Redox reactions

If you have seen a piece of rusty metal then you have seen the end result of a redox reaction (iron and oxygen forming iron oxide). Redox reactions are also used in electrochemistry and in biological reactions.

When some reactions occur, an exchange of electrons takes place. It is this exchange of electrons that leads to the change in charge that we noted in grade 10 (chapter 18, reactions in aqueous solution). When an atom gains electrons it becomes more negative and when it loses electrons it becomes more positive.

**Oxidation** is the *loss* of electrons from an atom, while **reduction** is the *gain* of electrons by an atom. In a reaction these two processes occur together so that one element or compound gains electrons while the other element or compound loses electrons. This is why we call this a redox reaction. It is a short way of saying reduction-oxidation reaction!

**Tip:**

You can remember this by using OiLRiG: Oxidation is Loss Reduction is Gain.

**Definition 1: Oxidation**

Oxidation is the *loss* of electrons by a molecule, atom or ion.
Definition 2: **Reduction**

Reduction is the *gain* of electrons by a molecule, atom or ion.

Before we look at redox reactions we need to first learn how to tell if a reaction is a redox reaction. In grade 10 you learnt that a redox reaction involves a change in the charge on an atom. Now we will look at why this change in charge occurs.

**Oxidation numbers**

By giving elements an oxidation number, it is possible to keep track of whether that element is losing or gaining electrons during a chemical reaction. The loss of electrons in one part of the reaction must be balanced by a gain of electrons in another part of the reaction.

Definition 3: **Oxidation number**

Oxidation number is the charge an atom would have if it was in a compound composed of ions.

There are a number of rules that you need to know about oxidation numbers, and these are listed below.

1. A molecule consisting of only one element always has an oxidation number of zero, since it is neutral.

   For example the oxidation number of hydrogen in $\text{H}_2$ is 0. The oxidation number of bromine in $\text{Br}_2$ is also 0.

2. Monatomic ions (ions with only one element or type of atom) have an oxidation number that is equal to the charge on the ion.

   For example, the chloride ion $\text{Cl}^-$ has an oxidation number of $-1$, and the magnesium ion $\text{Mg}^{2+}$ has an oxidation number of $+2$. 
3. In a molecule or compound, the sum of the oxidation numbers for each element in the molecule or compound will be zero.

For example the sum of the oxidation numbers for the elements in water will be 0.

4. In a polyatomic ion the sum of the oxidation numbers is equal to the charge.

For example the sum of the oxidation numbers for the elements in the sulfate ion (SO$_2^-$) will be $-2$.

5. An oxygen atom usually has an oxidation number of $-2$. One exception is in peroxides (e.g. hydrogen peroxide) when oxygen has an oxidation number of $-1$.

For example oxygen in water will have an oxidation number of $-2$ while in hydrogen peroxide (H$_2$O$_2$) it will have an oxidation number of $-1$.

6. The oxidation number of hydrogen is often $+1$. One exception is in the metal hydrides where the oxidation number is $-1$.

For example the oxidation number of the hydrogen atom in water is $+1$, while the oxidation number of hydrogen in lithium hydride (LiH) is $-1$.

7. The oxidation number of fluorine is $-1$.

Tip:

You will notice that some elements always have the same oxidation number while other elements can change oxidation numbers depending on the compound they are in.

Example 1: Oxidation numbers

Question
Give the oxidation number of sulfur in a sulfate (SO$_2^-$) ion

Answer
Determine the oxidation number for each atom
Oxygen will have an oxidation number of \(-2\). (Rule 5, this is not a peroxide.) The oxidation number of sulfur at this stage is uncertain since sulfur does not have a set oxidation number.

**Determine the oxidation number of sulfur by using the fact that the oxidation numbers of the atoms must add up to the charge on the compound**

In the polyatomic \(\text{SO}_2^4\) ion, the sum of the oxidation numbers must be \(-2\) (rule 4).

Let the oxidation number of sulfur be \(x\). We know that oxygen has an oxidation number of \(-2\) and since there are four oxygen atoms in the sulfate ion, then the sum of the oxidation numbers of these four oxygen atoms is \(-8\).

Putting this together gives:

\[
x + (-8) = -2 - 2 + 8 = +6
\]

So the oxidation number of sulfur is +6.

**Write down the final answer**

In the sulfate ion, the oxidation number of sulfur is +6.

**Example 2: Oxidation numbers**

**Question**

Give the oxidation number of both elements in ammonia (\(\text{NH}_3\)).

**Answer**

**Determine the oxidation number for each atom**

Hydrogen will have an oxidation number of +1 (rule 6, ammonia is not a metal hydride). At this stage we do not know the oxidation number for nitrogen.

**Determine the oxidation number of nitrogen by using the fact that the oxidation numbers of the atoms must add up to the charge on the compound**

In the compound \(\text{NH}_3\), the sum of the oxidation numbers must be 0 (rule 3).

Let the oxidation number of nitrogen be \(x\). We know that hydrogen has an oxidation number of +1 and since there are three hydrogen atoms in the ammonia molecule, then the sum of the oxidation numbers of these three hydrogen atoms is +3.

Putting this together gives:

\[
x + (+3) = 0 = -3
\]

So the oxidation number of nitrogen is \(-3\).
Write the final answer

Hydrogen has an oxidation number of +1 and nitrogen has an oxidation number of −3.

Example 3: Oxidation numbers

Question
Give the oxidation numbers for all the atoms in sodium chloride (NaCl).

Answer
Determine the oxidation number for each atom in the compound
This is an ionic compound composed of Na⁺ and Cl⁻ ions. Using rule 2 the oxidation number for the sodium ion is +1 and for the chlorine ion it is −1.

This then gives us a sum of 0 for the compound.

Write the final answer

The oxidation numbers for sodium is +1 and for chlorine it is −1.

Exercise 1: Oxidation numbers

Problem 1:

Give the oxidation numbers for each element in the following chemical compounds:

1. MgF₂
2. CaCl₂
3. CH₄
4. MgSO₄

Answer 1:

1. In the compound MgF₂, the oxidation number of fluorine is −1 (rule 7).
   Let the oxidation number of magnesium be x. We know that fluorine has an oxidation number of −1 and since there are two fluorine atoms in the compound, then the sum of the oxidation numbers of these two fluorine atoms is −2.
   Putting this together gives:
   
   \[ x + (-2) = 0 = +2 \]
So the oxidation number of magnesium is +2.

Magnesium has an oxidation number of +2 and fluorine has an oxidation number of −1.

2. This is an ionic compound composed of $\text{Ca}^{2+}$ and $\text{Cl}^−$ ions. Using rule 2 the oxidation number for the calcium ion is +2 and for the chlorine ion it is −1.

This then gives us a sum of 0 for the compound.

Calcium has an oxidation number of +2 and chlorine has an oxidation number of −1.

3. In the compound $\text{CH}_4$, the sum of the oxidation numbers must be 0 (rule 3).

Let the oxidation number of carbon be $x$. We know that hydrogen has an oxidation number of +1 (this is not a metal hydride) and since there are four hydrogen atoms in the molecule, then the sum of the oxidation numbers of these four hydrogen atoms is +4.

Putting this together gives:

$$x + (+4) = 0 = −4$$

So the oxidation number of carbon is −4.

Hydrogen has an oxidation number of +1 and carbon has an oxidation number of −4.

4. This is an ionic compound composed of $\text{Mg}^{2+}$ and $\text{SO}_4^{2−}$ ions. Using rule 2 the oxidation number for the magnesium ion is +2. In the polyatomic $\text{SO}_4^{2−}$ ion, the sum of the oxidation numbers must be −2 (rule 4).

Let the oxidation number of sulfur be $x$. We know that oxygen has an oxidation number of −2 (it is not in a peroxide) and since there are four oxygen atoms in the sulfate ion, then the sum of the oxidation numbers of these four oxygen atoms is −8.

Putting this together gives:

$$x + (−8)x = −2 = −2 + 8 = +6$$

So the oxidation number of sulfur is +6.

Putting all the information together we find that magnesium has an oxidation number of +2, oxygen has an oxidation number of −2 and sulfur has an oxidation number of +6.

Problem 2:
Compare the oxidation numbers of:

1. nitrogen in:
   \[ \text{NO}_2 \text{ and NO} \]

2. carbon in:
   \[ \text{CO}_2 \text{ and CO} \]

3. chromium in:
   \[ \text{Cr}_2\text{O}_7^- \text{ and CrO}_4^- \]

4. oxygen in:
   \[ \text{H}_2\text{O} \text{ and H}_2\text{O}_2 \]

5. hydrogen in:
   \[ \text{NaH} \text{ and H}_2\text{O} \]

Practise more questions like this

Answer 2:
1. In the compound \( \text{NO}_2 \), the sum of the oxidation numbers must be 0 (rule 3).
   Let the oxidation number of nitrogen be \( x \). We know that oxygen has an oxidation number of \(-2\) (this is not a peroxide) and since there are two oxygen atoms in the molecule, then the sum of the oxidation numbers of these two oxygen atoms is \(-4\). Putting this together gives:
   \[ x + (-4) = 0 = +4 \]
   So the oxidation number of nitrogen is \(+4\) in \( \text{NO}_2 \).

   In the compound \( \text{NO} \), the sum of the oxidation numbers must be 0 (rule 3).
   We know that oxygen has an oxidation number of \(-2\) (this is not a peroxide) and since there is only one oxygen atom in the molecule, then the nitrogen atom must have an oxidation number of \(+2\).
So the oxidation number of nitrogen is +2 in NO.
Nitrogen has an oxidation number of +4 in NO₂ and +2 in NO.

2. In the compound CO₂, the sum of the oxidation numbers must be 0 (rule 3).
   Let the oxidation number of carbon be \( x \). We know that oxygen has an oxidation number of −2 (this is not a peroxide) and since there are two oxygen atoms in the molecule, then the sum of the oxidation numbers of these two oxygen atoms is −4. Putting this together gives:
   \[
x + (-4) = 0 = +4
   \]
   So the oxidation number of carbon is +4 in CO₂.

   In the compound CO, the sum of the oxidation numbers must be 0 (rule 3).
   We know that oxygen has an oxidation number of −2 (this is not a peroxide) and since there is only one oxygen atom in the molecule, then the carbon atom must have an oxidation number of +2.

   So the oxidation number of carbon is +2 in CO.
   Carbon has an oxidation number of +4 in CO₂ and +2 in CO.

3. In the compound Cr₂O₇⁻, the sum of the oxidation numbers must be −2 (rule 4).
   Let the oxidation number of chromium be \( x \). We know that oxygen has an oxidation number of −2 (this is not a peroxide) and since there are seven oxygen atoms in the molecule, then the sum of the oxidation numbers of these seven oxygen atoms is −14. Putting this together gives:
   \[
   2x + (-14) = -2 = +6
   \]
   So the oxidation number of chromium is +6 in Cr₂O₇⁻.

   In the compound CrO⁻₄, the sum of the oxidation numbers must be −1 (rule 4).
   Let the oxidation number of chromium be \( x \). We know that oxygen has an oxidation number of −2 (this is not a peroxide) and since there are four oxygen atoms in the molecule, then the sum of the oxidation numbers of these four oxygen atoms is −8. Putting this together gives:
   \[
   x + (-8) = -1 = +7
   \]
   So the oxidation number of chromium is +7 in CrO⁻₄.
   Chromium has an oxidation number of +6 in Cr₂O₇⁻ and +7 in CrO⁻₄.

4. In the compound H₂O, the sum of the oxidation numbers must be 0 (rule 3).
This compound is not a metal hydride, so the oxidation number of hydrogen is +1 (rule 6). This compound is also not a peroxide, so the oxidation number of oxygen is −2 (rule 5).

We confirm that this gives us a sum of 0: \( 2(+1)+(-2)=0 \). So the oxidation number of oxygen is −2 in \( \text{H}_2\text{O} \).

In the compound \( \text{H}_2\text{O}_2 \), the sum of the oxidation numbers must be 0 (rule 4).

Let the oxidation number of oxygen be \( x \) (this is a peroxide and so oxygen does not have an oxidation number of −2). We know that hydrogen has an oxidation number of +1 (this is not a metal hydride) and since there are two hydrogen atoms in the molecule, the sum of the oxidation numbers of these two hydrogen atoms is +2.

Putting this together gives:

\[
2x+(+2)x=0=-1
\]

So the oxidation number of oxygen is −1 in \( \text{H}_2\text{O}_2 \). (Note that this confirms what has been stated about peroxides (see rule 5).)

Oxygen has an oxidation number of −2 in \( \text{H}_2\text{O} \) and −1 in \( \text{H}_2\text{O}_2 \).

5. In the compound \( \text{NaH} \), the sum of the oxidation numbers must be 0 (rule 3).

This compound is a metal hydride, so the oxidation number of hydrogen is −1 (rule 6). This compound is also an ionic compound with sodium ions and so sodium must have an oxidation number of +1 (rule 2).

We confirm that this gives us a sum of 0: \( +1+(-1)=0 \) So the oxidation number of hydrogen is −1 in \( \text{NaH} \).

In the compound \( \text{H}_2\text{O} \), the sum of the oxidation numbers must be 0 (rule 4).

This compound is not a metal hydride, so the oxidation number of hydrogen is +1 (rule 6). This compound is also not a peroxide, so the oxidation number of oxygen is −2 (rule 5).

So the oxidation number of hydrogen is +2 in \( \text{H}_2\text{O} \).

Hydrogen has an oxidation number of −1 in \( \text{NaH} \) and +1 in \( \text{H}_2\text{O} \).

Problem 3:
Give the oxidation numbers for each of the elements in all the compounds. State if there is any difference between the oxidation number of the element in the reactant and the element in the product.

1. \( \text{C (s)} + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) \)
2. \( \text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}) \)

**Practise more questions like this**

**Answer 3:**

1. In the compound C, the oxidation number of carbon is 0 (rule 1).
   In the compound \( \text{O}_2 \), the oxidation number of oxygen is 0 (rule 1).
   In the compound \( \text{CO}_2 \), the sum of the oxidation numbers must be 0 (rule 3).
   Let the oxidation number of carbon be \( x \). We know that oxygen has an oxidation number of \(-2\) (this is not a peroxide) and since there are two oxygen atoms in the molecule, then the sum of the oxidation numbers of these two oxygen atoms is \(-4\).
   Putting this together gives:
   \[
   x + (-4) = 0 = +4
   \]
   So the oxidation number of carbon is +4.

   The oxidation number of carbon in the products is 0 and in the reactants it is +4. The oxidation number has increased (become more positive).

   The oxidation number of oxygen in the products is 0 and in the reactants it is \(-2\). The oxidation number has decreased (become more negative).

2. In the compound \( \text{N}_2 \), the oxidation number of nitrogen is 0 (rule 1).
   In the compound \( \text{H}_2 \), the oxidation number of hydrogen is 0 (rule 1).
   In the compound \( \text{NH}_3 \), the sum of the oxidation numbers must be 0 (rule 3).
   Let the oxidation number of nitrogen be \( x \). We know that hydrogen has an oxidation number of +1 (this is not a metal hydride) and since there are three hydrogen atoms in the molecule, then the sum of the oxidation numbers of these three hydrogen atoms is +3.
   Putting this together gives:
   \[
   x + (+3) = 0 = -3
   \]
So the oxidation number of nitrogen is $-3$.

The oxidation number of hydrogen in the products is 0 and in the reactants it is $+1$ The oxidation number has increased (become more positive).

The oxidation number of nitrogen in the products is 0 and in the reactants it is $-3$ The oxidation number has decreased (become more negative).

**Redox reactions**

Now that we know how to determine the oxidation number of a compound, we will go on to look at how to use this knowledge in reactions.

**Oxidation and reduction**

By looking at how the oxidation number of an element changes during a reaction, we can easily see whether that element is being oxidised (lost electrons) or reduced (gained electrons).

If the oxidation number of a species becomes more positive, the species has been oxidised and if the oxidation number of a species becomes more negative, the species has been reduced.

**Tip:**

The word species is used in chemistry to indicate either a compound, a molecule, an ion, an atom or an element.

We will use the reaction between magnesium and chlorine as an example.

The chemical equation for this reaction is:

As a reactant, magnesium has an oxidation number of zero, but as part of the product magnesium chloride, the element has an oxidation number of $+2$. Magnesium has lost two electrons and has therefore been oxidised (note how the oxidation number becomes more positive). This can be written as a half-reaction. The half-reaction for this change is:

$$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$$
As a reactant, chlorine has an oxidation number of zero, but as part of the product magnesium chloride, the element has an oxidation number of −1. Each chlorine atom has gained an electron and the element has therefore been reduced (note how the oxidation number becomes more negative). The half-reaction for this change is:

\[
\text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^-
\]

**Definition 4: Half-reaction**

A half reaction is either the oxidation or reduction reaction part of a redox reaction.

In the two half-reactions for a redox reaction the number of electrons donated is exactly the same as the number of electrons accepted. We will use this to help us balance redox reactions.

Two further terms that we use in redox reactions and that you may see are reducing agents and oxidising agents.

An element that is oxidised is called a reducing agent, while an element that is reduced is called an oxidising agent.

You can remember this by thinking of the fact that when a compound is oxidised, it causes another compound to be reduced (the electrons have to go somewhere and they go to the compound being reduced).

**Redox reactions**

**Definition 5: Redox reaction**

A redox reaction is one involving oxidation and reduction, where there is always a change in the oxidation numbers of the elements involved. Redox reactions involve the transfer of electrons from one compound to another.

**Informal experiment 1: Redox reaction - displacement reaction**

**Aim**

To investigate the redox reaction between copper sulfate and zinc.
Materials
- A few granules of zinc
- 15 ml copper (II) sulfate solution (blue colour)
- glass beaker

Method

Add the zinc granules to the copper sulfate solution and observe what happens. What happens to the zinc granules? What happens to the colour of the solution?

Results
- Zinc becomes covered in a layer that looks like copper.
- The blue copper sulfate solution becomes clearer.

Cu$^{2+}$ ions from the CuSO$_4$ solution are reduced to form copper metal. This is what you saw on the zinc crystals. The reduction of the copper ions (in other words, their removal from the copper sulfate solution), also explains the change in colour of the solution (copper ions in solution are blue). The equation for this reaction is:

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

Zinc is oxidised to form Zn$^{2+}$ ions which are clear in the solution. The equation for this reaction is:

\[ \text{Zn} (s) \rightarrow \text{Zn}^{2+} (aq) + 2e^- \]

The overall reaction is:

\[ \text{Cu}^{2+} (aq) + \text{Zn} (s) \rightarrow \text{Cu} (s) + \text{Zn}^{2+} (aq) \]

Conclusion
A redox reaction has taken place. Cu\textsuperscript{2+} ions are reduced and the zinc is oxidised. This is a displacement reaction as the zinc replace the copper ions to form zinc sulfate.

**Informal experiment 2: Redox reaction - synthesis reaction**

**Aim**

To investigate the redox reaction that occurs when magnesium is burnt in air.

**Materials**

A strip of magnesium; bunsen burner; tongs; glass beaker.

**Method**

**Warning:**

Do not look directly at the flame.

1. Light the bunsen burner and use a pair of tongs to hold the magnesium ribbon in the flame.

2. Hold the lit piece of magnesium over a beaker. What do you observe?

**Results**

The magnesium burns with a bright white flame. When the magnesium is held over a beaker, a fine powder is observed in the beaker. This is magnesium oxide.

The overall reaction is:

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

**Conclusion**

A redox reaction has taken place. Magnesium is oxidised and the oxygen is reduced. This is a synthesis reaction as we have made magnesium oxide from magnesium and oxygen.

**Informal experiment 3: Redox reaction - decomposition reaction**

**Aim**

To investigate the decomposition of hydrogen peroxide.
**Materials**

Dilute hydrogen peroxide (about 3%); manganese dioxide; test tubes; a water bowl; stopper and delivery tube, Bunsen burner

![Figure 1](image)

**Warning:**

Hydrogen peroxide can cause chemical burns. Work carefully with it.

**Method**

1. Put a small amount (about 5 ml) of hydrogen peroxide in a test tube.

2. Set up the apparatus as shown above.

3. Very carefully add a small amount (about 0.5 g) of manganese dioxide to the test tube containing hydrogen peroxide.

**Results**
You should observe a gas bubbling up into the second test tube. This reaction happens quite rapidly.

The overall reaction is:

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

**Conclusion**

A redox reaction has taken place. \( \text{H}_2\text{O}_2 \) is both oxidised and reduced in this decomposition reaction.

Using what you have learnt about oxidation numbers and redox reactions we can balance redox reactions in the same way that you have learnt to balance other reactions. The following worked examples will show you how.

**Example 4: Balancing redox reactions**

**Question**

Balance the following redox reaction:

\[ \text{Fe}^{2+}(\text{aq}) + \text{Cl}_2(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Cl}^-(\text{aq}) \]

**Answer**

**Write a reaction for each compound**

\[ \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} \]

\[ \text{Cl}_2 \rightarrow \text{Cl}^- \]

**Balance the atoms on either side of the arrow**

We check that the atoms on both sides of the arrow are balanced:

\[ \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} \]

\[ \text{Cl}_2 \rightarrow 2\text{Cl}^- \]

**Add electrons to balance the charges**

We now add electrons to each reaction so that the charges balance.

We add the electrons to the side with the greater positive charge.
Balance the number of electrons

We now make sure that the number of electrons in both reactions is the same.

The reaction for iron has one electron, while the reaction for chlorine has two electrons. So we must multiply the reaction for iron by 2 to ensure that the charges balance.

\[ 2\text{Fe}^2+ \rightarrow 2\text{Fe}^3+ + 2e^- \]
\[ \text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^- \]

We now have the two half-reactions for this redox reaction.

The reaction for iron is the oxidation half-reaction as iron became more positive (lost electrons). The reaction for chlorine is the reduction half-reaction as chlorine has become more negative (gained electrons).

Combine the two half-reactions

\[
\begin{array}{c}
2\text{Fe}^2+ \\
+ \text{Cl}_2 + 2e^- \\
\end{array} \rightarrow \begin{array}{c}
2\text{Fe}^3+ + 2e^- \\
2\text{Cl}^- \\
\end{array}
\]

Write the final answer

Cancelling out the electrons gives:

\[ 2\text{Fe}^2+(aq) + \text{Cl}_2(aq) \rightarrow 2\text{Fe}^3+(aq) + 2\text{Cl}^-(aq) \]

Note that we leave the co-efficients in front of the iron ions since the charge on the left hand side has to be the same as the charge on the right hand side in a redox reaction.

In the example above we did not need to know if the reaction was taking place in an acidic or basic medium (solution). However if there is hydrogen or oxygen in the reactants and not in the products (or if there is hydrogen or oxygen in the products but not in the reactants)
then we need to know what medium the reaction is taking place in. This will help us to balance the redox reaction.

If a redox reaction takes place in an **acidic** medium then we can add water molecules to either side of the reaction equation to balance the number of oxygen atoms. We can also add hydrogen ions to balance the number of hydrogen atoms. We do this because we are writing the net ionic equation (showing only the ions involved and often only the ions containing the elements that change oxidation number) for redox reactions and not the net reaction equation (showing all the compounds that are involved in the reaction). If a Bronsted acid is dissolved in water then there will be free hydrogen ions.

If a redox reaction takes place in an **basic** medium then we can add water molecules to either side of the reaction equation to balance the number of oxygen atoms. We can also add hydroxide ions (OH\(^-\)) to balance the number of hydrogen atoms. We do this because we are writing the net ionic equation (showing only the ions involved and often only the ions containing the elements that change oxidation number) for redox reactions and not the net reaction equation (showing all the compounds that are involved in the reaction). If a Bronsted base is dissolved in water then there will be free hydroxide ions.

**Example 5: Balancing redox reactions**

**Question**
Balance the following redox reaction:

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

The reaction takes place in an acidic medium.

**Answer**
Write a reaction for each compound

- \(\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^3+\)
- \(\text{H}_2\text{S} \rightarrow \text{S}\)

Balance the atoms on either side of the arrow

We check that the atoms on both sides of the arrow are balanced.

In the first reaction we have 2 chromium atoms and 7 oxygen atoms on the left hand side. On the right hand side we have 1 chromium atom and no oxygen atoms. Since we are in an acidic medium we can add water to the right hand side to balance the number of oxygen
atoms. We also multiply the chromium by 2 on the right hand side to make the number of chromium atoms balance.

\[
\text{Cr}_2\text{O}_2^-7 \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

Now we have hydrogen atoms on the right hand side, but not on the left hand side so we must add 14 hydrogen ions to the left hand side (we can do this because the reaction is in an acidic medium):

\[
\text{Cr}_2\text{O}_2^-7 + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

We do not use water to balance the hydrogens as this will make the number of oxygen atoms unbalanced.

For the second part of the reaction we need to add 2 hydrogen ions to the right hand side to balance the number of hydrogens:

\[
\text{H}_2\text{S} \rightarrow \text{S} + 2\text{H}^+
\]

**Add electrons to balance the charges**

We now add electrons to each reaction so that the charges balance.

We add the electrons to the side with the greater positive charge.

\[
\text{Cr}_2\text{O}_2^-7 + 14\text{H}^+ + 6e^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\]

\[
\text{H}_2\text{S} \rightarrow \text{S} + 2\text{H}^+ + 2e^-
\]

**Balance the number of electrons**

We now make sure that the number of electrons in both reactions is the same.

The reaction for chromium has 6 electrons, while the reaction for sulfur has 2 electrons. So we must multiply the reaction for sulfur by 3 to ensure that the charges balances.

\[
3\text{H}_2\text{S} \rightarrow 3\text{S} + 6\text{H}^+ + 6e^-
\]

We now have the two half-reactions for this redox reaction.
The reaction involving sulfur is the oxidation half-reaction as sulfur became more positive (lost electrons). The reaction for chromium is the reduction half-reaction as chromium has become more negative (gained electrons).

**Combine the two half-reactions**

We combine the two half-reactions:

\[
\begin{align*}
\text{Cr}_2\text{O}_2^- - 7 + 14\text{H}^+ + 6e^- & \rightarrow 2\text{Cr}^3+ + 7\text{H}_2\text{O} \\
+ 3\text{H}_2\text{S} & \rightarrow 3\text{S} + 6\text{H}^+ + 6e^- \\
\text{Cr}_2\text{O}_2^- - 7 + 14\text{H}^+ + 6e^- + 3\text{H}_2\text{S} & \rightarrow 2\text{Cr}^3+ + 7\text{H}_2\text{O} + 3\text{S} + 6\text{H}^+ + 6e^- 
\end{align*}
\]

Table 2

**Write the final answer**

Crossing off the electrons and hydrogen ions gives:

\[\text{Cr}_2\text{O}_2^- (aq) + 3\text{H}_2\text{S} (aq) + 8\text{H}^+(aq) \rightarrow 2\text{Cr}^3+(aq) + 3\text{S} (s) + 7\text{H}_2\text{O} (l)\]

In Grade 12, you will go on to look at electrochemical reactions, and the role that electron transfer plays in this type of reaction.

**Exercise 2: Redox reactions**

Problem 1:

Consider the following chemical equations:

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

Which one of the following statements is correct?

1. Fe is oxidised and H$^+$ is reduced
2. Fe is reduced and O$_2$ is oxidised
3. Fe is oxidised and O$_2$ is reduced
4. Fe is reduced and H$^+$ is oxidised

(DoE Grade 11 Paper 2, 2007)
Answer 1:
Fe is oxidised and O₂ is reduced

Problem 2:
Which one of the following reactions is a redox reaction?

1. HCl (aq)+NaOH (aq)→NaCl (aq)+H₂O (l)
2. AgNO₃(aq)+NaI (aq)→AgI (s)+NaNO₃(aq)
3. 2FeCl₃(aq)+2H₂O (l)+SO₂(aq)→H₂SO₄(aq)+2HCl (aq)+2FeCl₂(aq)
4. BaCl₂(aq)+MgSO₄(aq)→MgCl₂(aq)+BaSO₄(s)

Answer 2:
FeCl₃(aq)+2H₂O (l)+SO₂(aq)→H₂SO₄(aq)+2HCl (aq)+2FeCl₂(aq)

Problem 3:
Balance the following redox reactions:

1. Zn (s)+Ag⁺(aq)→Zn²⁺(aq)+Ag (s)
2. Cu₂⁺(aq)+Cl⁻(aq)→Cu (s)+Cl₂(g)
3. Pb²⁺(aq)+Br⁻(aq)→Pb (s)+Br₂(aq)
4. HCl (aq)+MnO₂(s)→Cl₂(g)+Mn²⁺(aq)

This reaction takes place in an acidic medium.

Answer 3:
1. Write a reaction for each compound:
   Zn→Zn²⁺
   Ag⁺→Ag

   The atoms are balanced.
Add electrons to each reaction so that the charges balance. We add the electrons to the side with the greater positive charge.

\[
\begin{align*}
\text{Zn} & \rightarrow \text{Zn}^{2+} + 2e^- \\
\text{Ag}^{++} + e^- & \rightarrow \text{Ag}
\end{align*}
\]

We now make sure that the number of electrons in both reactions is the same.

The reaction for zinc has two electrons, while the reaction for silver has one electron. So we must multiply the reaction for silver by 2 to ensure that the charges balance.

\[
\begin{align*}
\text{Zn} & \rightarrow \text{Zn}^{2+} + 2e^- \\
2\text{Ag}^{++} + 2e^- & \rightarrow 2\text{Ag}
\end{align*}
\]

We combine the two half-reactions:

\[
\begin{array}{c}
\text{Zn} \\
+ 2\text{Ag}^{++} + 2e^- \\
\text{Zn} + \text{Ag}^{++} + 2e^- \\
\rightarrow \\
\rightarrow \\
\rightarrow \\
\rightarrow \\
\text{Zn}^{2+} + 2\text{Ag} + 2e^- \\
2\text{Ag} \\
\text{Zn}^{2+} + 2\text{Ag} + 2e^-
\end{array}
\]

Table 3

Cancelling out the electrons gives:

\[
\text{Zn (s)} + 2\text{Ag}^{+} (\text{aq}) \rightarrow \text{Zn}^{2+} (\text{aq}) + 2\text{Ag} (\text{s})
\]

2. Write a reaction for each compound:

\[
\begin{align*}
\text{Cu}^{2+} & \rightarrow \text{Cu} \\
\text{Cl}^- & \rightarrow \text{Cl}_2
\end{align*}
\]

Balance the atoms:

\[
\begin{align*}
\text{Cu}^{2+} & \rightarrow \text{Cu} \\
2\text{Cl}^- & \rightarrow \text{Cl}_2
\end{align*}
\]
Add electrons to each reaction so that the charges balance. We add the electrons to the side with the greater positive charge.

\[
\begin{align*}
\text{Cu}^2+ + 2e^- & \rightarrow \text{Cu} \\
2\text{Cl}^- & \rightarrow \text{Cl}_2 + 2e^-
\end{align*}
\]

We now make sure that the number of electrons in both reactions is the same. The reaction for copper has two electrons and the reaction for chlorine also has two electrons. So the charges balance.

We combine the two half-reactions:

<p>| | | |</p>
<table>
<thead>
<tr>
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</thead>
<tbody>
<tr>
<td>$\text{Cu}^2+ + 2e^-$</td>
<td>$\rightarrow$</td>
<td>$\text{Cu}$</td>
</tr>
<tr>
<td>$+$</td>
<td>$2\text{Cl}^-$</td>
<td>$\rightarrow$</td>
</tr>
<tr>
<td>$\text{Cu}^2+ + 2\text{Cl}^- + 2e^-$</td>
<td>$\rightarrow$</td>
<td>$\text{Cu} + \text{Cl}_2 + 2e^-$</td>
</tr>
</tbody>
</table>

Table 4

Cancelling out the electrons gives:

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

3. Write a reaction for each compound:

\[
\begin{align*}
\text{Pb}^2+ & \rightarrow \text{Pb} \\
\text{Br}^- & \rightarrow \text{Br}_2
\end{align*}
\]

Balance the atoms:

\[
\begin{align*}
\text{Pb}^2+ & \rightarrow \text{Pb} \\
2\text{Br}^- & \rightarrow \text{Br}_2
\end{align*}
\]

Add electrons to each reaction so that the charges balance. We add the electrons to the side with the greater positive charge.

\[
\begin{align*}
\text{Pb}^2+ + 2e^- & \rightarrow \text{Pb} \\
2\text{Br}^- & \rightarrow \text{Br}_2 + 2e^-
\end{align*}
\]
We now make sure that the number of electrons in both reactions is the same.

The reaction for lead has two electrons and the reaction for bromine also has two electrons. So the charges balance.

We combine the two half-reactions:

\[
\begin{align*}
\text{Pb}^{2+} + 2e^- & \rightarrow \text{Pb} \\
+ 2\text{Br}^- & \rightarrow \text{Br}_2 + 2e^- \\
\text{Pb}^{2+} + 2\text{Br}^- + 2e^- & \rightarrow \text{Pb} + \text{Br}_2 + 2e^-
\end{align*}
\]

Table 5

Cancelling out the electrons gives:

\[
\text{Pb}^{2+} (aq) + 2\text{Br}^- (aq) \rightarrow \text{Pb} (s) + \text{Br}_2 (g)
\]

4. Write a reaction for each compound:

\[
\begin{align*}
\text{HCl} & \rightarrow \text{Cl}_2 \\
\text{MnO}_2 & \rightarrow \text{Mn}^{2+}
\end{align*}
\]

Balance the atoms.

In the first reaction we have 1 chlorine atom on the left hand side and 2 chlorine atoms on the right hand side. So we multiply the left hand side by 2:

\[
2\text{HCl} \rightarrow \text{Cl}_2
\]

Now we note that there are 2 hydrogen atoms on the left hand side and no hydrogen atoms on the right hand so we add 2 hydrogen ions to the right hand side:

\[
2\text{HCl} \rightarrow \text{Cl}_2 + 2\text{H}^+
\]

The equation is now balanced.

For the second equation we need to add two water molecules to the right hand side:
\[ \text{MnO}_2 \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O} \]

We now need to add four hydrogen ions to the left hand side to balance the hydrogens:

\[ \text{MnO}_2 + 4\text{H}^+ \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O} \]

This equation is now balanced.

Add electrons to each reaction so that the charges balance. We add the electrons to the side with the greater positive charge.

\[ 2\text{HCl} \rightarrow \text{Cl}_2 + 2\text{H}^+ + 2e^- \]
\[ \text{MnO}_2 + 4\text{H}^+ + 2e^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O} \]

We now make sure that the number of electrons in both reactions is the same.

The reaction for chlorine has two electrons and the reaction for manganese also has two electrons. So the charges balance.

We combine the two half-reactions:

\[
\begin{array}{c}
2\text{HCl} \\
+ \text{MnO}_2 + 4\text{H}^+ + 2e^- \\
\end{array}
\rightarrow
\begin{array}{c}
\text{Cl}_2 + 2\text{H}^+ + 2e^- \\
\text{Mn}^{2+} + 2\text{H}_2\text{O} \\
\end{array}
\]

Table 6

Cancelling out the electrons and hydrogen ions gives:

\[ 2\text{HCl} + \text{MnO}_2 + 2\text{H}^+ \rightarrow \text{Cl}_2 + \text{Mn}^{2+} + 2\text{H}_2\text{O} \]