Redox Reactions

Redox reagents, equations, titrations, and electrolysis.

Index

Redox Equations

Redox reactions include reactions which involve the loss or gain of electrons.

The reactant giving away (donating) electrons is called the reducing agent (which is oxidised)

The reactant taking (accepting) electrons is called the oxidising agent (which is reduced)

Both oxidation and reduction happen simultaneously, however each is considered separately using ion-electron equations.

O.I.L. R.I.G. Oxidation is loss, reduction is gain of electrons e.g. $\frac{1}{2}O_2 + 2e^- \rightarrow O^{2-}$

$\frac{1}{2}Cl_2 + e^- \rightarrow Cl^-$

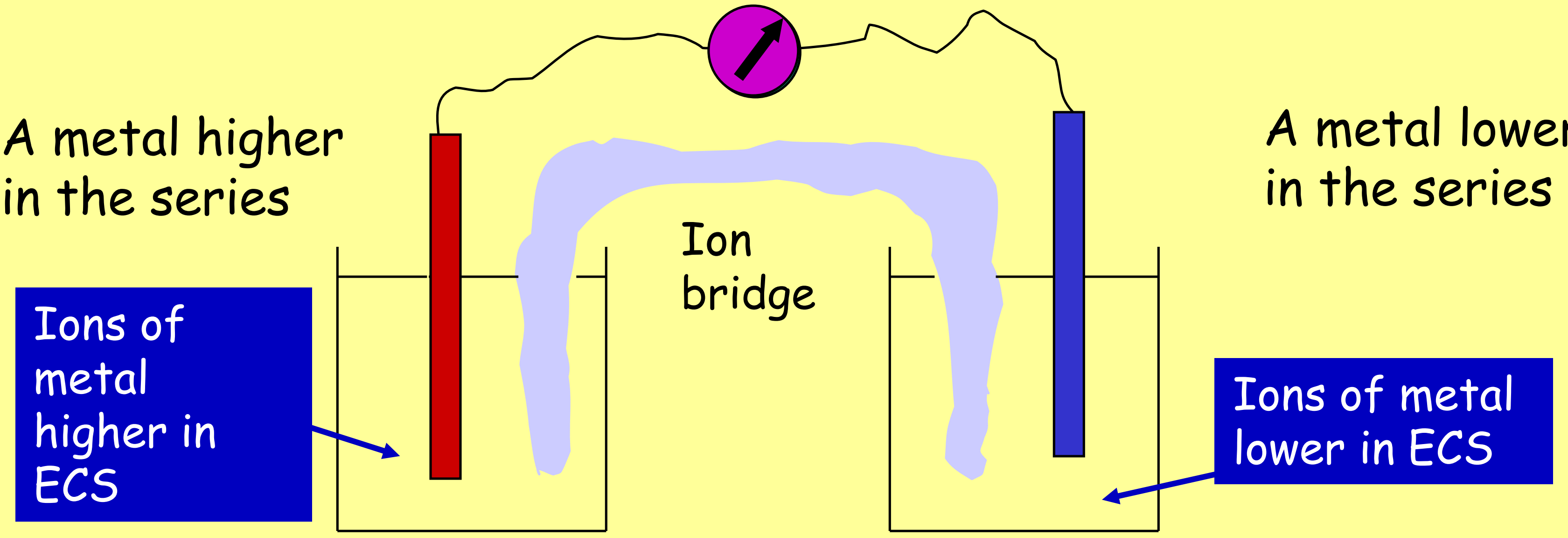
e.g.

$Mg \rightarrow Mg^{2+} + 2e^-$

$Al \rightarrow Al^{3+} + 3e^-$

- **Metals** on the **LHS** of the **Periodic Table** ionise by **electron loss** and are **called reducing agents**
- **Non-metals** on the **RHS** of the **Periodic Table** ionise by **electron gain** and are called **oxidising agents**

Cells and Redox



Metal atoms will be **oxidised.**
Metal atoms are the **reducing agent**.

Metal ions in solution will be **reduced**,
Metal ions are the **oxidising agent**.

e.g. $Mg \rightarrow Mg^{2+} + 2e^-$

E.g. $Cu^{2+} + 2e^- \rightarrow Cu$

Overall redox equation $\mathbf{Mg_{(s)} + Cu^{2+}_{(aq)} \rightarrow Mg^{2+}_{(aq)} + Cu_{(s)}}$

Cells and Redox

magnesium(s) + silver nitrate(aq) ➔ magnesium nitrate(aq) + silver(s).

The reducing agent in this reaction is the Mg as it will donate electrons to the silver ions .

The oxidising agent is the Ag^+ ions as they accept electrons from the Mg

Oxidation: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

Reduction: $2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

Half equations or ion-equations

$Mg_{(s)} + 2Ag^+_{(aq)} \rightarrow Mg^{2+}_{(aq)} + 2Ag_{(s)}$

Redox equation, electrons cancel out

Redox and the Electrochemical Series

E^o/V	Oxidising agents
-3.02v	$Li^+_{(aq)} + e \rightarrow Li_{(s)}$
-2.71v	$Na^+_{(aq)} + e \rightarrow Na_{(s)}$
-2.37v	$Mg^{2+}_{(aq)} + 2e \rightarrow Mg_{(s)}$
-0.13v	$Pb^{2+}_{(aq)} + 2e \rightarrow Pb_{(s)}$
0.00v	$2H^+_{(aq)} + 2e \rightarrow H_{2(g)}$
+0.34v	$Cu^{2+}_{(aq)} + 2e \rightarrow Cu_{(s)}$
+0.80v	$Ag^+_{(aq)} + e \rightarrow Ag_{(s)}$

Increasing powerful reducing agent (write the reaction **backwards**)

Hydrogen reference

Increasing powerful oxidising agent (write the reaction **as it appears**)

Considering the two ion-equations,

$Mg^{2+}_{(aq)} \rightarrow Mg_{(s)} + 2e^-$ **and** $Ag^+_{(aq)} + e^- \rightarrow Ag$,

Mg, being higher up the electrochemical series, would act as the reducing agent. (i.e. the ion-electron equation would be written backwards).
While Ag would be written as it appears in the electrochemical series.

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

Writing REDOX equations

Consider the reaction between sodium and water:

$$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + \tfrac{1}{2}H_{2(g)}$$

Consider how the ions are formed in this reaction

$$Na_{(s)} \rightarrow Na^+_{(aq)} + e^-$$

$$H_2O_{(l)} + e^- \rightarrow OH^-_{(aq)} + \tfrac{1}{2}H_{2(g)}$$

$$Na_{(s)} \rightarrow Na^+_{(aq)} + e^-$$

A sodium atom loses an electron

$$H_2O_{(l)} + e^- \rightarrow OH^-_{(aq)} + \tfrac{1}{2}H_{2(g)}$$

and, we could say that a water molecule must be accepting the electron

$Na_{(s)} \rightarrow Na^{+}_{(aq)} + e^{-}$

$H_2O_{(l)} + e^{-} \rightarrow OH^{-}_{(aq)} + \frac{1}{2}H_{2(g)}$

OIL

RIG

These are called **ion-electron equations** (or **ionic half equations**).

$$Na_{(s)} \rightarrow Na^{+}_{(aq)} + \cancel{e^-}$$

Electrons cancel!

$$H_2O_{(l)} + \cancel{e^-} \rightarrow OH^{-}_{(aq)} + \tfrac{1}{2}H_{2(g)}$$

Reduction and oxidation occur simultaneously. Adding the two equations together gives us the overall equation for a reaction.

$$Na_{(s)} + H_2O_{(l)} \rightarrow NaOH_{(aq)} + \tfrac{1}{2}H_{2(g)}$$

Balancing Redox equations

Most redox reaction you will come across will occur in neutral or acidic conditions.

1. Make sure there are the same number of atoms of each element being oxidised or reduce on each side of the half equation.

2. If there are any oxygen atoms present, balance them by adding water molecules to the other side of the half-equation.

3. If there are any hydrogen atoms present, balance them by adding hydrogen ions on the other side of the half-equation.

4. Make sure the half-reactions have the same overall charge on each side by adding electrons.

For basic solutions H atoms are balanced using H_2O and then the same number of OH^- ions to the opposite side to balance the oxygen atoms

1. Write down what you know....

sulphur dioxide is oxidised to sulphate ions

$$SO_{2(g)} \rightarrow SO_4{}^{2-}{}_{(aq)}$$

2. Balance the **oxygen** atoms by adding **water**

$$SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_4{}^{2-}{}_{(aq)}$$

3. Balance the **hydrogen atoms** by adding **hydrogen ions**

$$SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_4{}^{2-}{}_{(aq)} + 4H^+{}_{(aq)}$$

4. Balance the **charges** by adding **electrons**

$$SO_{2(g)} + 2H_2O_{(l)} \rightarrow SO_4{}^{2-}{}_{(aq)} + 4H^+{}_{(aq)} + 2e^-$$

charge is zero

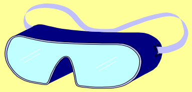
4 - and 4 + equals zero

Redox Titrations

Titration is a technique for measuring the concentration of a solution. A solution of known concentration is used to work out the unknown concentration of another solution.

Redox titrations involve solutions of reducing and oxidising agents.

At equivalence-point of a redox titration precisely enough electrons have been removed to oxidise all of the reducing agent.



Redox Titration

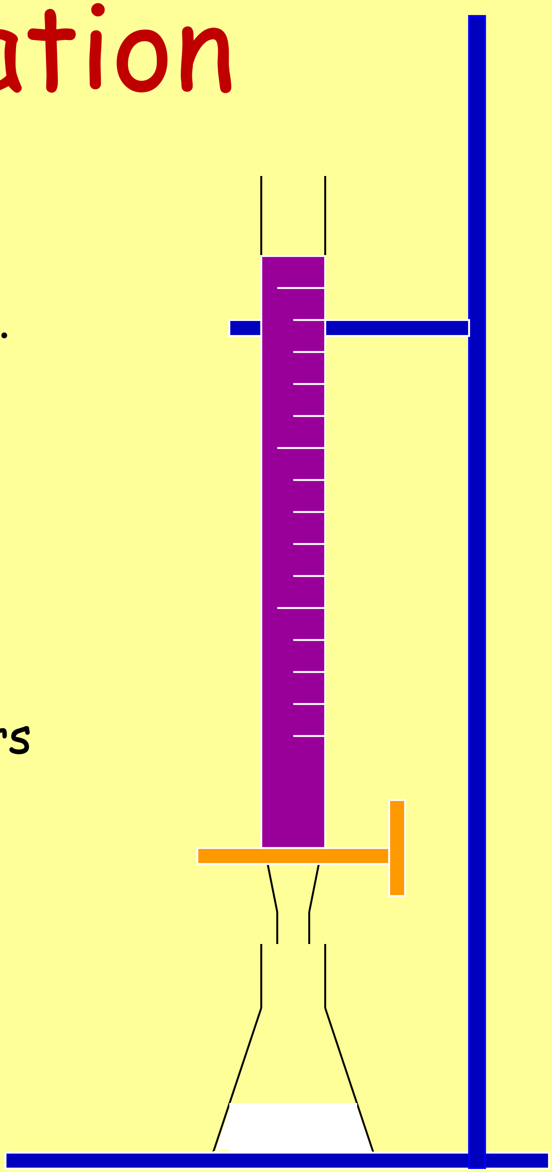
What to do:

1. Carefully fill the burette with **potassium permanganate** .

2. Carefully pipette exactly 20 ml of iron (II) sulphate into the conical flask.
Then add 20 ml 1 mol l^{-1} H_2SO_4

3. Add the permanganate until a permanent purple colour appears in the conical flask.

4. A rough titration is done first to give a rough equivalence-point (end-point), then repeated more accurately to give concordant results.



Redox Titrations

$5\,Fe^{2+}(aq) + 8H^+(aq) + MnO_4^-(aq) \rightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4H_2O(l)$

purple   colourless

Use a standard solution of potassium permanganate to find out the unknown concentration of an iron (II) sulphate solution

$n_y = 5$ $y = [Fe^{2+}{}_{(aq)}]$	$n_x = 1$ $x = [MnO_4^-{}_{(aq)}]$

$$\frac{V_y \times C_y}{n_y} = \frac{V_x \times C_x}{n_x}$$

Or $$C_y = \frac{V_x \times C_x \times n_y}{V_y \times n_x}$$

Redox Titrations, Vitamin C

$I_2 (aq) + 2e^- \rightarrow 2I^- (aq)$ reduction

$C_6H_8O_6 \rightarrow C_6H_6O_6 + 2H^+ (aq) + 2e^-$ oxidation

$I_2 (aq) + C_6H_8O_6 \rightarrow C_6H_6O_6 + 2H^+ + 2I^- (aq)$

Blue/Black (in the presence of starch)

colourless

Iodine, those concentration is known (in the burette) acts as an oxidising agent.

Vitamin C, the unknown concentration (in the conical flask) is a reducing agent.

Starch is added to show when the end-point is reached.

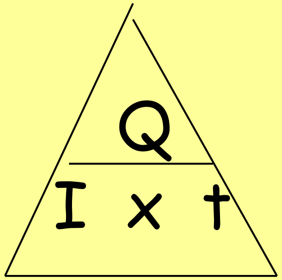
$$\frac{V_x \times C_x}{n_x} = \frac{V_y \times C_y}{n_y}$$

Electrolysis

Faraday was the first person to measure the amount of electrical charge needed to deposit a certain amount of substance at an electrode.

Amount of electrical charge (electrons) → (electrolysis) → Mass of substance deposited

Electrical charge is the amount of electrons



Electrolysis

Current is the flow of an electrical charge

The amount or quantity of charge (Q) is measured in Coulombs (C)

Quantity of charge = current x time

$$Q = I \times t$$

96,500 coulombs is called 1 Faraday (F).

The number of coulombs required to deposit 1 mole of atoms or molecules of an element is 96,500 x n. (F x n) n being either 1,2,3 or 4.

The multiplying factor n, can be equated to the number of electrons associated with the production of one atom or molecule of the element.

Electrolysis

Electrode reaction	Value of n (number of coulombs required to produce 1 mol of atoms)
$Na^{+} + e^{-} \Rightarrow Na$	1
$2H^{+} + 2e^{-} \Rightarrow H_2$	2
$Mg^{2+} + 2e^{-} \Rightarrow Mg$	2
$Al^{3+} + 3e^{-} \Rightarrow Al$	3
$4OH^{-}_{aq} \Rightarrow 2H_2O + O_2 + 4e^{-}$	4

96,500 coulombs = 1 mole of electrons

Electrolysis and Hydrogen

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

To produce 1 mole of H_2, 2 moles of electrons are needed.

So to produce 1 mole of H_2 , 96500 x 2 C of charge is needed.

It is possible to confirm that 96500 x 2 C of charge are needed to produce 1 mole of H_2 gas by electrolysing. The volume of hydrogen gas collected at the cathode is measured and converted to moles using the gases molar volume.
(The molar volume = 24 litres).
So knowing the volume of gas collected, you can work out the number of moles of gas collected.

Gases and Electrolysis

The mass or volume of an element discharged by electrolysis can be calculated from the quantity of electricity used and vice-versa.

Example: A solution of HCl is electrolysed. What current is needed to produce 2.4 litres of H_2 gas in 16min 5 sec? Molar volume = 24 l mol^{-1}

Since $2H^+ + 2e^- \rightarrow$ 1 mole of gas requires 2 moles of electrons.

i.e. 96500 x 2 C of charge is needed to produce 1 mole of gas

Since 2.4 litres is 0.1 mole of gas, so (96500 x 2) x 0.1 C is needed

$Q = I \times t$

So $I = Q/t$

$(96500 \times 2) \times 0.1 \;/\; (16 * 60) + 5$

Ans: **20 A**