Redox worksheet

Instructions:

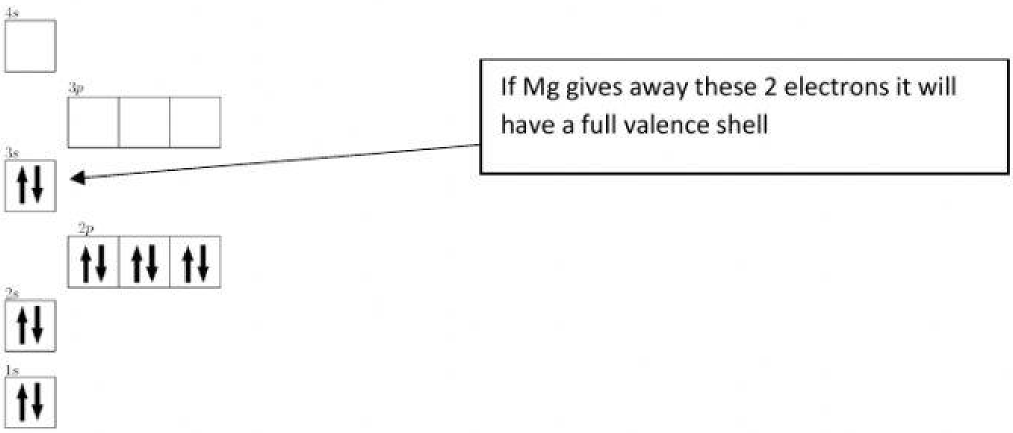
- ✓ **Fill in the missing words wherever space is provided.**
- ✓ **Fill numbers in as words**
- ✓ **When a charge is -1, just type –**
- ✓ **Exponents can just be left as regular numbers/symbols eg: Na+1**

Remember that all the elements on the periodic table are neutral.

They only become charged when they bond with another element.

Group 1,2,3 and transition elements really want to give off electrons. They have valence shells that are almost full and if they just give off 1,2 or 3 electrons they can become full.

Eg Mg



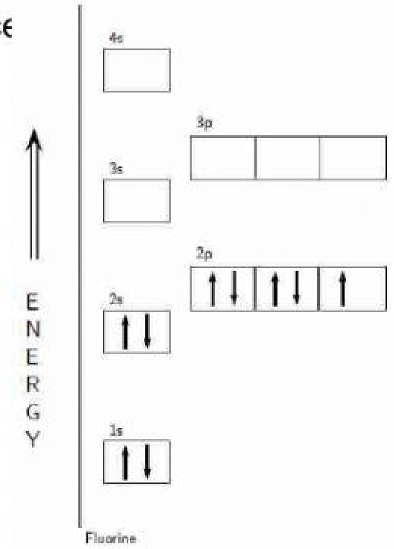
Magnesium is so close to having a full valance shell- all it needs to do is give away electrons.

Thus metals want to give off electrons (choose from: oxidise / reduce)

Non-metals

Non-metals on the other hand desperately want to gain electrons to have a full valence

Eg



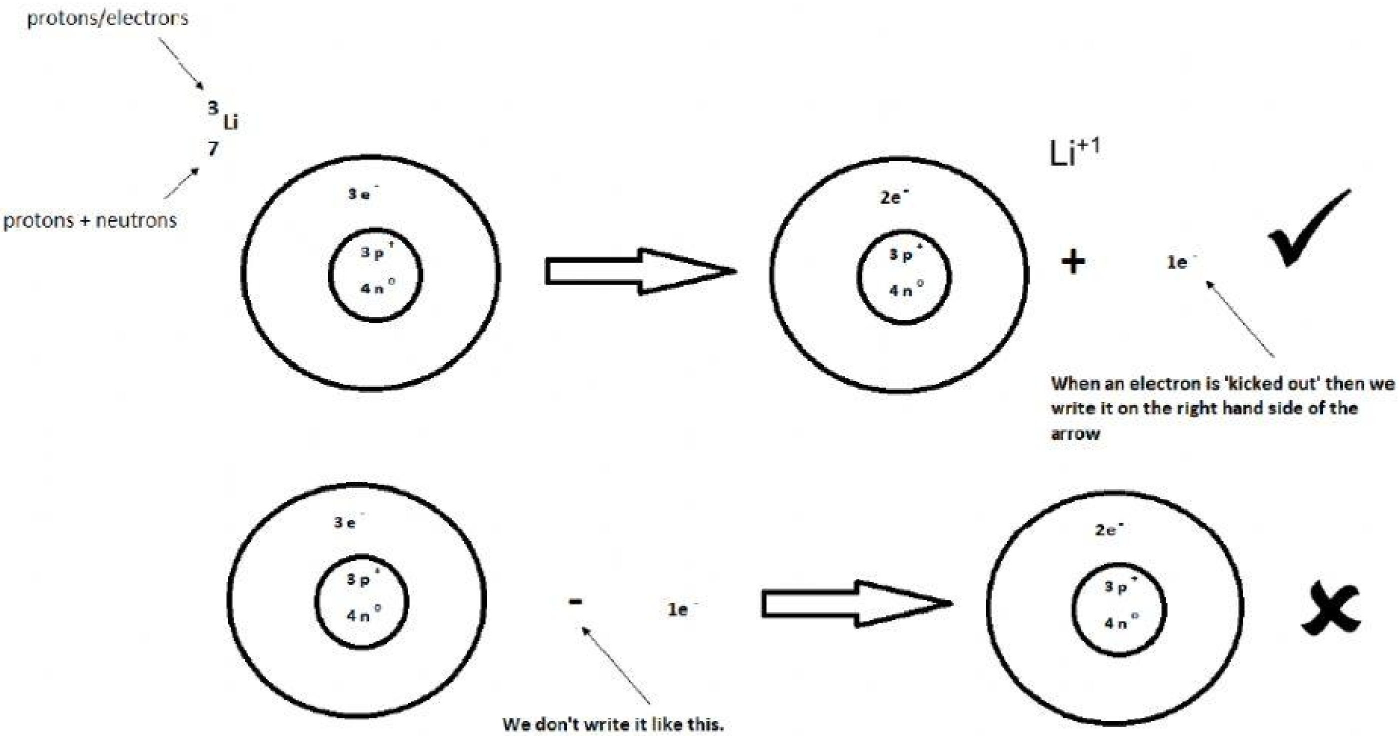
If fluorine can just gain   electron it the it would have a full valance shell.

Thus non metals want to gain electrons and thus oxidise   /   reduce

Now let's look at how you write it down when **group 1** elements give off electrons

Group 1

Li $\rightarrow$ Li^{+1} + $1e^-$



Na $\rightarrow$ Na^{+1} + $1e^-$

Group 2

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

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

Group 3

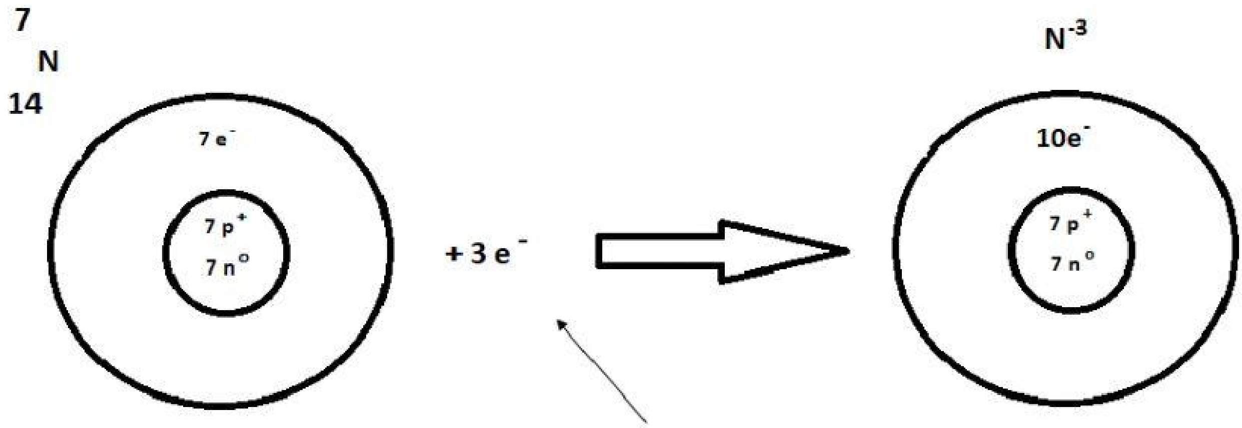
Al$\rightarrow$ Al^{+3} + $3e^-$

The metals thus oxidised and lost electrons

Non-metals

Remember that these want to gain electrons and reduce

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



If electron/s are gained by the element then you write it on the left hand side of the arrow

Group 6

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

Group 7

$F + 1e^- \rightarrow F^{-1}$

* In reality these elements are diatomic – so we need to multiply them by 2

Diatomic elements are: H, N, O, F, Cl, Br and I

$N_2 + 6e^- \rightarrow 2\,N^{-3}$

$O_2 + 4e^- \rightarrow 2O^{-2}$

$F_2 + 2e^- \rightarrow 2F^{-1}$

When the element is charged: we no longer write the 2 as a subscript, but rather as a coefficient

The non-metals reduced and gained electrons

Complete reactions

Now let's put the metal and non-metal together in a reaction

Example 1)

$Na + Cl_2 \rightarrow NaCl$

First things first – write the charge(also known as the oxidation no) above each element

$$\overset{0}{Na} + \overset{0}{Cl_2} \rightarrow \overset{+1}{Na}\overset{-1}{Cl}$$

If we break this up, then Na is becoming more positively charged and is losing one electron, thus it oxidises

Oxidation: $Na \rightarrow Na^{+1} + 1e^-$

And Chlorine is becoming more negatively charged and is gaining two electrons (since it is diatomic), thus it reduces

Reduction: $Cl_2 + 2e^- \rightarrow Cl^{-2}$

Note that there can't be an imbalance of electrons- thus is one element gives off 1 and the other gains 2, we have a bit of a problem.

To solve this, we multiply the electrons in the Na reaction by 2, but since you can't just multiply the electrons by 2, the entire reaction is multiplied by 2

$Na \rightarrow Na^{+1} + 1e^-$ (x2)

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

Redox: $2Na + Cl_2 \cancel{+ 2e^-} \rightarrow 2Na^{+1} + 2Cl^{-1} \cancel{+ 2e^-}$

(a redox reaction is one is which there is reduction and oxidation

Thus the oxidation and reduction reaction is put together and the electrons that are on the left and on the right of the arrow is cancelled.)

Watch this video on oxidation and reduction first before carrying on

https://www.youtube.com/watch?v=5rtJdjas-mY

Example 2

$Mg + Cl_2 \rightarrow MgCl_2$

Chlorine's charge is **-1**, not -2

It is multiplied by 2 because of the subscript

0 0 +2 -1(2)

$Mg + Cl_2 \rightarrow MgCl_2$

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

Red: $Cl_2 + 2e^- \rightarrow 2\ Cl^{-2}$

*check if the electrons are the same – if they are, then you can proceed to the redox reaction

Redox: $Mg + Cl_2 \rightarrow Mg^{+2} + 2Cl^{-1}$

Example 3

$Mg + O_2 \rightarrow MgO$

$$\overset{0}{Mg} + \overset{0}{O_2} \rightarrow \overset{+2}{Mg}\overset{-2}{O}$$

Oxi: $Mg \rightarrow Mg^{+2} + 2e^{-}$ (x2)

Red: $O_2 + 4e^{-} \rightarrow 2O^{-2}$

Redox: $2Mg + O_2 \rightarrow Mg^{+2} + 2O^{-2}$

Exercise 1: complete the following

1) $Ca + Cl_2 \rightarrow CaCl_2$

Oxi: $Ca \rightarrow \quad +$

Red: $Cl_2 + 2e^{-} \rightarrow$

Redox: $Ca + Cl_2 \rightarrow \quad +$

2) $Al + O_2 \rightarrow$

Oxi: $\rightarrow \quad +$

Red: $+ \quad \rightarrow$

Redox: $+ \quad \rightarrow \quad +$

3) $Al + Br_2 \rightarrow$

Oxi: $\rightarrow \quad +$

Red: $+ \quad \rightarrow$

Redox: $+ \quad \rightarrow \quad +$

Oxidising agent: A substance which undergoes oxidation / A substance which causes oxidation (itself is reduced)

Reducing agent: A substance which undergoes oxidation / A substance which causes oxidation (itself is oxidised)

Examples with spectator ions:

1) $Mg + ZnCl_2 \rightarrow Zn + MgCl_2$

step 1: figure out the oxidation numbers

Remember you need to know 2 transition elements' charges:

Zn^{+2}

Ag^{+1}

0 +2 -1(2) 0 +2 -1(2)

$Mg + ZnCl_2 \rightarrow Zn + MgCl_2$

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

red: $Zn^{+2} + 2e^- \rightarrow Zn$

redox: $Mg + Zn^{+2} \rightarrow Mg^{+2} + Zn$

Spectator ion: Cl^-

Oxidising agent: Zn^{+2} reducing agent: Mg

This the ion that does not oxidise or reduce

It is just a 'spectator' in the reaction

2) $Zn + Cu(NO_3)_2 \rightarrow Zn(NO_3)_2 + Cu$

0 $\boxed{}$ -1(2) +2 -1(2) 0

$Zn + Cu(NO_3)_2 \rightarrow Zn(NO_3)_2 + Cu$

It is going to become really important at this point to know your list of compound ions.

Now you need to reverse cross multiply in order to determine the charge of the transition elements

+2 -1

$Cu(NO_3)_2$

oxi: $Zn \rightarrow Zn^{2+} + 2e^-$

<u>red:</u> $Cu^{2+} + 2e^- \rightarrow Cu$

redox: $Zn + Cu^{+2} \rightarrow Cu + Zn^{+2}$

spectator ion:

oxidising agent: reducing agent:

Exercise 2:

1) $Cu + AgNO_3 \rightarrow Cu(NO_3)_2 + Ag$

Oxi: $\rightarrow$ +

Red: + $\rightarrow$

Redox: + $\rightarrow$ +

spectator ion:

oxidising agent: reducing agent:

2) $Pb + Cu(NO_3)_2 \rightarrow Pb(NO_3)_2 + Cu$

Oxi: $\rightarrow$ +

Red: + $\rightarrow$

Redox: + $\rightarrow$ +

spectator ion:

oxidising agent: reducing agent:

3) $Ni + Sn(NO_3)_2 \rightarrow Ni(NO_3)_2 + Sn$

Oxi: $\rightarrow$ +

Red: + $\rightarrow$

Redox: + $\rightarrow$ +

spectator ion:

oxidising agent: reducing agent:

4) $Fe + I_2 \rightarrow Fe^{3+} + 2I^-$

Oxi: $\rightarrow$ +

Red: + $\rightarrow$

Redox: + $\rightarrow$ +

spectator ion: none

oxidising agent: reducing agent:

You will notice from the above reactions that it is not always the metal that gives off electrons and non-metals that gain in each reaction. The reality is that 2 metals could react with each other and one will still oxidise and the other reduce.

How then do we determine which one will oxidise or reduce, if they don't give up a full reaction?

The reality is that not only transition elements can have multiple charges/oxidation numbers. Many other elements, such as nitrogen, sulphur, oxygen etc can have multiple charges.

Watch this video on oxidation numbers before you start the next exercise

https://www.youtube.com/watch?v=zpCy7KwqDO8