Half equations:

Learning outcomes:

Demonstrate knowledge, understanding and application of:

- *Rules for calculating oxidation number.*
- *Balancing half equations by adding electrons, H^+ ions, and water.*
- *Combining half equations*

Let's first recap some key definitions:

- → Oxidation = Loss of electrons or increase in oxidation number
- → Reduction = Gain of electrons or decrease in oxidation number
- → Oxidising agent = Gain of electrons – themselves reduced
- → Reducing agent = Loss of electrons – themselves oxidised.

Make sure you have fully understood Oxidation and Reduction before revising this section

Oxidising and reducing agents:

In a redox reaction, there will always be an oxidising agent and reducing agent.

Study Tip:

Oscar **A**te **G**reen **O**ranges = **O**xidising **A**gents **G**ive **O**xygen (or accept electrons)

Ruby **A**te **R**ed **O**ranges = **R**educing **A**gent **R**emove **O**xygen (or donate electrons)

For example: Let's have a look what is happening to the reaction between Ag^+ ions and Cu as shown below:

$$2\overset{+1}{Ag^+}(aq) + \underset{0}{Cu}(s) \rightarrow \underset{+2}{Cu^{2+}} + 2\overset{0}{Ag}(s)$$

- Silver's oxidation number goes from +1 to 0 therefore Ag^+ has been **reduced. = Oxidising agent. Therefore, Silver is the oxidising agent.**
- Copper's oxidation number goes from 0 to +2 so therefore Cu has been **oxidised = reducing agent**

Reminder:

Redox reaction in terms of electrons:

- Oxidation is the **gain** of electrons.
- Reduction is the **loss** of electrons.

Redox reaction in terms of Oxidation number:

- Reduction is the **decrease** in oxidation number. (e.g. 0 to -1)
- Oxidation is the **increase** in oxidation number. (e.g. -1 to 0)

Rules to remember when calculating the oxidation number:

- ✓ The oxidation number is always zero for elements.
  - ➔ **E.g.:** Na(s) has an oxidation state of 0
  - ➔ **E.g.:** H_2(g) has an oxidation state of 0
- ✓ Ions have an oxidation number of their charge.
  - ➔ **E.g.:** Mg^{2+} has a state of +2
- ✓ Hydrogen has an oxidation state of -1 if it is present with a metal.
  - ➔ **E.g.:** NaH, Na has an oxidation state of +1 so hydrogen must be -1 to make the species neutral.
- ✓ If hydrogen is present with a non-metal, it will have an oxidation state of +1
  - ➔ E.g.: H_2O, hydrogen must have an oxidation state of +1 to cancel out the charge.
- ✓ Oxygen will always have an oxidation state of -2.
- ✓ If oxygen is present with in a compound with fluorine, it will have an oxidation state of +2 or +1
- ✓ If oxygen is present in a peroxide compound, it will have an oxidation state of -1

Redox reaction means that one species is reduced, and the other is oxidised. Working out changes in oxidation number can help to deduce which species are oxidised and which are reduced. They can also help balance the equation for the reaction. You may also have spectator ions present in your equation.

There are 3 ways you can balance a chemical redox reaction:

1. **By adding electrons**
2. **By balancing the number of atoms on each side of the equation.**
3. **By adding H_2O and H^+**

Simple redox equations:

Let's first have a look at balancing simple redox reactions by adding electrons.

Worked example:

E.g.: Balance this half equation:

$$Cu^{2+}(aq) \rightarrow Cu(s)$$

Step 1: check if both sides have equal charges

- There are 2 positive charge on the left hand side of the equation

Step 2: Balance the charge by adding electrons

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$$

- We need to add 2 electrons to the left hand side to balance the equation

Step 3: Deduce if it is an oxidation/reduction reaction:

- This reaction shows an reduction reaction because copper gains 2 electrons

Worked example:

E.g.: Balance this half equation:

$$Fe^{3+} \rightarrow Fe^{2+}$$

Step 1: check if both sides have equal charges

- There is more positive on the left hand side than the right side.

Step 2: Balance the charge by adding electrons

$$Fe^{3+} \rightarrow Fe^{2+} + e-$$

- We need to add an electron to the right hand side to balance the charge

Step 3: Deduce if it is an oxidation/reduction reaction:

- This reaction shows an oxidation reaction because Iron (II) loses an electron

Oxidation and reduction in word equations:

In an exam, they may give you a word equation from which you need to deduce which species is oxidised and reduced.

Let's go through an example:

Worked example:

$$\overset{0}{Cl_2} + 2K\overset{-1}{Br} \rightarrow 2K\overset{-1}{Cl} + \overset{0}{Br_2}$$

- Oxidation number of Chlorine goes from 0 to -1 = reduced
- Oxidation number of Bromine goes from -1 to 0 = oxidised
- K^{+} remains K^{+} so it is a Spectator ion.

As you can see in the diagram, you can deduce whether an species is being oxidised or reduced by looking at its oxidation number.

Study Tip:

Reduction is the **decrease** in oxidation number. (e.g. 0 to -1)

Oxidation is the **increase** in oxidation number. (e.g. -1 to 0)

<u>Harder redox half equations:</u>

Oxidation numbers can be used to write and balance a redox equation.

By assigning oxidation numbers to all species in an equation, it is easy to see which species have been reduced and which have been oxidised.

In an exam, you may be asked to balance and combine two half equations together with unequal number of atoms, electrons, and charge.

To balance these reactions, it is recommended that you follow this order to balance the half equation and to avoid any mistakes:

1. **You start by balancing the number of each species on each side of the equation. (Avoid H_2 and O_2)**
2. **Balance the number of oxygen (if present) by adding H_20.**
3. **Balance the no. of hydrogen by adding H^+**
4. **Balance the charges on each side of the equation by adding electrons.**

Let's go through an example:

The redox reaction between hydrogen bromide and sulfuric acid has the following half equations. Write the overall redox reactions.

$HBr \rightarrow Br_2$

$SO_4^{2-} \rightarrow SO_2$

<u>Let's first balance the 1st half equation:</u>

Step 1: Balance the number of atoms on each side of the equation.

$HBr \rightarrow Br_2$

As you can see, there is one bromine on the left-hand side and 2 bromines on the right-hand side. To balance this, we need to add a big 2 in front of the HBr to make the no. of bromine equal.

$2HBr \rightarrow Br_2$

***** If you find balancing equations difficult, check out our balancing equation section*****

Step 2: Balance the number of hydrogens.

*** We don't need to balance oxygen because it is not present in this half equation***

Now there is an uneven distribution of hydrogen as there is more hydrogen on the left-hand side than the left.

We don't have any more hydrogen on the right-hand side as no hydrogen is present. However, we can add $2H^+$ ions to balance the total number of hydrogens. So now the equation becomes...

$2HBr \rightarrow Br_2 + 2H^+$

Now the number of hydrogens on each side of the equation are balanced.

Step 3: Balancing the charge.

$2HBr \rightarrow Br_2 + 2H^+$

Now we check if the charge of each side is the same.

As we can see, the right-hand side has an overall charge of 0 but the left-hand side has an overall charge of +2 due to the $2H^+$ ion.

To balance the overall charge of the equation, we need to add electrons to one side of the equation to make the RHS = LHS.

For this, we can add 2 electrons to the right-hand side like this:

$2HBr \rightarrow Br_2 + 2H^+ + 2e-$

This half equation is now balanced. Make sure to count the number of atoms on each side to ensure you have not made an error.

Let's now have a look at the 2nd half equation. This is harder to do but you have to be patient with it and count carefully.

$SO4^{2-} \rightarrow SO_2$

Step 1: Balance the no. of atoms.

All elements are balanced apart from oxygen. Remember avoiding balancing oxygen in the first step.

Step 2: Balance the no. of oxygen atoms.

$SO4^{2-} \rightarrow SO_2$

As you can see, the right-hand side has 4 oxygen atoms, but the left-hand side only has 2 oxygen atoms. We can only balance the no. of oxygen atoms by adding H_2O.

So, we need to add 2 H_20 molecules to the right-hand side to balance out the number of oxygen atoms like this…

$SO4^{2-} \rightarrow SO_2 + 2H_20$

Now the number of oxygen atoms are balanced.

Step 3: Balance the no. of hydrogens.

After we balanced the no. of oxygen atoms, we now have an unbalanced number of hydrogens on each side. We have 4 hydrogen atoms on the right-hand side and none of the left-hand side. So, we now need to add $4H^+$ ions to balance the equation like this…

$4H^+ + SO_4^{2-} \rightarrow SO_2 + 2H_20$

Now the number of hydrogen molecules are balanced on both sides of the equation.

Step 4: Balance the overall charge.

$4H^+ + SO_4^{2-} \rightarrow SO_2 + 2H_20$

As you can see, the left-hand side has an overall charge of +2 (you add the charges together to find the overall charge so in this case you add (-2) + 4 = +2), and the right-hand side has an overall neutral charge = no charge.

To make both sides have the same charge, we need to add 2 electrons to the left-hand side like this…

$4H^+ + SO_4^{2-} + 2e- \rightarrow SO_2 + 2H_20$

This half equation is now balanced.

Combining half equations together:

So now we have both of our half equations balanced, we can now combine them to make one full equation. For this, we need to compare both of the half equations and cancel out similar species.

As you can see, both half equations have 2 electrons present on both sides of the equation. We can cancel these out like this…

$$2HBr \rightarrow Br_2 + 2H^+ + \cancel{2e-}$$

$$4H^+ + SO_4^{2-} + \cancel{2e-} \rightarrow SO_2 + 2H_20$$

Now the spectator species have been cancelled out, all we need to do now is to combine all the reactants together and all the products together in one equation like this…

We can also simplify the no. of H^+ ions by bringing all the H^+ ions onto one side of the equation as well like this.

$$2HBr + \overset{2}{\cancel{4H^+}} + SO_4^{2-} \rightarrow Br_2 + \cancel{2H^+} + SO_2 + 2H_20$$

$$\Downarrow$$

$$2HBr + 2H^+ + SO_4^{2-} \rightarrow Br_2 + SO_2 + 2H_20$$