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J Giffney and I Torrie
Chemistry 1.3

AS 90171 Describe chemical reactions 4 Credits

Achievement Criteria

<table>
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<th>Achievement</th>
<th>Achievement with Merit</th>
<th>Achievement with Excellence</th>
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<td>• Describe chemical reactions.</td>
<td>• Interpret information about chemical reactions.</td>
<td>• Apply understanding of chemical reactions.</td>
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Explanatory Notes

1 Chemical reactions will be selected from:
   - Oxidation-reduction reactions. These are limited to: simple electron transfer involving elements and monatomic ions (such as Cl₂/Cl⁻, I₂/I⁻, Fe³+/Fe²⁺, metal/metal ion), simple oxygen transfer (such as between metal oxides and either hydrogen or carbon).
   - Precipitation reactions. These are limited to: formation of chlorides of silver and lead; sulfates of calcium, barium and lead, hydroxides and carbonates of copper(II), iron(II), iron(III), zinc, aluminium, calcium, and magnesium ions.
   - Thermal decomposition reactions. These are limited to: hydroxides, carbonates and hydrogen carbonates.

2 Describe, interpret information, and apply understanding of chemical reactions will include the carrying out of calculations.

3 Description and interpretation of chemical reactions may involve:
   - classification of reactions (as given in Explanatory Note 1)
   - relating observations to chemical reactions
   - writing word equations
   - writing formulae of reactants and/or products
   - predicting the formation of a precipitate when two solutions are mixed.

4 Application of understanding of chemical reactions may involve:
   - explanation of observations in terms of the reactants and products involved
   - justification of the classification of reactions using equations and/or observations.

5 For Achievement, calculations will be straightforward and involve:
   - balancing simple equations in which the formulae of reactants and products are given
   - calculating the molar mass of a substance when the formula and molar mass values of the elements are given.

6 For Achievement with Merit, calculations involve mass-mass calculations in which:
   - a given mass of a reactant or product is used to determine the mass of another substance in the same reaction. The balanced equation for the reaction will be given, and the known: unknown mole ratio will be 1:1
   - writing balanced equations from given named reactants or products.

7 For Achievement with Excellence, calculations will be complex and will involve mass-mass calculations in which:
   - a given mass of a reactant or product is used to determine the mass of another substance in the same reaction. The balanced equation for the reaction will be given, and the known: unknown mole ratio will be other than 1:1
   - the formula of a compound is determined.

8 The states of substances will be indicated in the question format, but are not required in student responses.

9 Solubility rules and a table of ions will be provided.

10 A periodic table showing symbols, atomic numbers and atomic masses only will be provided.
When is a change not a change?
A chemical reaction is just a reshuffling of atoms, where the bonds which join the atoms have been broken and rejoined in new ways. This chemical change results in new substances (the “products”) which usually have completely different properties (both physical and chemical) to the original reactants.

Just because different substances are mixed together, does not necessarily mean that a chemical reaction will take place. Sometimes the result is just a mixture in which the different chemicals are all jumbled up together but each component behaves exactly the same as before they were mixed. For example, iron and sulfur powders can be separated from a mixture using a magnet because the iron is still magnetic and the sulfur is not. In contrast, if a mixture of iron and sulfur is heated so that a chemical reaction occurs, the new product, (iron sulfide), is no longer magnetic.

Substances can also undergo a different type of change called physical change.

Exercise: Complete the diagram below which summarises common physical changes.

Mixing is also a physical change (where no change of individual properties occur) as in a salad dressing made from oil and vinegar or sand mixed with salt. Physical changes can often be fairly easily reversed (often by adding or removing energy) whereas chemical changes are often permanent or very difficult to reverse eg cement + sand + water → concrete!

Characteristics of Chemical change
Exercise: Complete the bubble map summarising ways of recognising chemical change.
Exercise: Tick each of the following situations where a chemical reaction occurs.

- water boiling
- a candle burning
- curtains fading
- baking bread
- eggs being beaten
- oil spilling on water
- a distress flare going off
- mowing the lawn
- lighting a gas BBQ
- adding limestone to some acid split on the floor
- fish marinating in lemon juice
- snow falling
- adding salt to french fries
- digesting a piece of toast
- switching on an electric light
- chopping wood for a fire
- photosynthesis in a leaf
- cutting out a magazine article
- producing iron from iron sand

Conservation of Mass

Because atoms are neither created nor destroyed in a chemical reaction, but are simply rearranged, this means that mass must be conserved.

\[
\text{total mass of products} = \text{total mass of reactants.}
\]

In some reactions the mass may appear to decrease because a product escapes as a gas (e.g., CO}_2\text{ or H}_2\text{O}) and so is no longer included in the weighing (unless it is collected). The mass may also appear to increase (e.g., the rusting of iron) if one of the reactants is a gas which was not included in the original weighing.

BALANCING EQUATIONS

Formulae show chemistry at a standstill whereas equations show chemistry in action.

Equations show:
1. The reactants which enter into a reaction.
2. The products which are formed by the reaction.
3. The amounts of each substance used and each substance produced.

Two important principles are that:

1. Every chemical compound has a formula which cannot be altered.
2. A chemical reaction must account for every atom that is used.

The 2\textsuperscript{nd} principle arises from the Law of Conservation of Matter which states that in a chemical reaction atoms are neither created nor destroyed.

Some other things to remember are:

1. Many elements exist as pairs of atoms e.g., H}_2, N}_2, O}_2, F}_2, Cl}_2, Br}_2, I}_2.
2. The formulae of ionic compounds are obtained from its ions using the “crossover” or “swap and drop” method to balance the charge.
3. The formulae of acids is obtained in a similar manner by matching the correct number of H}_+ ions with the charge of the parent ion of the acid family. e.g., sulfuric acid, H}_2SO}_4 from the sulfate ion, SO}_4^{2-}.
4. Before beginning to balance the equation each formula must be written correctly and these are NEVER changed during balancing.

Balancing is done by placing coefficients (numbers) in front of each formulae to ensure the same number of atoms of each element on both sides of the arrow.

**NOTE**

Where the number is 1, it is omitted.

1. Fractional coefficients such as 1/2 can be used although it is usual to convert these to integers at the end of the balancing process by multiplying all coefficients by the same factor.
HINT: Use the following order to balance the equation:
1. Metallic elements
2. Polyatomic ions that are **exactly the same** on both sides of the equation.
3. Non-metallic elements that are **uncombined** on one side of the equation.
4. Remaining elements other than hydrogen or oxygen.
5. Hydrogen.

Exercise For each of the following equations:

(a) Use Molymods to construct models of each of the reactants and products.
(b) Make duplicates where necessary until the number of each type of atom is the same for both reactants and products i.e. it is “balanced”.
(c) Draw simple diagrams to represent your models before (reactants) and after (products) the reaction.
(d) Write the correct numbers (this may be 1) in front of each reactant and product in the equation.

1. \( _{ } H_2 + _{ } O_2 \rightarrow _{ } H_2O \)

2. \( _{ } Mg + _{ } O_2 \rightarrow _{ } MgO \)

3. \( _{ } H_2 + _{ } Cl_2 \rightarrow _{ } HCl \)

4. \( _{ } N_2 + _{ } H_2 \rightarrow _{ } NH_3 \)

5. \( _{ } CH_4 + _{ } O_2 \rightarrow _{ } CO_2 + _{ } H_2O \)

6. \( _{ } Na + _{ } H_2O \rightarrow _{ } NaOH + _{ } H_2 \)
**Mass-mass calculations**

When hydrogen burns in oxygen, water is produced by the balanced equation:

\[ 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

This means that we need **two** hydrogen molecules for every **one** oxygen molecule. However, it does NOT mean that we need 2g of hydrogen for every 1g of oxygen because different atoms (and molecules) weigh different amounts. In fact each oxygen atom weighs 16x more than each hydrogen atom as it has a total of 16 protons and neutrons compared to hydrogen’s 1 proton. Similarly carbon atoms weigh 12x more than a hydrogen atom.

Masses of every element are measured on a “relative atomic mass” scale.

<table>
<thead>
<tr>
<th>Element</th>
<th>Relative Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>1</td>
</tr>
<tr>
<td>C</td>
<td>12</td>
</tr>
<tr>
<td>N</td>
<td>14</td>
</tr>
<tr>
<td>O</td>
<td>16</td>
</tr>
<tr>
<td>Na</td>
<td>23</td>
</tr>
<tr>
<td>Mg</td>
<td>24</td>
</tr>
<tr>
<td>S</td>
<td>32</td>
</tr>
<tr>
<td>Cl</td>
<td>35.5</td>
</tr>
<tr>
<td>Cu</td>
<td>63.5</td>
</tr>
</tbody>
</table>

Sometimes the values are not exact whole numbers (e.g. Cl = 35.5) because the value is an average of the different isotopes that exist in nature for that particular element. (e.g. chlorine consists of 75% $^{35}\text{Cl}$ and 25% $^{37}\text{Cl}$ with the latter having two extra neutrons in the nucleus resulting in a higher mass).

The mass of a molecule is obtained by adding up the individual relative atomic masses.

\[ \text{H}_2\text{O} = (2\times1) + (1\times16) = 18 \]

**Exercise**

Use the values given in the box above to calculate the relative molar masses for each of the following substances.

(a) $\text{H}_2$

(b) $\text{CO}_2$

(c) $\text{H}_2\text{SO}_4$

(d) $\text{O}_2$

(e) $\text{NaOH}$

(f) $\text{MgCl}_2$

(g) $\text{NH}_3$

(h) $(\text{NH}_4)_2\text{SO}_4$

(i) $\text{CuSO}_4$

(j) $\text{HCl}$

Example

In our original example of $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

- the total initial mass $= (2\times2) + (1\times16) = 36$
- and the final total mass $= (2\times18) = 36$

i.e mass is conserved (provided that the equation is balanced)

**NOTE** If the equation was not balanced correctly (e.g. $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$) then the masses would not be equal.

**Exercise**

Balance each of the following equations (many numbers may be 1) and then use your answers from Task A to show that the masses are the same on both sides of the equation.

(a) $\_\text{CuO} + \_\text{H}_2\text{SO}_4 \rightarrow \_\text{CuSO}_4 + \_\text{H}_2\text{O}$

(b) $\_\text{NH}_3 + \_\text{H}_2\text{SO}_4 \rightarrow \_ (\text{NH}_4)_2\text{SO}_4$

(c) $\_\text{MgCO}_3 + \_\text{HCl} \rightarrow \_\text{MgCl}_2 + \_\text{CO}_2 + \_\text{H}_2\text{O}$
We have shown previously that a mass of 4g of hydrogen gas reacts with 32g of oxygen gas to produce a total of 36g of water.
If instead of using 4g of hydrogen we used 40g, we would have also needed to use 10x as much oxygen (320g).
How much water would you expect to be produced? _________________

Does this still obey the law of conservation of mass? _________________

What would the masses be if we had only started with 1g of hydrogen gas?
Mass of oxygen required = _______ g
Mass of water produced =_______ g

Exercise
Use your answers from the previous exercise to calculate the missing masses for each of the following equations:

(a)  \[
\text{CuO} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{H}_2\text{O}
\]
\[
?g + 9.8g \rightarrow ?g + ?g
\]

(b)  \[
\text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4
\]
\[
?g + ?g \rightarrow 264g
\]

(c)  \[
\text{MgCO}_3 + \text{HCl} \rightarrow \text{MgCl}_2 + \text{CO}_2 + \text{H}_2\text{O}
\]
\[
42g + ?g \rightarrow ?g + ?g + ?g
\]

Types of Chemical reactions
Although there are millions of possible chemical reactions they can be classified into general types of reactions – much as the different groups in the Periodic Table summarise the properties of similar elements within the same “family”.

This achievement standard covers three types of reactions called oxidation-reduction (“redox”), precipitation and thermal decomposition reactions respectively.

(NOTE Another type of reaction called neutralisation or acid-base reactions is covered in C1.4)

Precipitation reactions
A solution is a mixture made up of a solute (usually a solid) dissolved in a solvent (usually a liquid).
Soluble substances are those which dissolve in water to form clear solutions.
Note that clear is not a colour – it simply means that you can see through it.
If a solution has no colour it should be described as colourless.

Many ionic compounds are soluble in water because the attractions between the water molecules and the ions are greater than the attractions between the positive and negative ions themselves.
Once the oppositely charged ions have been separated, these individual ions are far too small to be seen so the solution becomes transparent. The dissolving process can be written as
\[
\text{NaCl(s)} \xrightleftharpoons{H_2O} \text{Na}^+(aq) + \text{Cl}^-(aq)
\]
The (s) stands for solid and the (aq) stands for dissolved.
Note that as the water is unchanged it does not appear as a reactant but is shown above the arrow to indicate its presence is necessary.
If the water in a solution is allowed to evaporate, eventually there is not enough water present to keep the oppositely charged ions apart and they start to clump together to form ionic solids.
This process is called crystallisation.

Solubility
The solubility of a substance is the maximum amount that will dissolve in a given volume of water to form a “saturated” solution. If more solid is added to a saturated solution it will not dissolve (provided the temperature is not changed).
The solubility of an ionic compound in water depends on three factors:
- the specific ionic compound (is it sodium nitrate or silver chloride)
- the amount of water present (more “salt” will dissolve in a bucket than in a cup full of water).
- the temperature of the water (usually substances are more soluble in hot water)

**Insolubility**
Ionic compounds which do not easily dissolve in water, are called insoluble. This occurs because the attraction between the oppositely charged ions in the crystal is too strong to be overcome by the attraction between the water molecules and the ions.

Whether an ionic compound is soluble or insoluble can be determined by adding a tiny amount of the solid to half a test tube of water, shaking it and waiting to see if it dissolves to form a clear solution or whether the solid forms a cloudy suspension which eventually sinks to the bottom eg if sodium carbonate is used the compound dissolves because it is soluble, whereas if calcium carbonate is used a cloudy white mixture forms because it is insoluble.

**Solubility rules**
The solubility of different ionic compounds can be summarised by the following solubility “rules”, which are based on the results of experimental investigation.

**NOTE** These rules will be provided in all assessment tasks.

<table>
<thead>
<tr>
<th>Soluble</th>
<th>Insoluble</th>
</tr>
</thead>
<tbody>
<tr>
<td>(These compounds never precipitate)</td>
<td>(These compounds form precipitates)</td>
</tr>
<tr>
<td>1. All group 1 (ie Na⁺, K⁺) &amp; ammonium (NH₄⁺) compounds eg Na₂CO₃, NH₄Cl</td>
<td></td>
</tr>
<tr>
<td>2. All nitrate (NO₃⁻) compounds eg Pb(NO₃)₂, AgNO₃</td>
<td></td>
</tr>
<tr>
<td>3. Most sulfates (SO₄²⁻) except .............. CaSO₄, BaSO₄, PbSO₄, Ag₂SO₄</td>
<td></td>
</tr>
<tr>
<td>4. Most chlorides (Cl⁻) except .............. AgCl, PbCl₂</td>
<td></td>
</tr>
<tr>
<td>5. All carbonates (CO₃²⁻) except those covered by rule 1.</td>
<td></td>
</tr>
<tr>
<td>6. All hydroxides (OH⁻) except those covered by rule 1.</td>
<td></td>
</tr>
</tbody>
</table>

**Exercise**
Tick any of the following ionic compounds which are soluble (use the solubility rules above).

- **sodium chloride**
- **calcium carbonate**
- **copper(II) hydroxide**
- **potassium sulfate**
- **ammonium carbonate**
- **magnesium nitrate**
- **silver carbonate**
- **iron(III) hydroxide**
- **lead sulfate**
- **barium nitrate**
- **magnesium chloride**
- **lead(II) chloride**
**Precipitation**

Sometimes, when two (soluble) solutions are mixed together, the mixture goes cloudy, or a solid settles on the bottom of the container. This new insoluble solid is called a *precipitate*.

When the two solutions are mixed there are two new combination of ions possible: \( \text{A}^+ \text{Y}^- \) and \( \text{B}^+ \text{X}^- \). (Note that \( \text{A}^+ \text{B}^+ \) and \( \text{X}^- \text{Y}^- \) are not possible as like charges repel).

If either (or both) of these new combinations are insoluble the water will not be able to keep the ions apart and they will join together to form an ionic crystal which will settle out as a visible solid, e.g. \( \text{A}^+(aq) + \text{Y}^-(aq) \rightarrow \text{A}\text{Y}(s) \)

Note that if the insoluble product \( \text{AY} \) is filtered off, the filtrate will contain a solution of the soluble compound \( \text{BX} \), containing the ions \( \text{B}^+(aq) \) and \( \text{X}^-(aq) \).

These can be recovered as a solid by evaporating the water still present.

To decide whether or not a precipitate forms when two solutions are mixed, start by identifying the two new possible products that could form and use the solubility rules to decide whether they are both soluble or not. If either one is insoluble then a precipitate will occur.

**Exercise**

Use the solubility rules previously given, to predict if the following pairs of solutions will produce a precipitate when mixed.

In all cases where this occurs, name the precipitate formed. e.g. if silver nitrate and copper chloride are mixed the two possible products would be silver chloride and copper nitrate. The solubility rules tell us that AgCl is insoluble and so will form a precipitate and the copper nitrate will stay in solution.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Precipitate (✓ or ✗)</th>
<th>Name of precipitate</th>
</tr>
</thead>
<tbody>
<tr>
<td>calcium chloride + sodium sulfate</td>
<td>✗</td>
<td></td>
</tr>
<tr>
<td>aluminium nitrate + sodium hydroxide</td>
<td>✗</td>
<td></td>
</tr>
<tr>
<td>sodium sulfate + potassium carbonate</td>
<td>✗</td>
<td></td>
</tr>
<tr>
<td>silver nitrate + magnesium sulfate</td>
<td>✗</td>
<td></td>
</tr>
<tr>
<td>barium chloride + calcium nitrate</td>
<td>✗</td>
<td></td>
</tr>
<tr>
<td>potassium carbonate + zinc sulfate</td>
<td>✗</td>
<td></td>
</tr>
<tr>
<td>iron(III) chloride + sodium carbonate</td>
<td>✗</td>
<td></td>
</tr>
</tbody>
</table>

**Ionic equations**

Equations to represent precipitation reactions can be written in three different ways.

- **Word equation:** silver nitrate + sodium chloride \( \rightarrow \) silver chloride + sodium nitrate
- **Symbol equation:** \( \text{AgNO}_3(aq) + \text{NaCl}(aq) \rightarrow \text{AgCl(s) + NaNO}_3(aq) \)
- **Ionic equation:** \( \text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl(s)} \)

**NOTE** that the second version is unacceptable unless the states are included.

In the last version spectator ions are excluded from the expanded symbol equation, i.e., \( \text{Ag}^+ + \text{NO}_3^- + \text{Na}^+ + \text{Cl}^- \rightarrow \text{Ag}^+ + \text{Cl}^- + \text{Na}^+ + \text{NO}_3^- \)

The ionic equation can also be easily written by using the solubility rules to identify the precipitate formed and writing down the two ions that it is made from.

Note the balancing required in examples such as \( \text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{PbCl}_2(s) \)
Exercise
For each of the reactions on the previous page that produced a precipitate write down the ionic equation.

\[ \text{____________} + \text{____________} \rightarrow \text{____________} \]

\[ \text{____________} + \text{____________} \rightarrow \text{____________} \]

\[ \text{____________} + \text{____________} \rightarrow \text{____________} \]

\[ \text{____________} + \text{____________} \rightarrow \text{____________} \]

\[ \text{____________} + \text{____________} \rightarrow \text{____________} \]

Identifying a solution
The formation of a precipitate and its colour, can be used to identify the ions present in a solution. A flow chart, or key, such as given below can be used as it provides logical sequence of steps to follow, based on observable solubility properties. e.g. The version shown below will identify the anion (negative ion) present if it is either a chloride, sulfate, carbonate, hydroxide or nitrate.

Exercise
Use the anion flow chart to determine the anion present in a solution if it gives the following results:
- no reaction with nitric acid
- no precipitate with barium nitrate
- white precipitate with silver nitrate

Anion present ___________________
Oxidation – reduction reactions

Oxidation-reduction in terms of oxygen transfer
In its simplest sense oxidation occurs when an element or compound gains oxygen.

eg magnesium + oxygen → magnesium oxide
hydrogen + oxygen → water
sulfur + oxygen → sulfur dioxide

Conversely, reduction occurs when a compound loses oxygen.
It is important to realise that oxidation and reduction always occur together. e.g. if something gains oxygen, something else must have lost oxygen.
For example, when CO reduces Fe₂O₃ to Fe by removing its oxygen, the CO is oxidised to CO₂.

These reactions are often called “redox” reactions which is short for reduction-oxidation.

The substance which donates its oxygen is called the oxidising agent (or oxidant).
The substance which receives oxygen is called the reducing agent (or reductant).

An example would be the reaction used to produce iron from the iron(III) oxide present in iron sand

Fe₂O₃ + 3CO → 2Fe + 3CO₂

where the CO is the reducing agent (as it removes oxygen from the Fe₂O₃) and the Fe₂O₃ is the oxidising agent as it gives oxygen to the CO.

NOTE: In all cases, the reducing agent is itself oxidised and vice versa.

Exercises: For the reactions below, identify each reactant as either the oxidant or the reductant and complete the process boxes with the words oxidised and reduced as appropriate.

1. Mg + H₂O → MgO + H₂
   is _______ to
   is _______ to

2. CuO + H₂ → Cu + H₂O
   is _______ to
   is _______ to

3. PbO₂ + C → Pb + CO₂
   is _______ to
**Oxidation-reduction in terms of electron transfer**

When an iron nail is dipped into a solution of copper sulfate the red coating that forms on the nail is not “rust” (iron(III) oxide) but is actually metallic copper.

The overall reaction is \[ \text{Fe(s)} + \text{CuSO}_4(aq) \rightarrow \text{FeSO}_4(aq) + \text{Cu(s)}. \]

This is an example of a redox reaction although it is not obvious why this is. The reaction is actually between the Fe atom and Cu\(^{2+}\) ion and the SO\(_4^{2-}\) ion is not actually involved at all.

Ions which are not changed during the reaction are called **spectator ions** and can be omitted from the balanced equation for the reaction. We could have written the equation above as

\[ \text{Fe(s)} + (\text{Cu}^{2+} + \text{SO}_4^{2-})(aq) \rightarrow (\text{Fe}^{2+} + \text{SO}_4^{2-})(aq) + \text{Cu(s)} \]

which can be simplified to the ionic equation \[ \text{Fe(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Cu(s)} + \text{Fe}^{2+}(aq) \]

This means we would get exactly the same reaction whether we started with copper sulfate, copper chloride or copper nitrate as the SO\(_4^{2-}\), Cl\(^{-}\) and NO\(_3^{-}\) ions respectively, are not involved in the chemical change at all as they are all spectator ions.

In the reaction above, the Fe atom loses its two electrons (to form the Fe\(^{2+}\) ion) and the Cu\(^{2+}\) ion gains two electrons to form the Cu atom.

Why is this called redox when no oxygen has been transferred?

When magnesium reacts with oxygen to form magnesium oxide (or Mg\(^{2+}\) O\(^{2-}\)), the Mg atoms lose 2 electrons to form the Mg\(^{2+}\) ion and each O atom gains 2 electrons to form the O\(^{2-}\) ion.

\[ \text{Mg} \rightarrow \text{Mg}^{2+} + 2e^- \quad \text{and} \quad \text{O} + 2e^- \rightarrow \text{O}^{2-} \]

Notice that when the two “half equations” above are added the electrons cancel out to give \[ \text{Mg} + \text{O} \rightarrow \text{Mg}^{2+} + \text{O}^{2-} \] (or MgO).

However the equations written above would not be acceptable because the element oxygen actually exists as a pair of atoms (O\(_2\)) so we have to double the overall equation (Mg + O → MgO) to get the final balanced equation \[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}. \]

**NOTE:** We will look at the difference between the 2 in front of the Mg and the small subscript 2 after the O later, when we learn how to balance equations.

This means that oxidation can also be defined as **the loss of electrons in a reaction**.

Conversely reduction occurs when a substance gains electrons during a chemical reaction.

Some useful mnemonics (memory aids) for this are:

- LEO: the lion goes
- GERO: gain of electrons is reduction
- OXILoss

**NB:** Oxidation is gain of oxygen but loss of electrons.
Balancing redox reactions

When electrons are transferred from one substance to another there can’t be any “unused”
electrons left over. (This is the same principle that applies when we use the crossover method to
determine the formula of an ionic compound).

Consider the reaction between sodium metal and chlorine gas to produce the salt, sodium chloride.
Like most non-metals, chlorine exists as a diatomic molecule, Cl₂).

\[ \text{Na} + \text{Cl}_2 \rightarrow \text{NaCl} \quad \text{(or} \quad \text{Na}^+ + \text{Cl}^-) \]

The oxidation half equation is \( \text{Na} \rightarrow \text{Na}^+ + \text{e}^- \) (the Na atom has lost \( 1\text{e}^- \))
The reduction half equation is \( \text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^- \) (each Cl atom gains \( 1\text{e}^- \))

When writing these half equations the total charge on each side of the equation must be balanced.
e.g. for the oxidation half equation above there is no charge on the left side so the charge on the
right side must also be zero which it is. Note that it wouldn’t be zero if the electron wasn’t there or if
we had shown two electrons rather than one. Similarly when the Cl₂ is reduced it forms \( 2\text{Cl}^- \) ions
so 2 electrons must be added to the left hand side so that the charge on both sides is –2.
Finally, before we can combine the two equations together, the number of electrons lost must
equal the number of electrons gained so we need to double the first half equation so that both
involve two electrons.

\[ 2\text{Na} \rightarrow 2\text{Na}^+ + 2\text{e}^- \]

Now when we add the two half equations together we get

\[ 2\text{Na} + \text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Na}^+ + 2\text{Cl}^- + 2\text{e}^- \quad \text{or} \quad 2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl} \]

**NOTE**  Electrons don’t appear in the final balanced equation as they cancel out.

**Exercises:**  Complete and balance each of the following redox equations.

1. For the reaction \( \text{Fe} + \text{Cu}^{2+} \rightarrow \text{Fe}^{2+} + \text{Cu} \)

   oxidation half equation  ___________________

   reduction half equation  ___________________

   overall balanced redox equation  _____________________________

2. For the reaction \( \text{Mg} + \text{H}^+ \rightarrow \text{Mg}^{2+} + \text{H}_2 \)

   oxidation half equation  ___________________

   reduction half equation  ___________________

   overall balanced redox equation  _____________________________

3. For the reaction \( \text{Li} + \text{F}_2 \rightarrow \text{Li}^+ + \text{F}^- \)

   oxidation half equation  ___________________

   reduction half equation  ___________________

   overall balanced redox equation  _____________________________

4. For the reaction \( \text{Fe}^{2+} + \text{Cl}_2 \rightarrow \text{Fe}^{3+} + \text{Cl}^- \)

   oxidation half equation  ___________________

   reduction half equation  ___________________

   overall balanced redox equation  _____________________________

5. For the reaction \( \text{Cl}_2 + \text{Br}^- \rightarrow \text{Cl}^- + \text{Br}_2 \)

   oxidation half equation  ___________________

   reduction half equation  ___________________

   overall balanced redox equation  _____________________________
Competition between metals
Metal ions often react with a compound containing ions of another metal, e.g. when Fe reacts with Cu^{2+} ions, the iron metal is trying to get rid of its two valence electrons. To achieve this, the Cu^{2+} ions have to accept these two electrons and revert back to the elemental Cu form. However, like all metals (which lose electrons), copper is also stable in the form of the Cu^{2+} ion. In this situation, because iron is more “reactive” than copper, (i.e., it has a stronger tendency to lose its electrons) it is able to “force” the Cu^{2+} to accept the two electrons. There is a “tug of war” going on with Fe → Fe^{2+} + 2e^{-} being stronger than Cu → Cu^{2+} + 2e^{-} so overall Fe + Cu^{2+} → Fe^{2+} + Cu is the reaction that occurs. Whether this type of reaction proceeds or not can be predicted by using the activity series which is a list of all metals in order of their reactivity or desire to lose electrons, as shown below.

Activity series
A limited form of the metal activity series is shown below. It can be learnt using the mnemonic shown, but it will be provided in the assessment task.

| Na | > | Ca | > | Mg | > | Al* | > | Zn | > | Fe | > | Pb | > | [H2] | > | Cu | > | Ag | > | Au |
| React less vigorously with oxygen | React slowly | Don’t react |
| React with cold water | React with hot water or steam | React with acids only | Do not react at all

Note that as expected the most reactive metals are in Group 1 followed by Group 2 metals.

*Aluminium often does not appear to be as reactive as its position in the activity series suggests e.g. it is used extensively in buildings as windows and roofing without the need to be protected from air or water. This is because the metal is protected by a surface coating of an inert (unreactive) layer of aluminium oxide, which protects the aluminium metal below from further reaction.

In general a metal ion can be converted back into its metallic form by any metal above it in the activity series i.e. the most reactive metal always ends up in the ionic form as it loses its electrons most easily. e.g. Mg reacts with the Fe^{2+} present in a solution of Fe(NO_3)_2 to give Mg^{2+} + Fe whereas there is no reaction between Cu + Zn(NO_3)_2 as Cu is less reactive than Fe.

Note that the second example is simply the reverse of the successful reaction previously considered between Zn and Cu^{2+}.

Exercises
1. For each of the following reaction write a balanced equation. If no reaction happens, write “no reaction”.
   (a) Zn and Cu^{2+} ______________________________________
   (b) Ag and Cu^{2+} ______________________________________
   (c) Mg and Ag^{+} ______________________________________
   (d) Pb and 2H^{+} ______________________________________

2. Explain why it would not be advisable to store a solution of the fungicide, CuSO_4, in an iron container.
   ______________________________________
   ______________________________________
   ______________________________________
Galvanising
Unless iron is painted (to stop reaction with air and water) it will “corrode” to form rust. An alternative method of protecting iron, is to coat the metal with a thin layer of zinc to form “galvanised iron”. The zinc (which has its own protective, inert oxide layer like aluminium) prevents oxygen and water from reaching the iron below. If however, the zinc layer is scratched, or worn away, the zinc still protects the iron because it corrodes in preference to the iron as it is more reactive. This is called “sacrificial corrosion”.

What happens is that the reaction $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$ reaction occurs more easily than the reaction $\text{Fe} \rightarrow \text{Fe}^{2+} + 2e^-$ so the iron will not corrode until all of the zinc is used up.

For a similar reason, small blocks of zinc are often attached to the hulls of boats to help prevent the steel (an alloy of iron and carbon) corroding away in the wet, salt conditions. As the zinc corrodes away (in preference to the iron) the blocks are regularly replaced.

Reduction of metal ores
Apart from unreactive examples such as silver and gold, most metals exist in nature combined with non-elements to form compounds which are known as metal ores. The black sands on our West coast beaches contain a mixture of iron oxides. To extract the metallic iron from these compounds requires us to use an oxidising agent such as carbon (or carbon monoxide) to remove the oxygen as shown in the equation

$$\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$$

Decomposition reactions
Many ionic compounds, particularly metal carbonates, hydrogen carbonates and hydroxides (and some oxides) break apart when heated strongly. These are examples of decomposition reactions. This course is restricted to the first three examples described above, each of which follow the same general patterns described below.

1. metal carbonate $\xrightarrow{\text{heat}}$ metal oxide + carbon dioxide

$$\text{CuCO}_3 \rightarrow \text{CuO} + \text{CO}_2$$
\[\text{I. solid green black colourless}\]

$$\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$$
\[\text{I. solid white white colourless}\]

2. metal bicarbonate $\xrightarrow{\text{heat}}$ metal carbonate + carbon dioxide + water

$$2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$$
\[\text{I. white solid white solid colourless}\]

3. metal hydroxide $\xrightarrow{\text{heat}}$ metal oxide + water

$$\text{Cu(OH)}_2 \rightarrow \text{CuO} + \text{H}_2\text{O}$$
\[\text{I. blue solid black solid}\]

NB: In all of the above examples the gas produced can be confirmed as carbon dioxide by bubbling it through lime-water, which will initially turn milky because of the reaction.

$$\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$$
Types of reaction summary
The type of reaction can be determined either from observations or the equation itself.

In a precipitation reaction a cloudiness (or solid) appears when two solutions are mixed if the resulting ionic solid is insoluble.

In a redox reaction, the equation shows that oxygen is swapped between the reactants the reactants change from elements (eg metals) to ions (within compounds) which carry a OR charge.

In a thermal decomposition reaction the reactants are heated and at least two products form with either CO₂ or H₂O (or both) given off as gases leaving a new solid remaining.

Exercises:
Classify the following reactions as either precipitation, (P), oxidation-reduction, (R), or thermal decomposition, (D).

<table>
<thead>
<tr>
<th>Reaction equation</th>
<th>Type</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg(s) + H₂SO₄(aq) → MgSO₄(aq) + H₂(g)</td>
<td></td>
</tr>
<tr>
<td>Ca(HCO₃)₂(s) → CaCO₃(s) + CO₂(g) + H₂O(g)</td>
<td></td>
</tr>
<tr>
<td>CuSO₄(aq) + Mg → MgSO₄(aq) + Cu(s)</td>
<td></td>
</tr>
<tr>
<td>2C(s) + O₂(g) → 2CO(g)</td>
<td></td>
</tr>
<tr>
<td>CaCl₂(aq) + Na₂CO₃(aq) → 2NaCl(aq) + CaCO₃(s)</td>
<td></td>
</tr>
<tr>
<td>Fe₂O₃(s) + 3CO(g) → 2Fe(s) + 3CO₂(g)</td>
<td></td>
</tr>
<tr>
<td>CaCO₃(s) → CaO(s) + CO₂(g)</td>
<td></td>
</tr>
<tr>
<td>2Na(s) + Br₂(l) → 2NaBr(s)</td>
<td></td>
</tr>
<tr>
<td>2FeCl₃(aq) + 3 Ca(OH)₂(aq) → 3CaCl₂(aq) + 2Fe(OH)₃(s)</td>
<td></td>
</tr>
<tr>
<td>H₂SO₄(aq) + BaCl₂(aq) → BaSO₄(s) + 2HCl(aq)</td>
<td></td>
</tr>
</tbody>
</table>