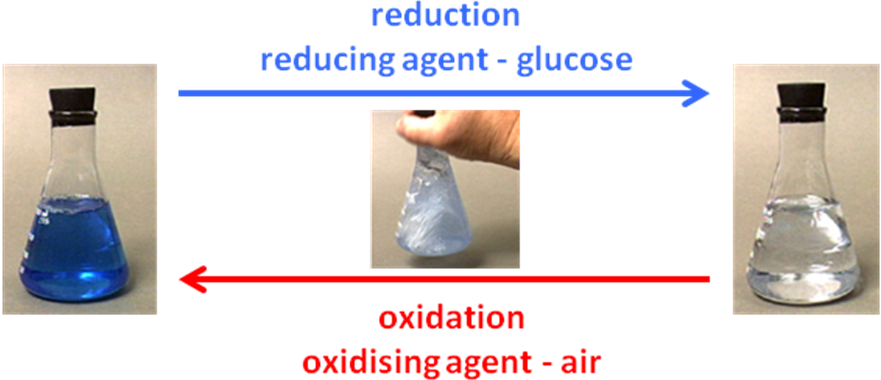
**3.1.7 Redox**

***Demo: Blue bottle experiment***



***Demo: Displacement Zn + CuO***

*Zn(s) + CuO(s) → Cu(s) + ZnO(s)*

**Zinc** has been **oxidised** and **copper oxide** has been **reduced**. Oxidation and reduction has taken in the same reaction. This is referred to as a **redox** reaction and is short for **reduction-oxidation**.

Originally oxidation was used to refer to reactions involving the **addition of oxygen** to metals.

2Mg + O2 → 2MgO

**Magnesium** has been **oxidised** to magnesium oxide, **oxygen** is called the **oxidising agent**.

The reverse process is **reduction**, in which **oxygen is removed**.

Fe2O3 + 2Al → Al2O3 + 2Fe

Aluminium is acting as a **reducing agent**.

This gave rise to these definitions:

|  |  |
| --- | --- |
| * **Oxidation**   + Addition of oxygen | * **Reduction**   + Removal of oxygen |
| * **Oxidising agent**   + Gives oxygen   + It is reduced | * **Reducing agent**   + Gains oxygen   + It is oxidised |

**Hydrogen** is sometimes **used** to **remove oxygen**, so the **addition** of **hydrogen** is also called **reduction**. The reverse, **removing hydrogen**, was called **oxidation**.

H2 + Cl2 → 2HCl

Consider these reactions:

SO2 + H2O + HgO → H2SO4 + Hg

SO2 + 2H2O + Cl2 → H2SO4 + 2HCl

In the first reaction HgO has been reduced (it has lost oxygen). In the second reaction chlorine has also been reduced (it has gained hydrogen). But what are the reducing agents in each reaction?

To fully understand the **definitions of oxidation and reduction need to be redefined**.

**OIL**

**RIG**

***Definitions***:

**Oxidation** is loss of electrons

**Reduced** is gain of electrons



Therefore:

***Definitions***:

**Reducing agent** is an electron donor

**Oxidising agent** is an electron acceptor

Since **redox** reactions always in involve the **movement of electrons** they are sometimes called **electron transfer reactions**.

**Example CGP135**

Using the new definitions the oxidising and reducing agents can be determined.

Na + ½ Cl2 → Na+ Cl-

e-

So sodium has **lost an electron**, it has been **oxidised** and is acting as the **reducing agent**. The chlorine **accepts the electron**, so it is **reduced** and acts as the **oxidising agent**.

***Task: Identify which species have been oxidised and reduced through the lost and gain of electrons and so what are the oxidising and reducing agents***

2Mg + O2 → 2MgO

Magnesium – lost 2 electrons – oxidised – reducing agent

Oxygen – gained electrons – reduced – oxidising agent

Fe2O3 + 2Al → Al2O3 + 2Fe

Aluminium – lost 2 electrons – oxidised – reducing agent

Iron(III) oxide – gained electrons – reduced – oxidising agent

**Oxidation states**

The idea of **oxidation states** is used to identify what has been oxidised and reduced in a redox reaction. Oxidation states are also called **oxidation numbers**. Each element in a compound is given an oxidation state. In an **ionic compound** it shows how many **electrons** the element has **gained** or **lost**. In a **molecule**, it shows the **distribution** of **electrons** between the elements, depending on their **electronegativity**. The **more** **electronegative** element is given the **negative** **oxidation** **state**.

The oxidation state is a number given to an atom or ion to describe its **relative state of oxidation or reduction**. A **change in oxidation states** shows **oxidation or reduction**.

Rules for finding oxidation states

|  |  |  |
| --- | --- | --- |
| **Species** | **Oxidation state** | **Example** |
| Element | 0 | O2 |
| Simple ions   * Group 1 * Group 2 * Aluminium | The charge of the ion  +1  +2  +3 | Na+  Mg2+  Al3+ |
| Fluorine | -1 | F- |
| Chlorine (except in compounds with F & O)  Chlorine in compounds with F & O | -1  Positive values (varies) | Cl- |
| Oxygen (except peroxides and OF2)  Oxygen in peroxides  Oxygen in OF2 | -2  -1  +2 | O2-  O-  O2+ |
| Hydrogen (except metal hydrides)  Hydrogen in metal hydrides | +1  -1 | H+  H- |
| Neutral compounds | Sum of all oxidation states totals zero |  |
| Ions | Sum of all oxidation states totals the charge of the ion |  |
| Central atom in a complex ion  (ion consisting of several atoms) | Charge it would have if it existed as a solitary simple ion without bonds to the other species |  |

Important information:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| more POSITIVE | |  | more NEGATIVE | |
| the **more** the element has been **oxidised** | |  | the **more** the element has been **reduced** | |
| + |  | | | - |

Finding oxidation states

**From systematic names**

Some elements have **multiple oxidation states** so when writing the name of the element in a compound **Roman numerals** are used to **show the oxidation state**.

1. = +1
2. = +2
3. = +3 etc.

**Example CGP137**

Iron(II) sulphate iron has oxidation state +2 so formula FeSO4

Iron(III) sulphate iron has oxidation state +3 so formula Fe2(SO4)3

This is useful for the group of compounds that end **–ate** (carbonate, sulphate, nitrate etc.) which contain **oxygen bonded to another non-metal** to form an ion. The other non-metal can exist in **different oxidation states**.

**Example CGP137**

Sulphate(VI) sulphur has oxidation state +6 ion formula SO42-

Sulphate(IV) sulphur has oxidation state +4 ion formula SO32-

Nitrate(III) nitrogen has oxidation state +3 ion formula NO2-

**From formula**

The formula of a **compound or ion** can be used to find the oxidation state of a particular element by looking at elements whose **oxidation states** you **know** from the rules and then **deduce** the oxidation states of any **other element**. Remember that for **neutral compounds** the **sum** of all **oxidation states** totals **zero** and for **ions** the **sum** of all **oxidation states** totals the **charge of the ion**.

**Examples**

Cu in CuO O = -2 total = 0 Cu = +2

Cu in Cu2O O = -2 total = 0 Cu = +1

O in H2O H = +1 total = 0 O = +5O in H2O2 H = +1 2 x +1 = +2 total = 0 O = -1

Zn in Zn(OH)2 H = +1, O = -2 2 x (+1-2) = -2 total = 0 Zn = +2

Oxo-ion complexes

**Example PO43-**

Overall charge is 3-, therefore sum of oxidation numbers equals -3:

Oxidation state of phosphorus + (4 x oxidation state of oxygen) = -3

Hence: oxidation state of phosphorous – (4 x -2) = -3

Thus: **oxidation state of phosphorous = +5**

***Task: Find oxidation state of the central atom and give the systematic name of the ion***

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Species** | **Number of oxygen atoms** | **Total oxidation number due to oxygen** | **Overall charge on the ion** | **Oxidation state of central atom** | **Name of species** |
| SO42- | 4 | -8 | -2 | +6 | sulphate(VI) |
| NO3- | 3 | -6 | -1 | +5 | nitrate(V) |
| CO32- | 3 | -6 | -2 | +4 | carbonate(IV) |
| ClO3- | 3 | -6 | -1 | +5 | chlorate(V) |
| ClO- | 1 | -2 | -1 | +1 | chlorate(I) |

***Whiteboard task: What’s the oxidation number?***

|  |  |
| --- | --- |
| 1. MgCl2 2. PCl5 3. H2S 4. CuO 5. HNO3 6. AlCl4- 7. VO3- 8. NH4+ 9. Cr2O72- 10. AsO43- | +2  +5  -2  +2  +5  +3  +5  -3  +6  +5 |

***Starter: 9.1.1 – Oxidation numbers***

***Sheet: Oxidation number***

***Sheet: HSW - Redox Q1-5***

***Sheet: Extension - Oxidation states***

***Application: CGP137 PQ1-6***

***Fact recall: CGP138 Q1-8***

**Redox equations**

Oxidation states can be used to **identify** which **element** has been **oxidised** and which **reduced** in a redox reaction and hence if a reaction is a redox. Redox reactions can be summarised:

|  |  |
| --- | --- |
| If the oxidation state of an element **increases** it has been **oxidised.**  If the oxidation state of an element **reduces** it has been **reduced.** |  |

The same element can exist in **different oxidation states**, the **transition elements** have many oxidation states.

***Demo: Different oxidation states of vanadium***

Ammonium metavanadate (NH4VO3) + HCl + Zn

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| |  |  |  |  |  |  |  | | --- | --- | --- | --- | --- | --- | --- | | VO3- | → | VO2+ | → | V2+ | → | V2+ | | yellow | → | blue | → | green | → | lilac | | +5 | → | +4 | → | +3 | → | +2 | |  |

Identifying redox reactions

**Example**

In the reaction: Fe2O3 + 2Al → Al2O3 + 2Fe

This can be written to show the metal ions: Fe3+ + Al → Fe + Al3+

The changes that occur with electrons are shown in the following half-equations:

Fe3+ + 3e- → Fe reduction

+3 0 reduced

Al → Al3+ + 3e- oxidation

0 +3 increased

The oxidation state of **iron** changes from **+3** in Fe2O3 **to 0** in Fe, **reduction** has occurred.

The oxidation state of **aluminium** changes from **0** in Al **to +3** in Al2O3 to, **oxidation** has occurred.

***Task: Explain why the following is a redox reaction & write an ionic equation***

Mg + 2HCl → MgCl2 + H2

0 +1 -1 +2 -1 0

Oxidation state of magnesium has increased so it has been oxidised – reducing agent

Oxidation state of hydrogen has reduced so it has been reduced – oxidising agent

-2 e- oxidation

Ionic equation: Mg + 2H+ → Mg2+ + H2

+2 e- reduction

**Example**

Explain why the following is not a redox reaction

MgO + 2HCl → MgCl2 + H2O

+2 -2 +1 -1 +2 -1 +1-2

It appears that this could be a redox reaction as magnesium oxides loses an oxygen. However, there is no change in oxidation states for any element. It is a reaction between an acid and a base.

Ionic equation: 2H+ + O2- → H2O

***Task: Work out if the following are redox reactions, state the oxidising/reducing agent and give the ionic equation***

PbO2 + 4HCl → PbCl2 + Cl2 + 2H2O

+4 -2 +1 -1 +2 -1 0 +1 -2

Oxidised: Cl- to Cl Oxidising agent: PbO2

Reduced: Pb4+ or Pb(IV) to Pb2+ Pb(II) Reducing agent: HCl

+2 e- reduction

Ionic: Pb4+ + 2Cl- → Pb2+ + Cl2

-2 e- oxidation

ZnO + 2HCl → ZnCl2 + H2O

+2 -2 +1 -1 +2 -1 +1 -2

No change in oxidation state – neutralisation

Ionic: O2- + 2H+ → H2O

In some reactions and **one** element may be **oxidised and reduced** to give two products with **different oxidation numbers**, this is called **disproportionation**.

***Sheet: Redox or not***

***Extension: Redox equations Q1***

**Constructing half-equations**

The overall **redox** equation can be **separated** into two **half-equations**; one shows **oxidation** and the other **reduction**. In each case they are **balanced** using **electrons** so that the **overall charge** on both sides of the half-equation is the **same**.

Each half-equation only shows the species involved in the reaction. Only the initial and final species in a redox equation are needed to construct half-equations.

When constructing a half-equation for reactions occurring in **aqueous solutions water provides a source of oxygen** and any **surplus oxygen** is **converted** into **water** by **reaction** with **hydrogen** ions from an **acid**.

Method:

* Write down initial and final species - only one element in each half-equation
* Balance for atoms
* Add oxidation numbers
* Balance for charge (add electrons)
* Balance oxygens (add water)
* Balance hydrogens (add H+)

**Example**

Deduce the half-equation for the reduction, in acid solution, of NO3- to NO.

Answer

NO3- → NO

+5 -2 +2-2

Oxidation state of nitrogen changes from +5 to +2, so nitrogen is reduced so 3 electrons are added to balance the charge

NO3- + 3e- → NO

+5 -3 +2

Oxidation state of oxygen is still -2 but two of the atoms combine with four hydrogen ions to make water

NO3- + 3e- → NO + 2H2O N.B. 1 oxygen atom in NO

NO3- + 3e- + 4H+ → NO + 2H2O

***Task: Deduce the half-equation for the reduction, in acid solution, of   
MnO4- to Mn2+***

Write down start and finish

MnO4-  → Mn2+

Calculate oxidation numbers

MnO4-  → Mn2+

+7 → +2

Add electrons to balance oxidation numbers

MnO4-  + 5e- → Mn2+

+7 + -5 → +2

Balance oxygens

MnO4-  + 5e- → Mn2+ + 4H2O

Balance hydrogens

MnO4-  + 5e- + 8H+ → Mn2+ + 4H2O

***Task (challenging): Deduce the half-equation for the reduction, in acid solution, of   
Cr2O72- to Cr3+***

N.B. dichromate v. powerful oxidising agent in acid – ethanol oxidised (breathalyser test)

Write down start and finish

Cr2O72- → Cr3+

yellow green

Balance for atoms

Cr2O72- → 2Cr3+

Calculate oxidation numbers

Cr2O72- → 2Cr3+

+6 +3

Add electrons to balance oxidation numbers

Cr2O72- + 6e- →2Cr3+ N.B. 2 moles so need 2x3e-

+12 -6 +6

Balance oxygens

Cr2O72- + 6e- → 2Cr3+ + 7H2O

Balance any hydrogens

Cr2O72- + 6e- + 14H+ → 2Cr3+ + 7H2O

***Starter: 9.1.2 – Redox: Half equations***

***Sheet: Redox – equations***

***Extension: RSC G&T Oxidation numbers Part 2***

***Extension: Redox equations Q3***

***Application: CGP140 PQ1-5***

***Fact recall: CGP140 Q1-2***

Constructing overall equations for redox reactions

**Overall equations** for any redox reaction can be obtained by **putting the 2 half-equations** together. They must **balance for atoms** of each element and the **total charge** on each side must be the **same**. So the number of **electrons given** by the **reducing agent** exactly **balances** the number of **electrons gained** by the **oxidising agent**. There may be a need to **cancel** for **water** or **hydrogen ions**.

**Example**

Write half-equations for the oxidation of bromide ions and reduction of chlorine, and use these to deduce the overall equation for the reaction.

Answer

The half equation for the reduction of chlorine is:

Cl2 → 2Cl-

0 -1

Cl2 + 2e- → 2Cl-

The half equation for the oxidation of bromide ions is:

Br- → Br2

-1 0

2Br- → Br2 + 2e-

Make sure the electrons balance, so they will cancel out:

Cl2 + 2e- → 2Cl-

2Br- → Br2 + 2e-

Cl2 + 2Br- → 2Cl- + Br2

OA RA

**Example**

Write half-equation for the oxidation of copper and reduction of nitric acid, and use these to deduce the overall equation for the reaction.

Answer

The half equation for the oxidation of copper is:

Cu → Cu2+

0 +2

Cu → Cu2+ + 2e-

The half equation for the reduction of nitric acid to nitrogen dioxide is:

HNO3 → NO2

+5 +4

HNO3 + e- + H+ → NO2 + H2O

Make sure the electrons balance, so they will cancel out:

Cu → Cu2+ + 2e-

2HNO3 + 2e-  + 2H+ → 2NO2 + 2H2O N.B. Electrons must balance so x2

Cu + 2HNO3 + 2H+ → Cu2+ + 2NO2 + 2H2O

RA OA

***Task: From the given half-equations deduce the overall equation***

Ethanedioic acid [C2O42- or (COO)22-] reacting with manganate (VII) ions

|  |  |  |  |
| --- | --- | --- | --- |
| R | MnO4-  + 5e- + 8H+ | → | Mn2+ + 4H2O |

|  |  |  |  |
| --- | --- | --- | --- |
| O | C2O42- | → | 2CO2 + 2e- |

Overall equation:

|  |  |  |  |
| --- | --- | --- | --- |
| x2 | 2MnO4-  + 10e- + 16H+ | → | 2Mn2+ + 8H2O |

|  |  |  |  |
| --- | --- | --- | --- |
| x5 | 5C2O42- | → | 10CO2 + 10e- |

|  |  |  |  |
| --- | --- | --- | --- |
|  | 2MnO4-  + 5C2O42- + 16H+ | → | 2Mn2+ + 10CO2 + 8H2O |

***Starter: 9.1.3 – Half equations to overall equations***

***Sheet: Redox – equations***

***Sheet: Oxidation states & redox***

***Extension: Redox equations Q2***

***Sheet: Redox PPQ1-2***

***Exam questions: Oxford p134-135 Q1-5***