-equation for the reaction at the anode, during the electrolysis of these molten compounds:
   a potassium chloride   b calcium oxide
Q3 Give the two half-equations for the electrolysis of:
   a a concentrated solution of hydrochloric acid, HCl
   b a dilute solution of sodium nitrate, NaNO$_3$
   c a dilute solution of copper(II) chloride, CuCl$_2$
8.4 The electrolysis of brine

What is brine?
Brine is a concentrated solution of sodium chloride, or common salt. It can be obtained by pumping water into salt mines to dissolve the salt, or by evaporating seawater.

Brine might not sound very exciting – but from it, we get chemicals needed for thousands of products we use every day. When it undergoes electrolysis, the overall reaction is:

\[ 2\text{NaCl (aq)} + 2\text{H}_2\text{O (l)} \xrightarrow{\text{electrolysis}} 2\text{NaOH (aq)} + \text{Cl}_2 (g) + \text{H}_2 (g) \]

The electrolysis

The diagram below shows one type of cell used for this electrolysis. The anode is made of titanium, and the cathode of steel. Now look at the diaphragm down the middle of the cell. Its function is to let ions through, but keep the gases apart. (So the cell is called a diaphragm cell.)

The ions present are \( \text{Na}^+ \) and \( \text{Cl}^- \) from the salt, and \( \text{H}^+ \) and \( \text{OH}^- \) from the water. The reactions at the electrodes are exactly as shown at the bottom of page 106. (Look back at them.)

At the cathode  
Hydrogen is discharged in preference to sodium:

\[ 2\text{H}^+ (aq) + 2e^- \rightarrow \text{H}_2 (g) \]

As usual at the cathode, this is a reduction.

At the anode  
Chlorine is discharged in preference to oxygen:

\[ 2\text{Cl}^- (aq) \rightarrow \text{Cl}_2 (g) + 2e^- \]

As usual at the anode, this is an oxidation.

The two gases bubble off. \( \text{Na}^+ \) and \( \text{OH}^- \) ions are left behind, giving a solution of sodium hydroxide. Some of the solution is evaporated to give a more concentrated solution, and some is evaporated to dryness, giving solid sodium hydroxide.
What the products are used for
The electrolysis of brine is an important process, because the products are so useful. Look at these:

**Chlorine, a poisonous yellow-green gas**
Used for making ...
- the plastic PVC (nearly 1/3 of it used for this)
- solvents for degreasing and drycleaning
- medical drugs (a large % of these involve chlorine)
- weedkillers and pesticides (most of these involve chlorine)
- paints and dyestuffs
- bleaches
- hydrogen chloride and hydrochloric acid
It is also used as a sterilising agent, to kill bacteria in water supplies and swimming pools.

**Sodium hydroxide solution, alkaline and corrosive**
Used in making ...
- soaps
- detergents
- viscose (rayon) and other textiles
- paper (like the paper in this book)
- ceramics (tiles, furnace bricks, and so on)
- dyes
- medical drugs

**Hydrogen, a colourless flammable gas**
Used ...
- in making nylon
- to make hydrogen peroxide
- to ‘harden’ vegetable oils to make margarine
- as a fuel in hydrogen fuel cells

Of the three chemicals, chlorine is the most widely used. Around 50 million tonnes of it are produced each year, around the world.

---

**Q**

1. What is brine? Where is it obtained from?
2. Write a word equation for the electrolysis of brine.
3. Draw a rough sketch of the diaphragm cell. Mark in where the oxidation and reduction reactions take place in it, and write the half-equations for them.
4. What is the diaphragm for, in the diaphragm cell?
5. The electrolysis of brine is a very important process. 
   a. Explain why.
   b. Give three uses for each of the products.
6. Your job is to keep a brine electrolysis plant running safely and smoothly. Try to think of three or four safety precautions you might need to take.
Two more uses of electrolysis

When electrodes are not inert

A solution of copper(II) sulfate contains blue Cu²⁺ ions, SO₄²⁻ ions, and H⁺ and OH⁻ ions from water. Electrolysis of the solution will give different results, depending on the electrodes. Compare these:

A Using inert electrodes (carbon or platinum)

At the cathode  Copper ions are discharged:
2Cu²⁺ (aq) + 4e⁻ → 2Cu (s)
The copper coats the electrode.

At the anode  Oxygen bubbles off:
4OH⁻ (aq) → 2H₂O (l) + O₂ (g) + 4e⁻

So copper and oxygen are produced. This fits the rules on page 105. The blue colour of the solution fades as the copper ions are discharged.

B Using copper electrodes

At the cathode  Again, copper is formed, and coats the electrode:
Cu²⁺ (aq) + 2e⁻ → Cu (s)

At the anode  The anode dissolves, giving copper ions in solution:
Cu (s) → Cu²⁺ (aq) + 2e⁻

So this time, the electrodes are not inert. The anode dissolves, giving copper ions. These move to the cathode, to form copper. So copper moves from the anode to the cathode. The colour of the solution does not fade.

The idea in B leads to two important uses of electrolysis: for refining (or purifying) copper, and for electroplating.

Refining copper

The anode is made of impure copper. The cathode is pure copper. The electrolyte is dilute copper(II) sulfate solution.

The copper deposited on the cathode is over 99.9% pure.
The sludge may contain valuable metals such as platinum, gold, silver, and selenium. These are recovered and sold.
The purer it is, the better copper is at conducting electricity. Highly refined copper is used for the electrics in cars. A car like this will contain more than 1 km of copper wiring.

**Electroplating**

Electroplating means using electricity to coat one metal with another, to make it look better, or to prevent corrosion. For example, steel car bumpers are coated with chromium. Steel cans are coated with tin to make tins for food. And cheap metal jewellery is often coated with silver.

The drawing on the right shows how to electroplate a steel jug with silver. The jug is used as the cathode. The anode is made of silver. The electrolyte is a solution of a soluble silver compound, such as silver nitrate.

**At the anode** The silver dissolves, forming silver ions in solution:

\[
\text{Ag} (s) \rightarrow \text{Ag}^+ (aq) + e^- 
\]

**At the cathode** The silver ions are attracted to the cathode. There they receive electrons, forming a coat of silver on the jug:

\[
\text{Ag}^+ (aq) + e^- \rightarrow \text{Ag} (s)
\]

When the layer of silver is thick enough, the jug is removed.

![Silverplating: electroplating with silver. When the electrodes are connected to a power source, electroplating begins.](image)

**To electroplate**

In general, to electroplate an object with metal X, the set-up is:

- **cathode** – object to be electroplated
- **anode** – metal X
- **electrolyte** – a solution of a soluble compound of X.

---

1. Copper(II) ions are blue. When copper(II) sulfate solution is electrolysed, the blue solution:
   - a loses its colour when carbon electrodes are used
   - b keeps its colour when copper electrodes are used
   Explain each of these observations.
2. If you want to purify a metal by electrolysis, will you make it the anode or the cathode? Why?
3. Describe the process of refining copper.
4. What does electroplating mean?
5. Steel cutlery is often electroplated with nickel. Why?
6. You plan to electroplate steel cutlery with nickel.
   - a What will you use as the anode?
   - b What will you use as the cathode?
   - c Suggest a suitable electrolyte.
Checkup on Chapter 8

Revision checklist

Core curriculum

Make sure you can …

- define the terms conductor and insulator
- give examples of how we make use of conductors and insulators
- explain what these terms mean:
  - electrolysis
  - electrolyte
  - electrode
  - inert electrode
  - anode
  - cathode
- explain why an ionic compound must be melted, or dissolved in water, for electrolysis
- predict what will be obtained at each electrode, in the electrolysis of a molten ionic compound
- say what halides are
- say why the products of electrolysis may be different, when a compound is dissolved in water, rather than melted
- give the general rules for the products at the anode and cathode, in the electrolysis of a solution
- name the product at each electrode, for the electrolysis of:
  - concentrated hydrochloric acid
  - a concentrated solution of sodium chloride
- explain what electroplating is, and why it is used
- describe how electroplating is carried out

Extended curriculum

Make sure you can also …

- predict the products, for the electrolysis of halides in dilute and concentrated solutions
- describe the reactions at the electrodes, during the electrolysis of:
  - a molten halide such as lead bromide
  - a dilute solution of a halide such as sodium chloride
  - a concentrated solution of a halide
- and write half-equations for them
- describe the electrolysis of brine, and name the three products, and give some uses for them (you will not be asked for a diagram of the cell)
- describe the differences, when the electrolysis of copper(II) sulfate is carried out:
  - using inert electrodes (carbon or platinum)
  - using copper electrodes
- describe how electrolysis is used to refine impure copper, and say why this is important

Questions

Core curriculum

1. Electrolysis of molten lead bromide is carried out:

   - The bulb will not light until the lead bromide has melted. Why not?
   - What will be seen at the anode?
   - Name the substance in b.
   - What will be formed at the cathode?

2. Six substances A to F were dissolved in water, and connected in turn into the circuit below. A represents an ammeter, which is used to measure current. The table shows the results.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Current (amperes)</th>
<th>At cathode</th>
<th>At anode</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>0.8</td>
<td>copper</td>
<td>chlorine</td>
</tr>
<tr>
<td>B</td>
<td>1.0</td>
<td>hydrogen</td>
<td>chlorine</td>
</tr>
<tr>
<td>C</td>
<td>0.0</td>
<td>—</td>
<td>—</td>
</tr>
<tr>
<td>D</td>
<td>0.8</td>
<td>copper</td>
<td>oxygen</td>
</tr>
<tr>
<td>E</td>
<td>1.2</td>
<td>hydrogen</td>
<td>oxygen</td>
</tr>
<tr>
<td>F</td>
<td>0.7</td>
<td>silver</td>
<td>oxygen</td>
</tr>
</tbody>
</table>

   - a Which solution conducts best?
   - b Which solution is a non-electrolyte?
   - c Which solution could be:
     - i silver nitrate?
     - ii copper(II) sulfate?
     - iii copper(II) chloride?
     - iv sodium hydroxide?
     - v sugar?
     - vi concentrated hydrochloric acid?
   - d Explain how the current is carried:
     - i within the electrolytes
     - ii in the rest of the circuit

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Extended curriculum

3 The electrolysis below produces gases A and B.

\[\text{gas A} \quad \text{carbon anode} \quad \text{carbon cathode} \quad \text{gas B} \quad \text{concentrated solution of sodium chloride} \]

\[\text{power supply} \]

a Why does the solution conduct electricity?
b Identify each gas, and describe a test you could carry out to confirm its identity.
c Name one product manufactured from:
   i gas A
   ii gas B
d i Write half-equations to show how the two gases are produced.
   ii The overall reaction is a redox reaction. Explain why.
e The solution remaining after the electrolysis will turn litmus paper blue.
   i What is the name of this solution?
   ii State one chemical property for it.

4 a List the ions that are present in concentrated solutions of:
   i sodium chloride    ii copper(II) chloride
b Explain why and how the ions move, when each solution is electrolysed using platinum electrodes.
c Write the half-equation for the reaction at:
   i the anode
   ii the cathode during the electrolysis of each solution.
d Explain why the anode reactions for both solutions are the same.
e i The anode reactions will be different if the solutions are made very dilute. Explain why.
   ii Write the half-equations for the new anode reactions.
f Explain why copper is obtained at the cathode, but sodium is not.
g Name another solution that will give the same products as the concentrated solution of sodium chloride does, on electrolysis.
h Which solution in a could be the electrolyte in an electroplating experiment?

5 Molten lithium chloride contains lithium ions (Li\(^+\)) and chloride ions (Cl\(^-\)).
a Copy the following diagram and use arrows to show which way:
   i the ions move when the switch is closed
   ii the electrons flow in the wires

\[\text{chloride ion} \quad + \quad \text{lithium ion} \quad \text{switch} \]

b i Write equations for the reaction at each electrode, and the overall reaction.
   ii Describe each of the reactions using the terms reduction, oxidation and redox.

6 This question is about the electrolysis of a dilute aqueous solution of lithium chloride.
a Give the names and symbols of the ions present.
b Say what will be formed, and write a half-equation for the reaction:
   i at the anode
   ii at the cathode
c Name another compound that will give the same products at the electrodes.
d How will the products change, if a concentrated solution of lithium chloride is used?

7 An experiment is needed, to see if an iron object can be electroplated with chromium.
a Suggest a solution to use as the electrolyte.
b i Draw a labelled diagram of the apparatus that could be used for the electroplating.
   ii Show how the electrons will travel from one electrode to the other.
c Write half-equations for the reactions at each electrode.
d At which electrode does oxidation take place?
e The concentration of the solution does not change. Why not?

8 Nickel(II) sulfate (NiSO\(_4\)) is green. A solution of this salt is electrolysed using nickel electrodes.
a Write a half-equation for the reaction at each electrode.
b At which electrode does reduction take place?
   Explain your answer
c What happens to the size of the anode?
d The colour of the solution does not change, during the electrolysis. Explain why.
e Suggest one industrial use for this electrolysis.
Energy changes in reactions

During a chemical reaction, there is always an energy change.

Energy is given out or taken in. The energy is usually in the form of heat.
(But some may be in the form of light and sound.)
So reactions can be divided into two groups: exothermic and endothermic.

Exothermic reactions

Exothermic reactions give out energy. So there is a temperature rise.

Here are three examples:

To start off the reaction between iron and sulfur, you must heat the mixture. But soon it glows red hot — without the Bunsen burner!

Mixing silver nitrate and sodium chloride solutions gives a white precipitate of silver chloride — and a temperature rise.

When you add water to lime (calcium oxide) heat is given out, so the temperature rises. Here the rise is being measured.

These reactions can be described as:

reactants → products + energy

The total energy is the same on each side of the arrow, in a reaction. So in exothermic reactions, the products have lower energy than the reactants.

This is shown on the energy level diagram on the right.

The energy change

Energy is measured in kilojoule (kJ). For reaction A above:

Fe (s) + S (s) → FeS (s)  the energy change = −100 kJ

So 100 kJ of energy is given out when the amounts of reactants in the equation (56 g of iron and 32 g of sulfur, or 1 mole of each) react together.

The minus sign shows that energy is given out.

Other examples of exothermic reactions

All these are exothermic:

- the neutralisation of acids by alkalis.
- the combustion of fuels. We burn fuels to obtain heat for cooking, heating homes, and so on. The more energy they give out, the better!
- respiration in your body cells. It provides the energy to keep your heart and lungs working, and for warmth and movement.
**Endothermic reactions**

Endothermic reactions take in energy from their surroundings. Here are three examples:

When barium hydroxide reacts with ammonium chloride, the temperature falls so sharply that water under the beaker will freeze!

Sherbet is citric acid plus the base sodium hydrogen carbonate. The neutralisation that occurs takes in heat – so your tongue cools.

The crucible contains calcium carbonate. If you keep on heating, it will all decompose to calcium oxide and carbon dioxide.

These reactions can be described as:

**reactants + energy → products**

The energy is transferred from the surroundings: in D from the air and wet wood, in E from your tongue, and in F from the Bunsen burner. Since energy is taken in, the products must have higher energy than the reactants. This is shown on the energy level diagram on the right.

**The energy change**

For reaction F above:

\[ \text{CaCO}_3 (s) \rightarrow \text{CaO} (s) + \text{CO}_2 (g) \quad \text{the energy change} = +178 \, \text{kJ} \]

So 178 kJ of energy is needed to make 100 g (or 1 mole) of CaCO₃ decompose. The **plus sign shows that energy is taken in**.

**Other examples of endothermic reactions**

Reactions D and E above are **spontaneous**. They start off on their own. But many endothermic reactions are like F, where energy must be put in start the reaction and keep it going. For example:

- reactions that take place in cooking.
- photosynthesis. This is the process in which plants convert carbon dioxide and water to glucose. It depends on the energy from sunlight.

**Q**

1. Is it exothermic or endothermic?
   - a the burning of a candle
   - b the reaction between sodium and water
   - c the change from raw egg to fried egg

2. Which unit is used to measure energy changes?

3. \[ 2\text{Na} (s) + \text{Cl}_2 (g) \rightarrow 2\text{NaCl} (s) \]
   - The energy change for this reaction is \(-822.4 \, \text{kJ}\).
   - What can you conclude about the reaction?

4. Draw an energy level diagram for:
   - a an endothermic reaction
   - b an exothermic reaction
9.2 Explaining energy changes

Making and breaking bonds

In a chemical reaction, bonds must first be broken. Then new bonds form. Breaking bonds takes in energy. Making bonds releases energy.

Example 1: an exothermic reaction

Hydrogen reacts with chlorine in sunshine, to form hydrogen chloride:

1. First, the bonds in the hydrogen and chlorine molecules must be broken. Energy must be taken in, for this. (Energy from sunshine will do!)

But the energy taken in for step 1 is less than the energy given out in step 2. So this reaction gives out energy, overall. It is exothermic.

If the energy taken in to break bonds is less than the energy released in making bonds, the reaction is exothermic.

Example 2: an endothermic reaction

If you heat ammonia strongly, it breaks down to nitrogen and hydrogen. Here we use lines to show the bonds. (Note the triple bond in nitrogen.)

1. First, the bonds in ammonia must be broken. Energy must be taken in, for this. (You supply it by heating.)

This time, the energy taken in for step 1 is greater than the energy given out in step 2. So the reaction takes in energy, overall. It is endothermic.

If the energy taken in to break bonds is greater than the energy released in making bonds, the reaction is endothermic.

Bond energies

The energy needed to make or break bonds is called the bond energy. Look at the list on the right. 242 kJ must be supplied to break the bonds in a mole of chlorine molecules, to give chlorine atoms. If these atoms join again to form molecules, 242 kJ of energy are given out.

The bond energy is the energy needed to break bonds, or released when these bonds form. It is given in kJ/mole.

<table>
<thead>
<tr>
<th>Bond energy (kJ/mole)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H–H</td>
</tr>
<tr>
<td>Cl–Cl</td>
</tr>
<tr>
<td>H–Cl</td>
</tr>
<tr>
<td>C–C</td>
</tr>
<tr>
<td>C–Cl</td>
</tr>
<tr>
<td>C–O</td>
</tr>
<tr>
<td>C–H</td>
</tr>
<tr>
<td>O=O</td>
</tr>
<tr>
<td>O–H</td>
</tr>
<tr>
<td>N=N</td>
</tr>
<tr>
<td>N–H</td>
</tr>
</tbody>
</table>
Calculating the energy changes in reactions
So let’s calculate the energy change for those reactions on page 116.

1. **The exothermic reaction between hydrogen and chlorine**
   \[ H - H + Cl - Cl \rightarrow 2 H - Cl \]
   - **Energy in** to break each mole of bonds:
     - \( 1 \times H - H \): 436 kJ
     - \( 1 \times Cl - Cl \): 242 kJ
     - **Total energy in** = 678 kJ
   - **Energy out** from the two moles of bonds forming:
     - \( 2 \times H - Cl \): \( 2 \times 431 = 862 \) kJ
   - **Energy in – energy out** = \( 678 \) kJ – \( 862 \) kJ = \( -184 \) kJ
   - So the reaction gives out \( 184 \) kJ of energy, overall. Its energy level diagram is shown on the right.

2. **The endothermic decomposition of ammonia**
   \[ 2 N - H \rightarrow N\equiv N + 3 H - H \]
   - **Energy in** to break the two moles of bonds:
     - \( 6 \times N - H \): \( 6 \times 391 = 2346 \) kJ
   - **Energy out** from the four moles of bonds forming:
     - \( 1 \times N\equiv N \): 946 kJ
     - \( 3 \times H - H \): \( 3 \times 436 = 1308 \) kJ
     - **Total energy out** = 2254 kJ
   - **Energy in – energy out** = \( 2346 \) kJ – \( 2254 \) kJ = \( +92 \) kJ
   - So the reaction takes in \( 92 \) kJ of energy, overall. Look at its energy level diagram.

**Starting a reaction off**
To start a reaction, bonds must be broken. As you saw, this needs energy.

- For some reactions, not much energy is needed. Just mix the reactants at room temperature. (For example, reactions B and C on page 114.)
- Some exothermic reactions need heat from a Bunsen burner just to start bonds breaking. Then the energy given out by the reaction breaks further bonds. (For example, reaction A on page 114.)
- But for endothermic reactions like the decomposition of calcium carbonate (reaction F on page 115), you must continue heating until the reaction is complete.

**Q**
1. Two steps must take place, to go from reactants to products. What are they?
2. Some reactions are endothermic. Explain why, using the ideas of bond breaking and bond making.
3. Hydrogen reacts with oxygen. Draw the equation for the reaction as above, with lines to show the bonds.
4. Now see if you can calculate the energy change for the reaction in 3, using the bond energy table on page 116.
9.3 Energy from fuels

What is a fuel?
A fuel is any substance we use to provide energy. We convert the chemical energy in the fuel into another form of energy. We burn most fuels, to obtain their energy in the form of heat.

The fossil fuels
The fossil fuels — coal, petroleum (oil), and natural gas (methane) — are the main fuels used around the world. We burn them to release heat.

We burn fossil fuels in power stations, to heat water to make steam. A jet of steam drives the turbines that generate electricity. We burn them in factories to heat furnaces, and in homes for cooking and heating. (Kerosene, from petroleum, is also used in lamps.) Petrol and diesel (from petroleum) are burned in engines, to give the hot gas that moves the pistons. These then make the wheels turn.

The world uses up enormous quantities of the fossil fuels. For example, nearly 12 million tonnes of petroleum every day!

So what makes a good fuel?
These are the main questions to ask about a fuel:

- **How much heat does it give out?** We want as much heat as possible, per tonne of fuel.
- **Does it cause pollution?** If it causes a lot of pollution, we may be better off without it!
- **Is it easily available?** We need a steady and reliable supply.
- **Is it easy and safe to store and transport?** Most fuels catch fire quite easily, so safety is always an issue.
- **How much does it cost?** The cheaper the better.

The fossil fuels give out a lot of heat. But they cause pollution, with coal the worst culprit. The pollutants include carbon dioxide, which is linked to global warming, and other gases that cause acid rain. (See page 214.)

What about availability? We are using up the fossil fuels fast. Some experts say we could run out of petroleum and gas within 50 years. But there is probably enough coal to last several hundred years.

The burning of fuel is an exothermic reaction. The more heat given out the better — as long as the fuel is safe to use.
Two fuels growing in importance
Because of fears about global warming, and dwindling supplies of petroleum and gas, there is a push to use new fuels. Like these two:

**Ethanol** This is an alcohol, with the formula C₂H₅OH. It can be made from any plant material. For example, it is made from sugar cane in Brazil, and from corn (maize) in the USA. It is used in car engines, on its own or mixed with petrol. See pages 256 – 257 for more.

**Hydrogen** This gas burns explosively in oxygen, giving out a lot of energy – so it is used to fuel space rockets. It is also used in *fuel cells* (without burning) to give energy in the form of electricity. See page 121 for more.

Different amounts of heat
Some fuels give out a lot more heat than others. Compare these:

<table>
<thead>
<tr>
<th>Fuel</th>
<th>Equation for burning in oxygen</th>
<th>Heat given out per gram of fuel/kJ</th>
</tr>
</thead>
<tbody>
<tr>
<td>natural gas (methane)</td>
<td>CH₄ (g) + 2O₂ (g) → CO₂ (g) + 2H₂O (l)</td>
<td>− 55</td>
</tr>
<tr>
<td>ethanol</td>
<td>C₂H₅OH (l) + 3O₂ (g) → 2CO₂ (g) + 3H₂O (l)</td>
<td>− 86</td>
</tr>
<tr>
<td>hydrogen</td>
<td>2H₂ (g) + O₂ (g) → 2H₂O (g)</td>
<td>− 143</td>
</tr>
</tbody>
</table>

**Nuclear fuels**

*Nuclear fuels* are *not* burned. They contain unstable atoms called *radioisotopes* (page 34). Over time, these break down naturally into new atoms, giving out radiation and a lot of energy.

But you can also *force* radioisotopes to break down, by shooting neutrons at them. That is what happens in a *nuclear power station*. The energy given out is used to heat water, to make jets of steam to drives the turbines for generating electricity.

The radioisotope uranium-235 is commonly used in nuclear fuels. When it decays, the new atoms that form are also unstable, and break down further.

Nuclear fuel has two big advantages:
- It gives out huge amounts of energy. A pellet of nuclear fuel the size of a pea can give as much energy as a tonne of coal.
- No carbon dioxide or other polluting gases are formed.

But it is not all good news. An explosion in a nuclear power station could spread radioactive material over a huge area, carried in the wind.

The waste material produced in a nuclear power station is also radioactive, and may remain very dangerous for hundreds of years. Finding a place to store it safely is a major problem.

**Q**

1. Sketch an energy level diagram that you think shows:
   - a good fuel
   - a very poor fuel

2. What else do you need to think about, to decide whether a substance would make a good fuel?

3. Look at the table above. From *all* the information given, which of the three fuels do you think is best? Explain.

4. The fuel butane (C₄H₁₀) burns to give the same products as methane. Write a balanced equation for its combustion.
9.4 Giving out energy as electricity

Electricity: a form of energy
Electricity is a current of electrons. Like heat, it is a form of energy.
When you burn a fuel, chemical energy is converted to heat.
But a reaction can also give out energy as electricity.

Electricity from a redox reaction

Connect a strip of magnesium, a strip of copper, and a light bulb, like this. (Note: no battery!)
Nothing happens.

So what is going on?

1. Magnesium is more reactive than copper. (See the list on the right.)
   That means it has a stronger drive to form ions. So the magnesium atoms give up electrons, and go into solution as ions:
   \[ \text{Mg} (s) \rightarrow \text{Mg}^{2+} (aq) + 2e^- \quad \text{(magnesium is oxidised)} \]

2. The electrons flow along the wire to the copper strip, as a current.

3. The solution contains Na\(^+\) and Cl\(^-\) ions from sodium chloride, and some H\(^+\) and OH\(^-\) ions from water. Hydrogen is less reactive than sodium, so the H\(^+\) ions accept electrons from the copper strip:
   \[ 2\text{H}^+ (aq) + 2e^- \rightarrow \text{H}_2 (g) \quad \text{(hydrogen ions are reduced)} \]

So a redox reaction is giving out energy in the form of a current.

A simple cell
The metal strips, wire, and beaker of solution above form a simple cell.
Electrons flow from the magnesium strip, so it is called the positive pole.
The copper strip is the negative pole. The solution is the electrolyte.

A simple cell consists of two metals and an electrolyte. The more reactive metal is the negative pole of the cell. Electrons flow from it.

Other metals can also be used, as long as they differ in reactivity.
And any solution can be used, as long as it contains ions.

You could connect a voltmeter into the circuit, to measure the voltage.
The bigger the difference in reactivity of the metals, the larger the voltage, and the more brightly the bulb will light. Find out more on page 190.
The hydrogen fuel cell

In the hydrogen fuel cell, hydrogen and oxygen combine without burning. It is a redox reaction. The energy is given out as an electric current.

Like the simple cell, the fuel cell has a negative pole that gives out electrons, a positive pole that accepts them, and an electrolyte.

Both poles are made of carbon.

The negative pole is surrounded by hydrogen, and the positive pole by oxygen (in air). The electrolyte contains OH\(^-\) ions.

At the negative pole

Hydrogen loses electrons to the OH\(^-\) ions. It is oxidised:

\[
2\text{H}_2 (g) + 4\text{OH}^- (aq) \rightarrow 4\text{H}_2\text{O} (l) + 4e^- 
\]

A current of electrons flows through the wire to the positive pole. You can make use of it on the way. For example, pass it through light bulbs to light your home.

At the positive pole

The electrons are accepted by oxygen molecules. Oxygen is reduced to OH\(^-\) ions:

\[
\text{O}_2 (g) + 2\text{H}_2\text{O} (l) + 4e^- \rightarrow 4\text{OH}^- (aq)
\]

But the concentration of OH\(^-\) ions in the electrolyte does not increase. Why not?

Adding the two half-equations gives the full equation for the redox reaction:

\[
2\text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{H}_2\text{O} (l)
\]

So the overall reaction is that hydrogen and oxygen combine to form water.

Advantages of the hydrogen fuel cell

- Only water is formed. No pollutants!
- The reaction gives out plenty of energy. 1 kg of hydrogen gives about 2.5 times as much energy as 1 kg of natural gas (methane).
- We will not run out of hydrogen. It can be made by the electrolysis of water with a little acid added. Solar power could provide cheap electricity for this. Scientists also hope to make it from waste plant material, using bacteria.

But there is a drawback. Hydrogen is very flammable. A spark or lit match will cause a mixture of hydrogen and air to explode. So it must be stored safely.

- This car has a hydrogen fuel cell instead of a petrol engine.

Q

2. In a simple cell, which metal gives up electrons to produce the current: the more reactive or less reactive one?
3. A wire connects strips of magnesium and copper, standing in an electrolyte. Bubbles appear at the copper strip. Why?
5. a. In the hydrogen fuel cell, what is the fuel?
   b. How are the electrons transferred in this cell?
   c. What type of electrolyte is used?
The batteries in your life

Batteries and you
We depend a lot on batteries. Cars and buses will not start without them. Torches need them. So do mobile phones, laptops, cameras, iPods …
The diagram on the right shows a simple model of a battery (or cell). All batteries contain two solid substances of different reactivity, and an electrolyte. The more reactive substance gives up electrons more readily. These flow out of the battery as an electric current.
Since the more reactive substance provides the electrons, it is called the negative pole, or negative electrode, or negative terminal.
The simple cell, shown on page 120, is the simplest battery of all. But it is not very practical. You could not use it in a torch, for example, and it does not have enough voltage to start a car. You need other types of battery.

A torch battery
Torch batteries are ‘dry’, and easy to carry around:

The metal case is the negative pole. It is usually zinc.
The electrolyte is often sodium or potassium hydroxide, made into a paste that will not leak.
(So these batteries are called alkaline batteries.)

The current does some work (for example makes a bulb light)
flow of electrons (electricity)
more reactive
less reactive
electrolyte

Gotcha! Thanks to redox reactions in the torch battery.

A car battery
A car battery consists of plates of lead and lead(IV) oxide, standing in a solution of sulfuric acid, as shown on the right. This is what happens:

1. The lead plate reacts with the sulfuric acid, giving lead(II) sulfate:
   \[ \text{Pb} (s) + \text{H}_2\text{SO}_4 (aq) \rightarrow \text{PbSO}_4 (s) + 2\text{H}^+ + 2e^- \]
The lead(II) sulfate coats the plate.

2. The electrons go off through the wire as an electric current. It gets the car’s starter motor working.

3. The electrons flow back through the wire to the lead(IV) oxide plate. This also reacts with the acid to form lead(II) sulfate, which coats the plate:
   \[ \text{PbO}_2 (s) + \text{H}_2\text{SO}_4 (aq) + 2\text{H}^+ + 2e^- \rightarrow \text{PbSO}_4 (s) + 2\text{H}_2\text{O} (l) \]

In fact the car battery usually has six sets of plates linked together, giving a total voltage of 12 volts.
**Recharging the car battery**

While the car battery is running, the plates are being coated with lead(II) sulfate, and the sulfuric acid is being used up. So if it runs for long enough, the battery will stop working, or ‘go flat’.

But it needs to run for only a short time, to start the car. And then something clever happens: electricity generated by the motor causes the reactions to reverse. The lead(II) sulfate on the plates is converted back to lead and lead(IV) oxide, ready for next time.

**A button battery**

You probably have a button battery in your watch. Button batteries often use lithium as the negative terminal. Here is a cross-section through one:

![Button battery diagram](image)

Lithium is a good choice because it is highly reactive: it gives up electrons easily. These flow out through the top of the steel case, to the connection in your watch. They flow back through the lower part of the case, and $\text{Mn}^{4+}$ ions accept them, to become $\text{Mn}^{3+}$ ions.

**A lithium-ion battery**

Lithium-ion batteries are **rechargeable**. So they are used in laptops, mobile phones, and iPods.

The battery consists of thin sheets of lithium cobalt oxide ($\text{LiCoO}_2$), and graphite (carbon). The electrolyte is a solution of a lithium salt in an organic solvent. This is how the battery works:

<table>
<thead>
<tr>
<th>When it is charging</th>
<th>When you use it</th>
</tr>
</thead>
<tbody>
<tr>
<td><img src="image" alt="Diagram" /></td>
<td><img src="image" alt="Diagram" /></td>
</tr>
</tbody>
</table>

When your phone is charging, the graphite becomes negative, and attracts lithium ions from the lithium cobalt oxide. When you use it, the lithium ions flow back to lithium cobalt oxide, and electrons flow from the graphite to power your phone.

So your calls and texts depends on those lithium ions moving. Remember that, next time you use your mobile!

![Keeping in touch, via lithium ions.](image)
9.5 Reversible reactions

When you heat copper(II) sulfate crystals ...

The blue crystals above are hydrated copper(II) sulfate. On heating, they turn to a white powder. This is anhydrous copper(II) sulfate:

\[ \text{CuSO}_4\cdot 5\text{H}_2\text{O} (s) \rightarrow \text{CuSO}_4 (s) + 5\text{H}_2\text{O} (l) \]

The reaction is easy to reverse: add water! The anhydrous copper(II) sulfate gets hot and turns blue. The reaction is:

\[ \text{CuSO}_4 (s) + 5\text{H}_2\text{O} (l) \rightarrow \text{CuSO}_4\cdot 5\text{H}_2\text{O} (s) \]

So the reaction can go in either direction: it is reversible. The reaction we start with (1 above) is called the forward reaction. Reaction 2 is the back reaction.

We use the symbol \( \rightleftharpoons \) instead of a single arrow, to show that a reaction is reversible. So the equation for the reaction above is:

\[ \text{CuSO}_4\cdot 5\text{H}_2\text{O} (s) \rightleftharpoons \text{CuSO}_4 (s) + 5\text{H}_2\text{O} (l) \]

What about the energy change?

Reaction 1 above requires heat — it is endothermic. In 2, the white powder gets hot and spits when you drip water on it — so that reaction is exothermic. It gives out the same amount of heat as reaction 1 took in.

A reversible reaction is endothermic in one direction, and exothermic in the other. The same amount of energy is transferred each time.

Some important reversible reactions

Many important reactions are reversible. Here are some examples:

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Comments</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{N}_2 (g) + 3\text{H}_2 (g) \rightarrow 2\text{NH}_3 (g) )</td>
<td>This is a very important reaction, because ammonia is used to make nitric acid and fertilisers.</td>
</tr>
<tr>
<td>( 2\text{SO}_2 (g) + \text{O}_2 (g) \rightarrow 2\text{SO}_3 (g) )</td>
<td>This is a key step in the manufacture of sulfuric acid.</td>
</tr>
<tr>
<td>( \text{CaCO}_3 (s) \rightleftharpoons \text{CaO} (s) + \text{CO}_2 (g) )</td>
<td>This is a thermal decomposition: it needs heat. Calcium oxide (called lime, or quicklime) has many uses (page 240).</td>
</tr>
</tbody>
</table>

Water of crystallisation

- The water in blue copper(II) sulfate crystals is called water of crystallisation.
- Hydrated means it has water molecules built into its structure.
- Anhydrous means no water is present.

Two tests for water

Water will turn:
- white anhydrous copper(II) sulfate blue
- blue cobalt chloride paper pink.

Both compounds add on water of crystallisation, giving the colour change. To reverse, just heat!
Reversible reactions and equilibrium

As you saw in the last table, the reaction between nitrogen and hydrogen to make ammonia is **reversible**:

\[ \text{N}_2 (g) + 3\text{H}_2 (g) \rightleftharpoons \text{2NH}_3 (g) \]

So let’s see what happens during the reaction:

Three molecules of hydrogen react with one of nitrogen to form two of ammonia. So if you put the correct mixture of nitrogen and hydrogen into a closed container …

… will it all turn into ammonia? No! Once a certain amount of ammonia is formed, the system reaches a state of **dynamic equilibrium**. From then on …

… every time two ammonia molecules form, another two break down into nitrogen and hydrogen. So the level of ammonia remains unchanged.

**Equilibrium** means there is *no overall change*. The amounts of nitrogen, hydrogen and ammonia remain steady. But **dynamic** means there is continual change: ammonia molecules continually break down, while new ones form.

In a **closed system**, a reversible reaction reaches a state of **dynamic equilibrium**, where the forward and back reactions take place at the same rate. So there is *no overall change*.

The term **dynamic equilibrium** is usually shortened to **equilibrium**.

**A challenge for industry**

Imagine you run a factory that makes ammonia. You want the yield of ammonia to be as high as possible.

But the reaction between nitrogen and hydrogen is *never complete*. Once equilibrium is reached, a molecule of ammonia breaks down every time a new one forms.

This is a problem. What can you do to increase the yield of ammonia? You will find out in the next unit.

---

**Q**

1. What is a reversible reaction?
2. Write a word equation for the reaction between solid copper(II) sulfate and water.
3. How would you turn hydrated copper(II) sulfate into anhydrous copper(II) sulfate?
4. What will you observe if you place pink cobalt chloride paper in warm oven?
5. Explain the term **dynamic equilibrium**.
6. Nitrogen and hydrogen are mixed, to make ammonia.
   a. Soon, two reactions are going on in the mixture. Give the equations for them.
   b. For a time, the rate of the forward reaction is greater than the rate of the back reaction. Has equilibrium been reached? Explain.
9.6 Shifting the equilibrium

The challenge
Reversible reactions present a challenge to industry, because they never complete. Let’s look at that reaction between nitrogen and hydrogen again:

\[ \text{N}_2 (g) + 3\text{H}_2 (g) \rightleftharpoons 2\text{NH}_3 (g) \]

This represents the reaction mixture at equilibrium. The amount of ammonia in it will not increase … because every time a new molecule of ammonia forms, another breaks down.

What can be done?
You want as much ammonia as possible. So how can you increase the yield? This idea, called Le Chatelier’s principle, will help you:

When a reversible reaction is in equilibrium and you make a change, the system acts to oppose the change, and restore equilibrium. A new equilibrium mixture forms.

A reversible reaction always reaches equilibrium, in a closed system. But by changing conditions, you can shift equilibrium, so that the mixture contains more product. Let’s look at four changes you could make.

1. Change the temperature
Will raising the temperature help you obtain more ammonia? Let’s see.

\[ \text{N}_2 + 3\text{H}_2 \xrightarrow{\text{heat out}} 2\text{NH}_3 \]
\[ 2\text{NH}_3 \xrightarrow{\text{heat in}} \text{N}_2 + 3\text{H}_2 \]

The forward reaction is exothermic – it gives out heat. The back reaction is endothermic – it takes it in. Heating speeds up both reactions … but if you heat the equilibrium mixture, it acts to oppose the change. More ammonia breaks down in order to use up the heat you add.

What if you lower the temperature? The system acts to oppose the change: more ammonia forms, giving out heat. Great! But if the temperature is too low, the reaction takes too long to reach equilibrium. Time is money, in a factory. So it is best to choose a moderate temperature.
2 Change the pressure

Pressure is caused by the gas molecules colliding with the walls of the container. The more molecules present, the higher the pressure.

When you increase the pressure, the equilibrium mixture acts to oppose this. More ammonia forms, which means fewer molecules. So the amount of ammonia in the mixture has increased. Equilibrium has shifted to the right. Well done. You are on the right track.

3 Remove the ammonia

The equilibrium mixture is a balance between nitrogen, hydrogen, and ammonia. Suppose you cool the mixture. Ammonia condenses first, so you can run it off as a liquid. Then warm the remaining nitrogen and hydrogen again. More ammonia will form, to restore the balance.

Choosing the optimum conditions

So to get the best yield of ammonia, it is best to:

- use high pressure, and remove ammonia, to improve the yield
- use a moderate temperature, and a catalyst, to get a decent rate.

Page 227 shows how these ideas are applied in an ammonia factory.

A note about rate

By now, you should realize that:

- a change in temperature always shifts equilibrium.
- a change in pressure will shift equilibrium only if the number of molecules is different on each side of the equation.

But how do these changes affect the rate? Raising the temperature or pressure increases the rate of both the forward and back reactions, so equilibrium is reached faster. (A temperature rise gives the molecules more energy. An increase in pressure forces them closer. So in both cases, the number of successful collisions increases.)

1 The reaction between nitrogen and hydrogen is reversible. This causes a problem for the ammonia factory. Why?
2 What is Le Chatelier’s principle? Write it down.
3 In manufacturing ammonia, explain why:
   a high pressure is used   b ammonia is removed
4 Sulfur dioxide (SO₂) and oxygen react exothermically to form sulfur trioxide (SO₃). The reaction is reversible.
   a Write the symbol equation for this reaction.
   b What happens to the yield of sulfur trioxide if you:
      i increase the pressure?  ii raise the temperature?
Checkup on Chapter 9

Revision checklist

Core curriculum
Make sure you can …
- explain what these terms mean: exothermic reaction, endothermic reaction
- give examples of exothermic and endothermic reactions and draw energy level diagrams for them
- state the unit used for measuring energy
- say what the $+\text{ and } -$ signs mean, in energy values
- explain what the purpose of a fuel is
- name the fossil fuels, and say how we use them
- explain what nuclear fuels are, and where we use them, and name one
- give advantages and disadvantages of nuclear fuels
- say how hydrogen and ethanol are used as fuels
- explain what a reversible reaction is, with examples
- write the symbol for a reversible reaction
- describe how to change hydrated copper(II) sulfate to the anhydrous compound, and back again
- explain why anhydrous copper(II) sulfate can be used to test for the presence of water

Extended curriculum
Make sure you can also …
- use the idea of bond making and bond breaking to explain why a reaction is exo- or endothermic
- define bond energy
- calculate the energy change in a reaction, given the equation, and bond energy values
- describe a simple cell, and explain that the current comes from a redox reaction
- predict which metal will be the negative pole in a simple cell
- give half-equations for reactions that take place in a simple cell (like the one on page 120)
- describe the hydrogen fuel cell, and give the overall reaction that takes place in it
- explain that a reversible reaction never completes, in a closed container – it reaches equilibrium
- give ways to obtain more product in a gaseous reversible reaction
- predict the effect of a change in temperature and pressure, for a given reversible reaction
- say how a catalyst will affect a reversible reaction
- predict the effect of a change in conditions for a reversible reaction in solution

Questions

Core curriculum

1. Look at this reaction:
   \[ \text{NaOH (aq)} + \text{HCl (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)} \]
   a. Which type of reaction is it?
   b. It is exothermic. What does that mean?
   c. What will happen to the temperature of the solution, as the chemicals react?
   d. Draw an energy diagram for the reaction.

2. Water at 25°C was used to dissolve two compounds. The temperature of each solution was measured immediately afterwards.
<table>
<thead>
<tr>
<th>Compound</th>
<th>Temperature of solution/°C</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonium nitrate</td>
<td>21</td>
</tr>
<tr>
<td>calcium chloride</td>
<td>45</td>
</tr>
</tbody>
</table>
   a. List the apparatus needed for this experiment.
   b. Calculate the temperature change on dissolving each compound.
   c. i. Which compound dissolved exothermically?
        ii. How did you decide this?
        iii. What can you say about the energy level of its ions in the solution, compared with in the solid compound?
   d. For each solution, estimate the temperature of the solution if:
      i. the amount of water is halved, but the same mass of compound is used
      ii. the mass of the compound is halved, but the volume of water is unchanged
      iii. both the mass of the compound, and the volume of water, are halved.

3. Hydrated copper(II) sulfate crystals were heated:
   a. What is the ice for?
   b. What colour change will occur in the test-tube?
   c. The reaction is reversible. What does that mean?
   d. How would you show that it is reversible?
   e. Write the equation for the reversible reaction.
Extended curriculum

4 The fuel natural gas is mostly methane. Its combustion in oxygen is exothermic:
\[
CH_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O (l)
\]
(a) Explain why this reaction is exothermic, in terms of bond breaking and bond making.
(b) 
(i) Copy and complete this energy diagram for the reaction, indicating:
   A the overall energy change
   B the energy needed to break bonds
   C the energy given out by new bonds forming.

(ii) Methane will not burn in air until a spark or flame is applied. Why not?
(c) When 1 mole of methane burns in oxygen, the energy change is \(-890\) kJ.
   (i) What does the \(-\) sign tell you?
   (ii) Which word describes a reaction with this type of energy change?
(d) How much energy is given out when 1 gram of methane burns? \((A_r: \text{C} = 12, \text{H} = 1.)\)

5 Strips of copper foil and magnesium ribbon were cleaned with sandpaper and then connected as shown below. The bulb lit up.

(a) Why were the metals cleaned?
(b) Name the electrolyte used.
(c) Explain why the bulb lit up.
(d) Which metal releases electrons into the circuit?
(e) In this arrangement, energy is being changed from one form to another. Explain.
(f) What is this type of arrangement called?
(g) Give reasons why the set-up shown above would not be used as a torch battery.

6 The gas hydrazine, \(N_2H_4\), burns in oxygen like this:
\[
N-N \ (g) + O \ (g) \rightarrow N=N \ (g) + 2O \ (g)
\]
(a) Count and list the bonds broken in this reaction.
(b) Count and list the new bonds formed.
(c) Calculate the total energy:
   (i) required to break the bonds
   (ii) released when the new bonds form.
   (The bond energies in kJ/mole are: \(N-H\) 391; \(N-N\) 158; \(N\equiv N\) 945; \(O-H\) 464; \(O=O\) 498.)
(d) Calculate the energy change in the reaction.
(e) Is the reaction exothermic, or endothermic?
(f) Where is energy transferred from, and to?
(g) Comment on the suitability of hydrazine as a fuel.

7 Hydrogen and bromine react reversibly:
\[
H_2 (g) + Br_2 (g) \rightleftharpoons 2HBr (g)
\]
(a) Which of these will favour the formation of more hydrogen bromide?
   (i) add more hydrogen
   (ii) remove bromine
   (iii) remove the hydrogen bromide as it forms
(b) Explain why increasing the pressure will have no effect on the amount of product formed.
(c) However, the pressure is likely to be increased, when the reaction is carried out in industry. Suggest a reason for this.

8 Ammonia is made from nitrogen and hydrogen. The energy change in the reaction is \(-92\) kJ/mole. The reaction is reversible, and reaches equilibrium.
(a) Write the equation for the reaction.
(b) Is the forward reaction endothermic, or exothermic? Give your evidence.
(c) Explain why the yield of ammonia:
   (i) rises if you increase the pressure
   (ii) falls if you increase the temperature
(d) What effect does increasing:
   (i) the pressure
   (ii) the temperature have on the rate at which ammonia is made?
(e) Why is the reaction carried out at 450°C rather than at a lower temperature?

9 The dichromate ion \(Cr_2O_7^{2-}\) and chromate ion \(CrO_4^{2-}\) exist in equilibrium, like this:
\[
Cr_2O_7^{2-} (aq) + H_2O (l) \rightleftharpoons 2CrO_4^{2-} (aq) + 2H^+ (aq)
\]
orange \(\rightarrow\) yellow
(a) What would you see if you added dilute acid to a solution containing chromate ions?
(b) How would you reverse the change?
10.1 Rates of reaction

Fast and slow
Some reactions are fast and some are slow. Look at these examples:

The precipitation of silver chloride, when you mix solutions of silver nitrate and sodium chloride. This is a very fast reaction.

Concrete setting. This reaction is quite slow. It will take a couple of days for the concrete to fully harden.

Rust forming on an old car. This is usually a very slow reaction. It will take years for the car to rust completely away.

But it is not always enough to know just that a reaction is fast or slow. In factories where they make products from chemicals, they need to know exactly how fast a reaction is going, and how long it will take to complete. In other words, they need to know the rate of the reaction.

What is rate?
Rate is a measure of how fast or slow something is. Here are some examples.

This plane has just flown 800 kilometers in 1 hour. It flew at a rate of 800 km per hour.

This petrol pump can pump out petrol at a rate of 50 litres per minute.

This machine can print newspapers at a rate of 10 copies per second.

From these examples you can see that:

Rate is a measure of the change that happens in a single unit of time.

Any suitable unit of time can be used – a second, a minute, an hour, even a day.
When zinc is added to dilute sulfuric acid, they react together. The zinc disappears slowly, and a gas bubbles off.

As time goes by, the gas bubbles off more and more slowly. This is a sign that the reaction is slowing down. Finally, no more bubbles appear. The reaction is over, because all the acid has been used up. Some zinc remains behind.

The gas that bubbles off is hydrogen. The equation for the reaction is:

$$\text{zinc} + \text{sulfuric acid} \rightarrow \text{zinc sulfate} + \text{hydrogen}$$

$$\text{Zn (s)} + \text{H}_2\text{SO}_4 (aq) \rightarrow \text{ZnSO}_4 (aq) + \text{H}_2 (g)$$

Both zinc and sulfuric acid get used up in the reaction. At the same time, zinc sulfate and hydrogen form.

You could measure the rate of the reaction, by measuring:

- the amount of zinc used up per minute or
- the amount of sulfuric acid used up per minute or
- the amount of zinc sulfate produced per minute or
- the amount of hydrogen produced per minute.

For this reaction, it is easiest to measure the amount of hydrogen produced per minute, since it is the only gas that forms. It can be collected as it bubbles off, and its volume can be measured.

In general, to find the rate of a reaction, you should measure:

- the amount of a reactant used up per unit of time or
- the amount of a product produced per unit of time.

Q1
Here are some reactions that take place in the home. Put them in order of decreasing rate (the fastest one first).

- a raw egg changing to hard-boiled egg
- b fruit going rotten
- c cooking gas burning
- d bread baking
- e a metal tin rusting

Q2
Which of these rates of travel is slowest?

- 5 kilometres per second
- 20 kilometres per minute
- 60 kilometres per hour

Q3
Suppose you had to measure the rate at which zinc is used up in the reaction above. Which of these units would be suitable? Explain your choice.

- a litres per minute
- b grams per minute
- c centimetres per minute

Q4
Iron reacts with sulfuric acid like this:

$$\text{Fe (s)} + \text{H}_2\text{SO}_4 (aq) \rightarrow \text{FeSO}_4 (aq) + \text{H}_2 (g)$$

- a Write a word equation for this reaction.
- b Write down four different ways in which the rate of the reaction could be measured.
Measuring the rate of a reaction

A reaction that produces a gas

The rate of a reaction is found by measuring the amount of a reactant used up per unit of time, or the amount of a product produced per unit of time. Look at this reaction:

\[
\text{magnesium + hydrochloric acid} \rightarrow \text{magnesium chloride + hydrogen} \\
\text{Mg (s) + 2HCl (aq) $\rightarrow$ MgCl}_2 (aq) + \text{H}_2 (g)
\]

Here hydrogen is the easiest substance to measure, because it is the only gas in the reaction. It bubbles off and can be collected in a gas syringe, where its volume is measured.

The experiment

Clean the magnesium with sandpaper. Put dilute hydrochloric acid in the flask. Drop the magnesium into the flask, and insert the stopper and syringe immediately. Start the clock at the same time.

Hydrogen begins to bubble off. It rises up the flask and into the gas syringe, pushing the plunger out:

At the start, no gas has yet been produced or collected. So the plunger is all the way in. Now the plunger has been pushed out to the 20 cm\(^3\) mark. 20 cm\(^3\) of gas have been collected.

The volume of gas in the syringe is noted at intervals – for example every half a minute. How will you know when the reaction is complete?

Typical results

<table>
<thead>
<tr>
<th>Time / minutes</th>
<th>0</th>
<th>1/2</th>
<th>1</th>
<th>1 1/2</th>
<th>2</th>
<th>2 1/2</th>
<th>3</th>
<th>3 1/2</th>
<th>4</th>
<th>4 1/2</th>
<th>5</th>
<th>5 1/2</th>
<th>6</th>
<th>6 1/2</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume of hydrogen / cm(^3)</td>
<td>0</td>
<td>8</td>
<td>14</td>
<td>20</td>
<td>25</td>
<td>29</td>
<td>33</td>
<td>36</td>
<td>38</td>
<td>39</td>
<td>40</td>
<td>40</td>
<td>40</td>
<td>40</td>
</tr>
</tbody>
</table>

This table shows some typical results for the experiment. You can tell quite a lot from this table. For example, you can see that the reaction lasted about five minutes. But a graph of the results is even more helpful. The graph is shown on the next page.
Notice these things about the results:

1. In the first minute, 14 cm\(^3\) of hydrogen are produced.
   So the rate for the first minute is 14 cm\(^3\) of hydrogen per minute.
   In the second minute, only 11 cm\(^3\) are produced. \((25 - 14 = 11)\)
   So the rate for the second minute is 11 cm\(^3\) of hydrogen per minute.
   The rate for the third minute is 8 cm\(^3\) of hydrogen per minute.
   So the rate decreases as time goes on.
   The rate changes all through the reaction. It is greatest at the start, but decreases as the reaction proceeds.

2. The reaction is fastest in the first minute, and the curve is steepest then. It gets less steep as the reaction gets slower.
   The faster the reaction, the steeper the curve.

3. After 5 minutes, no more hydrogen is produced, so the volume no longer changes. The reaction is over, and the curve goes flat.
   When the reaction is over, the curve goes flat.

4. Altogether, 40 cm\(^3\) of hydrogen are produced in 5 minutes.
   The average rate for the reaction = \(\frac{\text{total volume of hydrogen}}{\text{total time for the reaction}}\)
   = \(\frac{40 \text{ cm}^3}{5 \text{ minutes}}\)
   = 8 cm\(^3\) of hydrogen per minute.

Note that this method can be used for any reaction where one product is a gas.

Q

1. For the experiment in this unit, explain why:
   a. the magnesium ribbon is cleaned first
   b. the clock is started the moment the reactants are mixed
   c. the stopper is replaced immediately

2. From the graph at the top of this page, how can you tell when the reaction is over?

3. Look again at the graph at the top of the page.
   a. How much hydrogen is produced in the first:
      i. 2.5 minutes?
      ii. 4.5 minutes?
   b. How long did it take to get 20 cm\(^3\) of hydrogen?
   c. What is the rate of the reaction during:
      i. the fourth minute?
      ii. the sixth minute?