OXIDATION STATES

Used to
- tell if oxidation or reduction has taken place
- work out what has been oxidised and/or reduced
- construct half equations and balance redox equations

Atoms / simple ions **The number of electrons which must be added or removed to become neutral**

atoms	Na in Na	=	0	*neutral already ... no need to add any electrons*
cations	Na in Na^+	=	+1	*need to add 1 electron to make Na^+ neutral*
anions	Cl in Cl^-	=	−1	*need to take 1 electron away to make Cl^- neutral*

Q.1 ***What is the oxidation state of the elements in ?***

a) N ***b) Fe^{3+}***

c) S^{2-} ***d) Cu***

e) Cu^{2+} ***f) Cu^+***

Molecules **Sum of oxidation states adds up to zero**

Elements H in H_2 = 0

Compounds C in CO_2 = +4 and O = -2 [i.e. +4 and 2(-2) = 0]

- *CO_2 is neutral, so the sum of the oxidation states must be zero*
- *one element must have a positive OS, the other must be negative*
- *the more electronegative species will have the negative value*
- ***electronegativity increases across a period and decreases down a group***
- *O is further to the right in the periodic table so it has the negative value (-2)*
- *C is to the left so it has the positive value (+4)*
- *one needs two O's at -2 each to balance one C at +4*

Complex ions **Sum of oxidation states adds up to the charge on the ion**

in SO_4^{2-} S = +6 , O = -2 [i.e. +6 + 4(-2) = -2] ; therefore the ion has a 2- charge

Example *What is the oxidation state (O.S.) of Mn in MnO_4^- ?*

- *the O.S. of oxygen in most compounds is -2*
- *there are 4 O's so the sum of the O.S.'s = -8*
- *the overall charge on the ion is -1, ∴ the sum of all the O.S.'s must add up to -1*
- *the O.S. of Mn plus the sum of the O.S.'s of the four O's must equal -1*
- *therefore the O.S. of Manganese in MnO_4^- = +7*

WHICH OXIDATION STATE ?

- elements can exist in more than one oxidation state
- certain elements can be used as benchmarks

HYDROGEN	**(+1)**	except	0	atom (H) and molecule (H_2)
			-1	hydride ion, H^- [in sodium hydride, NaH]
OXYGEN	**(-2)**	except	0	atom (O) and molecule (O_2)
			-1	in hydrogen peroxide, H_2O_2
			+2	in F_2O
FLUORINE	**(-1)**	except	0	atom (F) and molecule (F_2)

Metals
- have positive values in compounds
- value is usually that of the Group Number — *Al is +3*
- values can go go no higher than the Group No. — *Mn can be +2,+4,+6,+7*

Non metals
- mostly negative based on their usual ion — *Cl is usually -1*
- can have values up to their Group No. — *Cl can be +1, +3, +5, +7*

- to avoid ambiguity, the oxidation state is often included in the name of a species

e.g. *manganese(IV) oxide shows Mn is in the +4 oxidation state in MnO_2*
sulphur(VI) oxide for SO_3
dichromate(VI) for $Cr_2O_7^{2-}$
phosphorus(V) chloride for PCl_5.

Q.2 ***What is the theoretical maximum oxidation state of the following elements ?***

Na P Ba Pb S Mn Cr

What will be the usual and maximum oxidation state in compounds of ?

	Li	***Br***	***Sr***	***O***	***B***	***N***
USUAL						
MAXIMUM						

Q.3 ***Give the oxidation state of the element other than O,H or F in***

SO_2	NH_3	NO_2	NH_4^+	IF_7	Cl_2O_7
MnO_4^{2-}	NO_3^-	NO_2^-	SO_3^{2-}	$S_2O_3^{2-}$	$S_4O_6^{2-}$

What is odd about the value of the oxidation state of S in $S_4O_6^{2-}$?
Can it have such a value ? Can you provide a suitable explanation ?

Q.4 ***What is the oxidation state of each element in the following compounds ?***

CH_4	PCl_3	NCl_3	CS_2	ICl_5	BrF_3
$MgCl_2$	H_3PO_4	NH_4Cl	H_2SO_4	$MgCO_3$	$SOCl_2$

REDOX REACTIONS

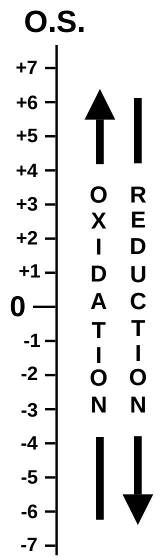
Redox When reduction and oxidation take place

Oxidation Removal of electrons; species will get less negative / more positive

Reduction Gain of electrons; species will become more negative / less positive

REDUCTION in O.S. Species has been REDUCED
e.g. Cl is reduced to Cl^- (0 to -1)

INCREASE in O.S. Species has been OXIDISED
e.g. Na is oxidised to Na^+ (0 to +1)



OIL RIG **O**xidation **I**s the **L**oss
Reduction **I**s the **G**ain of electrons

Q.5 ***Classify the following (unbalanced) changes as oxidation, reduction or neither.***

a) $Mg \longrightarrow Mg^{2+}$ ***b)*** $O^{2-} \longrightarrow O$ ***c)*** $Al^{3+} \longrightarrow Al$
d) $Fe^{3+} \longrightarrow Fe^{2+}$ ***e)*** $Ti^{3+} \longrightarrow Ti^{4+}$ ***f)*** $2Q \longrightarrow Q_2$

Q.6 ***What change takes place in the oxidation state of the underlined element ?***
Classify the change as oxidation, reduction or neither.

a) $\underline{N}O_3^- \longrightarrow NO$ ***b)*** $H\underline{N}O_3 \longrightarrow N_2O$ ***c)*** $\underline{C}H_4 \longrightarrow CO$
d) $\underline{Mn}O_2 \longrightarrow Mn^{2+}$ ***e)*** $\underline{Cr}_2O_7^{2-} \longrightarrow Cr^{3+}$ ***f)*** $\underline{S}O_3^{2-} \longrightarrow SO_4^{2-}$
g) $\underline{Cr}_2O_7^{2-} \longrightarrow CrO_4^{2-}$ ***h)*** $H_2\underline{O}_2 \longrightarrow H_2O$ ***i)*** $H_2\underline{O}_2 \longrightarrow O_2$

How to balance redox half equations

Step **1** Work out the formula of the species before and after the change

2 Work out the oxidation state of the element before and after the change

3 Add electrons to one side of the equation so that the oxidation states balance

4 If the charges on all the species (ions and electrons) on either side of the equation do not balance then add sufficient H^+ ions to one of the sides to balance the charges

5 If the equation still doesn't balance, add sufficient water molecules to one side

Example 1 *Iron(II) being oxidised to iron(III).*

Step 1	$Fe^{2+} \longrightarrow Fe^{3+}$	
Step 2	+2 +3	
Step 3	$Fe^{2+} \longrightarrow Fe^{3+} + e^-$	*now balanced*

Example 2 *MnO_4^- being reduced to Mn^{2+} in acidic solution*

Step 1	$MnO_4^- \longrightarrow Mn^{2+}$	
Step 2	+7 +2	
Step 3	$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$	
Step 4	$MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+}$	
Step 5	$MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$	*now balanced*

Q.7 ***Balance the following half equations***

$I_2 \longrightarrow I^-$

$C_2O_4^{2-} \longrightarrow 2CO_2$

$H_2O_2 \longrightarrow O_2$

$H_2O_2 \longrightarrow H_2O$

$Cr_2O_7^{2-} \longrightarrow Cr^{3+}$

$SO_4^{2-} \longrightarrow SO_2$

Combining half equations

A combination of two ionic half equations, one involving oxidation and the other reduction, produces a balanced REDOX equation. The equations can be balanced as follows...

Step **1** Write out the two half equations
2 Multiply the equations so that the number of electrons in each is the same
3 Add the two equations and cancel out the electrons on either side of the equation
4 If necessary, cancel out any other species which appear on both sides of the equation

Example *The reaction between manganate(VII) and iron(II).*

Step 1

$$Fe^{2+} \longrightarrow Fe^{3+} + e^{-} \qquad \text{Oxidation}$$
$$MnO_4^{-} + 5e^{-} + 8H^{+} \longrightarrow Mn^{2+} + 4H_2O \qquad \text{Reduction}$$

Step 2

$$5Fe^{2+} \longrightarrow 5Fe^{3+} + 5e^{-} \qquad \textbf{multiplied by 5}$$
$$MnO_4^{-} + 5e^{-} + 8H^{+} \longrightarrow Mn^{2+} + 4H_2O \qquad \textbf{multiplied by 1}$$

Step 3

$$MnO_4^{-} + 5e^{-} + 8H^{+} + 5Fe^{2+} \longrightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^{-}$$

Step 4

$$MnO_4^{-} + \cancel{5e^{-}} + 8H^{+} + 5Fe^{2+} \longrightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + \cancel{5e^{-}}$$

gives

$$\mathbf{MnO_4^{-} + 8H^{+} + 5Fe^{2+} \longrightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}}$$

Q.8 ***Construct balanced redox equations for the reactions between***

a) **Mg and H^{+}** ***b)*** **$Cr_2O_7^{2-}$ and Fe^{2+}** ***c)*** **H_2O_2 and MnO_4^{-}**
d) **$C_2O_4^{2-}$ and MnO_4^{-}** ***e)*** **$S_2O_3^{2-}$ and I_2** ***f)*** **$Cr_2O_7^{2-}$ and I^{-}**