OXIDATION NUMBERS

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 and 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 **'The sum of the oxidation numbers adds up to zero'**

Elements H in H_2 = 0

Compounds C in CO_2 = +4 and O = -2 +4 and 2(-2) = 0

- *CO_2 is neutral, so the sum of the oxidation numbers must be zero*
- *one element must have a positive ON, 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*

Q.2 *If the oxidation number of O is -2, state the oxidation number of the other element in...*

a) SO_2 *b) SO_3* *(c) NO* *d) NO_2*

e) N_2O *f) MnO_2* *g) P_4O_{10}* *h) Cl_2O_7*

Complex ions **'The sum of the oxidation numbers adds up to the charge on the ion'**

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

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

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

WHICH OXIDATION NUMBER ?

- 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 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 number is often included in the name

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

Na P Ba Pb S Mn Cr

State the most common and the maximum oxidation number in compounds of...

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

Q.4 *Give the oxidation number 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-}	$S_2O_3^{2-}$	$S_4O_6^{2-}$

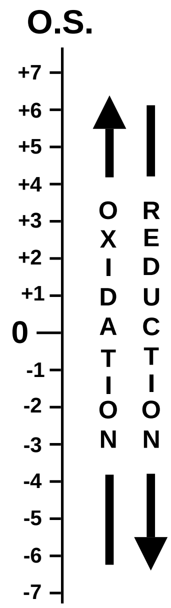
What is odd about the value of the oxidation state of S in $S_4O_6^{2-}$ *?*

Can it have such a value ? Can you provide a suitable explanation ?

Q.5 *What is the oxidation number of each element in the following compounds ?*

CH_4	***C*** =	PCl_3	***P*** =	NCl_3	***N*** =
	H =		***Cl*** =		***Cl*** =
CS_2	***C*** =	ICl_5	***I*** =	BrF_3	***Br*** =
	S =		***Cl*** =		***F*** =
$MgCl_2$	***Mg*** =	H_3PO_4	***H*** =	NH_4Cl	***N*** =
	Cl =		***P*** =		***H*** =
			O =		***Cl*** =
H_2SO_4	***H*** =	$MgCO_3$	***Mg*** =	$SOCl_2$	***S*** =
	S =		***C*** =		***O*** =
	O =		***O*** =		***Cl*** =

REDOX REACTIONS



Redox When reduction and oxidation take place

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

Reduction Gain of electrons; species becomes more negative / less positive

REDUCTION in O.N. Species has been REDUCED

e.g. Cl is reduced to Cl^- (0 to -1)

INCREASE in O.N. 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.6 *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.7 *What change takes place in the oxidation state of the underlined element ? Classify the change as oxidation (O), reduction (R) or neither (N).*

a) $\underline{\mathbf{N}}O_3^- \longrightarrow NO$ b) $H\underline{\mathbf{N}}O_3 \longrightarrow N_2O$

c) $\underline{\mathbf{C}}H_4 \longrightarrow CO$ d) $\underline{\mathbf{Cr}}_2O_7^{2-} \longrightarrow Cr^{3+}$

e) $\underline{\mathbf{S}}O_3^{2-} \longrightarrow SO_4^{2-}$ f) $\underline{\mathbf{Cr}}_2O_7^{2-} \longrightarrow CrO_4^{2-}$

g) $H_2\underline{\mathbf{O}}_2 \longrightarrow H_2O$ h) $H_2\underline{\mathbf{O}}_2 \longrightarrow O_2$

How to balance redox half equations

Step

1 Work out the formula of the species before and after the change;

2 If different numbers of the relevant species are on both sides, balance them

3 Work out the oxidation number of the element before and after the change

4 Add electrons to one side of the equation so the oxidation numbers balance

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

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

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

Steps1/2	$Fe^{2+} \longrightarrow Fe^{3+}$	
Step 3	+2 +3	
Step 4	$Fe^{2+} \longrightarrow Fe^{3+} + e^-$	*now balanced*

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

Steps 1/2	$MnO_4^- \longrightarrow Mn^{2+}$	
Step 3	+7 +2	
Step 4	$MnO_4^- + 5e^- \longrightarrow Mn^{2+}$	
Step 5	$MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+}$	
Step 6	$MnO_4^- + 5e^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O$	*now balanced*

Q.8 *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 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

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

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

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

Step 2 $5Fe^{2+} \longrightarrow 5Fe^{3+} + 5e^-$ ***multiplied by 5***

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

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

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

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

Q.9 *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^-