

Redox, Group 2 & Group 7

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Special thanks to Ms Williams and Ms Matrella for their constant support and inspiration

Please note that these guides are a collation of my personal notes, teachers' notes, chemistry books, and websites such as chemguide, chemsheets, chemwiki and wikipedia.

Redox

OIL - **Oxidation** Is **Loss** of electrons

RIG - **Reduction** Is **Gain** of electrons

Oxidation and reduction can be seen as movement up or down a scale of oxidation states

Also note that **Oxidation is gain in oxygen or loss of hydrogen**. **Reduction is loss of oxygen or gain of hydrogen**. However these conditions may not necessary apply. If an element gains oxygen but has its oxidation state reduced, it is said to be reduced.

Rules

Here are the rules for determining oxidation states of elements:

Element	Oxidation State
GROUP 1	+1
GROUP 2	+2
F	-1
Cl	-1 (Except in ClF and Cl ₂ O)
O	-2 (Except in F ₂ O and H ₂ O ₂)
H	+1
H in Metal Hydride	-1

Writing Half Ionic Equations

1. Write an equation for the species
2. Work out oxidation states for the element on both sides of the equation
3. Balance O with H₂O
4. Balance H with H⁺
5. Balance charges with e⁻

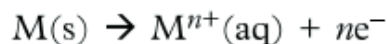
Combining Half Equations

If the electrons on both half equations balance out then add the equations and cancel out the e⁻. If not, use an appropriate scalar.

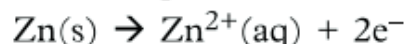
Common Reducing Agents

These will react to **reduce another species** and are **themselves oxidized**. They do so by giving away electrons.

Metals



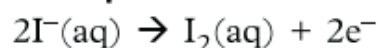
For example:



Iron(II) salts



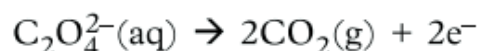
Acidified potassium iodide



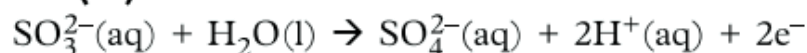
Thiosulfate



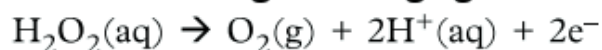
Ethanedioic acid and ethanedioates



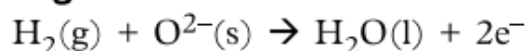
Sulfuric(IV) acid



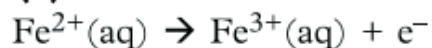
Hydrogen peroxide in the presence of an acid and absence of a strong oxidising agent



Hydrogen



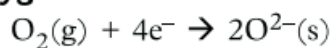
Iron(II) salts



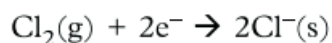
Common Oxidizing Agents

These will react to **oxidize another species** and are **themselves reduced**. They do so by gaining electrons.

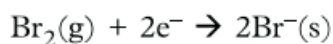
Oxygen



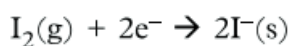
Chlorine



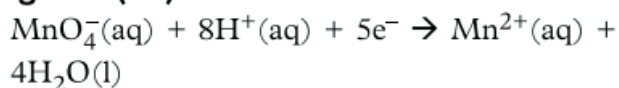
Bromine



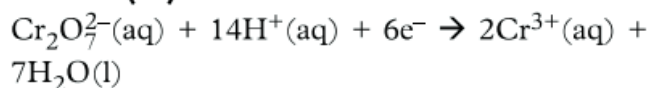
Iodine



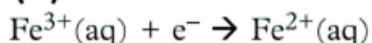
Manganate(VII) in acid solution



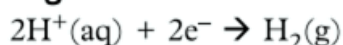
Dichromate(VI) in acid solution



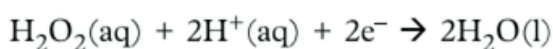
Iron(III) salts



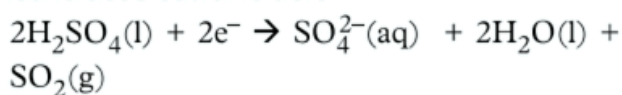
Hydrogen ions



Hydrogen peroxide in absence of another oxidising agent



Concentrated sulfuric acid

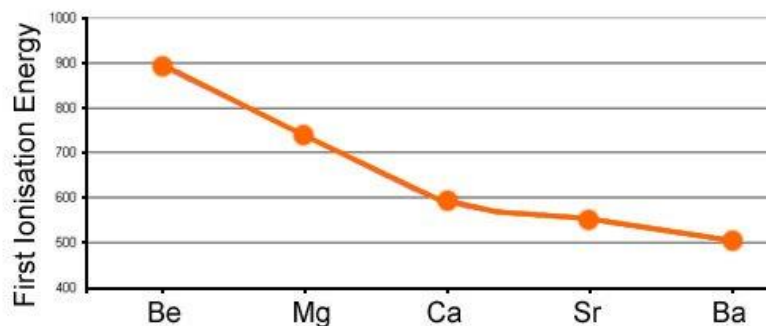


Group 2

Ionisation Energy (IE)

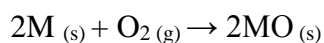
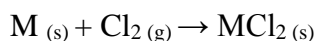
IE decreases down the group because more full electron shells are present each time. Hence more shielding effect will take place. Also, the atomic radius will increase. Both of these will decrease the effective nuclear charge which in turn decreases the attraction between the

nucleus and the valency electron. Therefore the outermost electron is more easily removed. Hence reactivity increases down the group.



Reactions

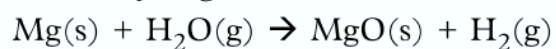
M = Group 2 Metal



Element	Reaction with water or steam
Beryllium	No reaction due to a thick oxide layer on the surface of the metal
Magnesium	Very, very slow reaction with cold water – rapid reaction with steam
Calcium	Reaction with cold water to produce hydrogen – reactivity increases down the group
Strontium	
Barium	

table 2.5.4 The reaction of the group 2 metals with water.

Magnesium reacts with steam to form magnesium oxide and hydrogen:



Calcium, strontium and barium react to form the hydroxide and hydrogen, for example:



In each case the hydroxide forms an alkaline solution.

Solubility

	nitrates	chlorides	sulphates	hydroxides
magnesium				
calcium	all soluble	all soluble	solubility decreases down group	solubility increases down group
strontium				
barium				

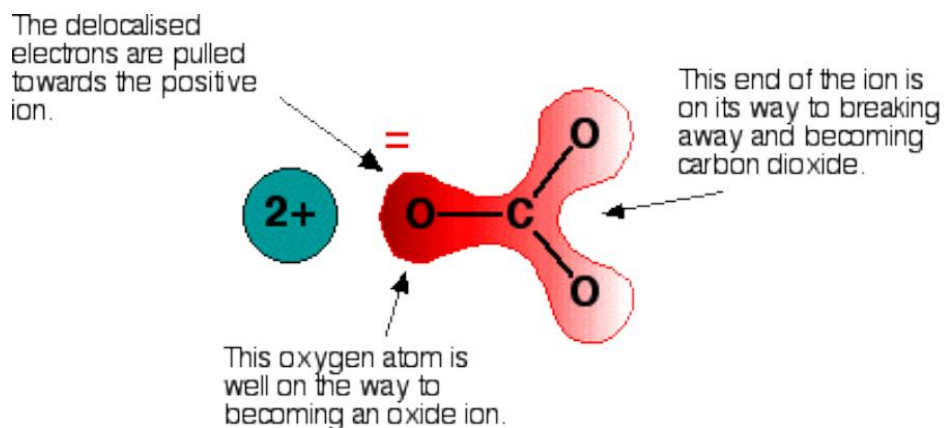
A more generalised picture for solubility:

Symbol	Soluble	Insoluble
Na/K/NO ₃ /NH ₄ ⁺	ALL	NO EXCEPTIONS
SO ₄ ²⁻	MOST	Pb/Ba/Ca
Cl ⁻	MOST	Pb/Ag
O ²⁻ /OH ⁻ /CO ₃ ²⁻	Na/K/ Ba ²⁺ , Sr ²⁺ , Ca ²⁺ are considered slightly soluble.	THE REST

Thermal Stability

Group 2 Carbonates

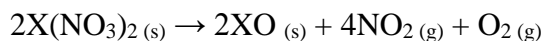
The carbonates become more stable to heat as you go down the Group. The smaller the positive ion is, the higher the charge density, and the greater effect it will have on the carbonate ion. As the ionic radius of the cations increases as you go down Group 2, they have less effect on the carbonate ions near them. This means that you have to heat the compound more in order to help the carbon dioxide to break free and leave the metal oxide.



Group 1 Carbonates

Group 1 carbonates are **more** stable to heat than Group 2 carbonates. Only Li_2CO_3 decomposes when heated with a bunsen burner. This is because the charge on cations in Group 1 is only +1 and the polarising power is weaker.

Group 2 Nitrates



Like the carbonates, the nitrates become more stable to heat as you go down Group 2.

Group 1 Nitrates

In Group 1, only LiNO_3 breaks down to lithium oxide, NO_2 and O_2 when heated. The rest of Group 1 nitrates break down to form the nitrite and O_2 .



Note: The nitrates are easier to break down than the carbonates.

Flame Tests

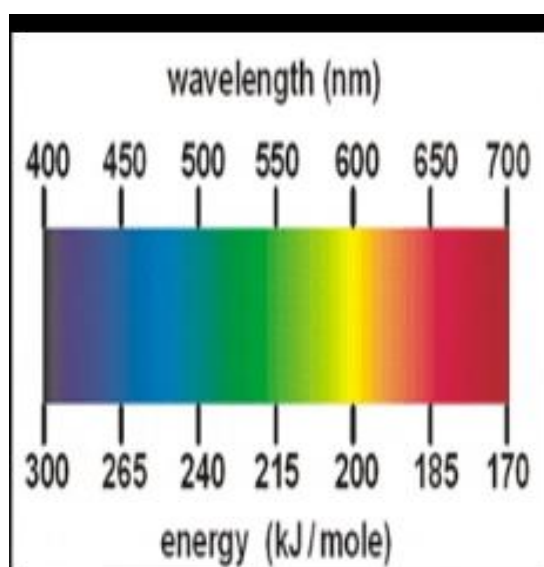
Dip a clean nichrome or platinum wire into a sample that has been mixed with a few drops of concentrated hydrochloric acid. This is to remove impurities. Hold the wire in a hot blue bunsen flame to see the colours.

How?

Flame colours are produced from the movement of electrons in the metal ions. When heated, electrons gain energy and can jump into empty orbitals at higher levels. We say the atom is excited. When an electron drops back to its ground state, the atom emits a photon of light. Each element has its own characteristic wavelength of light, e.g. 800 nm for lithium ions.

The spacing between energy levels in an atom determines the sizes of the transitions that occur, and thus the energy and wavelengths of the collection of photons emitted. Larger transition – higher energy photon released (shorter wavelength). Small transition – lower energy photon released (longer wavelength).

Smaller transitions tend to occur in the top of the table. From there we say that going down a group, a blue shift tends to occur in Group 1. As for Group 2, the pattern is harder to explain.



Cation	Flame colour
Lithium	Red
Sodium	Yellow
Potassium	Lilac
Magnesium	No colour
Calcium	(Yellow-) red
Strontium	Red
Barium	(Pale) green

Group 7

Strong oxidizing agents due to a very high electronegative character. They only need one electron to be stable. The oxidizing ability decreases down a group since the valence electron is further away from nucleus, more shielded, etc...

Going down a group, the reactivity increases. The colour gets darker and boiling temperature increases: F(pale yellow gas) --> Cl (greenish yellow gas) --> Br (dark red liquid) --> I (shiny/greyish-black crystalline solid).

Reactions

Displacement

As the reactivity decreases down a group, the top elements will displace the bottom ones from their compounds. Sometimes the colour change can be confused, for example Bromine molecules for Iodine molecules. Hence an organic solvent is used to distinguish between

them. In the image, the order is Cl_2 , Br_2 , I_2 . The top layer is organic solvent + halogen while the bottom layer is aqueous solution. Another test for presence of iodine molecules is adding starch. If I_2 is present, starch goes from colourless to blue/black. This test is essential when carrying out iodine/thiosulfate titration as it sharpens the end point of the reaction.

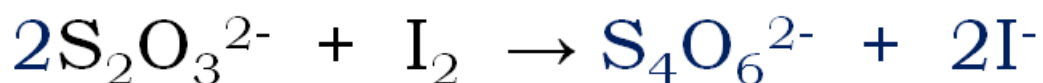
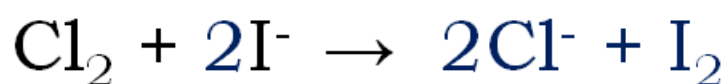
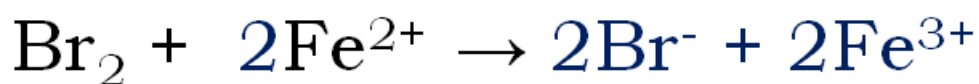
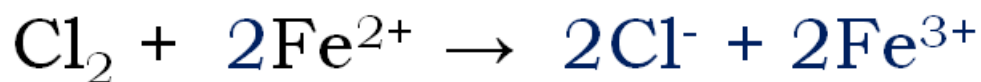
	$\text{Cl}_2(\text{aq})$	$\text{Br}_2(\text{aq})$	$\text{I}_2(\text{aq})$
$\text{Cl}^-(\text{aq})$		Stays yellow solution (no reaction)	Stays brown solution (no reaction)
$\text{Br}^-(\text{aq})$	Yellow solution forms (Br_2 forms) $\text{Cl}_2 + 2 \text{Br}^- \rightarrow 2 \text{Cl}^- + \text{Br}_2$		Stays brown solution (no reaction)
$\text{I}^-(\text{aq})$	Brown solution forms (I_2 forms) $\text{Cl}_2 + 2 \text{I}^- \rightarrow 2 \text{Cl}^- + \text{I}_2$	Brown solution forms (I_2 forms) $\text{Br}_2 + 2 \text{I}^- \rightarrow 2 \text{Br}^- + \text{I}_2$	



Here, the top layer is organic and the bottom is aqueous...The far left is Cl, Br then I

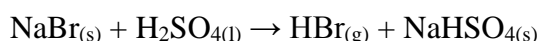
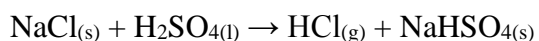
Disproportionation Reactions

In the 3rd reaction, you can see Cl being oxidized to be Cl^- and also reduced to Cl^+ . This is called a disproportionation reactions where a species is both oxidized and reduced at the same time.

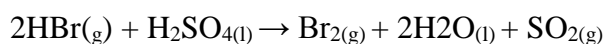


Testing for the halides and solubility in $\text{NH}_3(\text{aq})$

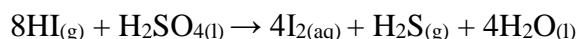
Halide ion	Action of $\text{AgNO}_3(\text{aq})$	Action of dilute $\text{NH}_3(\text{aq})$	Action of conc. $\text{NH}_3(\text{aq})$
$\text{F}^-(\text{aq})$	No precipitate		
$\text{Cl}^-(\text{aq})$	$\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$ white ppt	Soluble $\text{AgCl}(\text{s}) + 2 \text{NH}_3(\text{aq}) \rightarrow$ $[\text{Ag}(\text{NH}_3)_2]^+(\text{aq}) + \text{Cl}^-(\text{aq})$	Soluble $\text{AgCl}(\text{s}) + 2 \text{NH}_3(\text{aq}) \rightarrow$ $[\text{Ag}(\text{NH}_3)_2]^+(\text{aq}) + \text{Cl}^-(\text{aq})$
$\text{Br}^-(\text{aq})$	$\text{Ag}^+(\text{aq}) + \text{Br}^-(\text{aq}) \rightarrow \text{AgBr}(\text{s})$ cream ppt	Insoluble	Soluble $\text{AgBr}(\text{s}) + 2 \text{NH}_3(\text{aq}) \rightarrow$ $[\text{Ag}(\text{NH}_3)_2]^+(\text{aq}) + \text{Br}^-(\text{aq})$
$\text{I}^-(\text{aq})$	$\text{Ag}^+(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{AgI}(\text{s})$ yellow ppt	Insoluble	Insoluble

Reactions Of Halides With Concentrated H_2SO_4 

But we can distinguish between NaBr and NaCl using conc. sulphuric acid because for bromides, the H_2SO_4 acts as an oxidising agent on the HBr formed to form Br_2 :



And iodides get oxidised even more. The sulphuric acid is reduced to hydrogen sulphide as the HI is a stronger reducing agent than HBr.



halide	products	observation	reaction type	equation
Cl ⁻	HCl	steamy fumes	acid-base	$\text{NaCl} + \text{H}_2\text{SO}_4 \rightarrow \text{NaHSO}_4 + \text{HCl}$
Br ⁻	HBr	steamy fumes	acid-base	$\text{NaBr} + \text{H}_2\text{SO}_4 \rightarrow \text{NaHSO}_4 + \text{HBr}$
	Br ₂	brown fumes	oxidation of Br ⁻	$2 \text{Br}^- + \text{H}_2\text{SO}_4 + 2 \text{H}^+ \rightarrow \text{Br}_2 + \text{SO}_2 + 2 \text{H}_2\text{O}$
	SO ₂	colourless gas	reduction of H ₂ SO ₄	$2 \text{Br}^- + \text{H}_2\text{SO}_4 + 2 \text{H}^+ \rightarrow \text{Br}_2 + \text{SO}_2 + 2 \text{H}_2\text{O}$
I ⁻	HI	steamy fumes	acid-base	$\text{NaI} + \text{H}_2\text{SO}_4 \rightarrow \text{NaHSO}_4 + \text{HI}$
	I ₂	purple fumes	oxidation of I ⁻	$2 \text{I}^- + \text{H}_2\text{SO}_4 + 2 \text{H}^+ \rightarrow \text{I}_2 + \text{SO}_2 + 2 \text{H}_2\text{O}$
	SO ₂	colourless gas	reduction of H ₂ SO ₄	$2 \text{I}^- + \text{H}_2\text{SO}_4 + 2 \text{H}^+ \rightarrow \text{I}_2 + \text{SO}_2 + 2 \text{H}_2\text{O}$
	S	yellow solid	reduction of H ₂ SO ₄	$6 \text{I}^- + \text{H}_2\text{SO}_4 + 6 \text{H}^+ \rightarrow 3 \text{I}_2 + \text{S} + 4 \text{H}_2\text{O}$
	H ₂ S	gas (bad egg smell)	reduction of H ₂ SO ₄	$8 \text{I}^- + \text{H}_2\text{SO}_4 + 8 \text{H}^+ \rightarrow 4 \text{I}_2 + \text{H}_2\text{S} + 4 \text{H}_2\text{O}$