OXIDATION NUMBERS

Solns

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 0 *b) Fe^{3+}* +3 *c) S^{2-}* −2

d) Cu 0 *e) Cu^{2+}* +2 *f) Cu^+* +1

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 +4 *b) SO_3* +6 *(c) NO* +2 *d) NO_2* +4

e) N_2O +1 *f) MnO_2* +4 *g) P_4O_{10}* +5 *h) Cl_2O_7* +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 +1 *P* +5 *Ba* +2 *Pb* +4 *S* +6 *Mn* +7 *Cr* +6

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

	Li	*Br*	*Sr*	*O*	*B*	*N*
COMMON	+1	−1	+2	−2	+3	−3
MAXIMUM	+1	+7	+2	+2	+3	+5.

Q.4 *Give the oxidation number of the element other than O, H or F in*

SO_2	+4	NH_3	−3	NO_2	+4	NH_4^+	−3
IF_7	+7	Cl_2O_7	+7	MnO_4^{2-}	+6	NO_3^-	+5
NO_2^-	+3	SO_3^{2-}	+4	$S_2O_3^{2-}$	+2	$S_4O_6^{2-}$	+2.5

What is odd about the value of the oxidation state of S in $S_4O_6^{2-}$? not whole number.

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

Yes one Sulphur is +3 and another Sulphur +2

depending where the S=O is

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

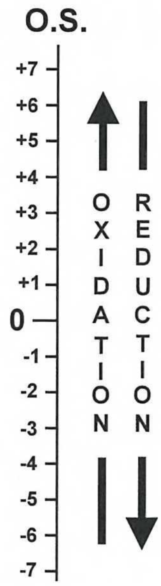
CH_4	**C** = −4	PCl_3	**P** = +3	NCl_3	**N** = +3
	H = +1		**Cl** = −1		**Cl** = −1
CS_2	**C** = +4	ICl_5	**I** = +5	BrF_3	**Br** = +3
	S = −2		**Cl** = −1		**F** = −1
$MgCl_2$	**Mg** = +2	H_3PO_4	**H** = +1	NH_4Cl	**N** = −3
	Cl = −1		**P** = +5		**H** = +1
			O = −2		**Cl** = −1
H_2SO_4	**H** = +1	$MgCO_3$	**Mg** = +2	$SOCl_2$	**S** = +4
	S = +6		**C** = +4		**O** = −2
	O = −2		**O** = −2		**Cl** = −1

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 **R**eduction **I**s the **G**ain of electrons

Q.6 *Classify the following (unbalanced) changes as oxidation, reduction or neither.*

a) $Mg \longrightarrow Mg^{2+}$ oxi

b) $O^{2-} \longrightarrow O$ oxi

c) $Al^{3+} \longrightarrow Al$ red

d) $Fe^{3+} \longrightarrow Fe^{2+}$ red

e) $Ti^{3+} \longrightarrow Ti^{4+}$ oxi

f) $2Q \longrightarrow Q_2$ non-redox

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) $\underline{N}O_3^- \longrightarrow NO$ (+5 → +2) R

b) $H\underline{N}O_3 \longrightarrow N_2O$ (+5 → +1) R

c) $\underline{C}H_4 \longrightarrow CO$ (-4 → +2) O

d) $\underline{Cr}_2O_7^{2-} \longrightarrow Cr^{3+}$ (+6 → +3) R

e) $\underline{S}O_3^{2-} \longrightarrow SO_4^{2-}$ (+4 → +6) O

f) $\underline{Cr}_2O_7^{2-} \longrightarrow CrO_4^{2-}$ (+6 → +6) N

g) $H_2\underline{O}_2 \longrightarrow H_2O$ (-1 → -2) R

h) $H_2\underline{O}_2 \longrightarrow O_2$ (-1 → 0) O.

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*

$I_2 + 2e^- \longrightarrow 2I^-$

$2e^- + 4H^+ + C_2O_4^{2-} \longrightarrow 2CO_2 + 2H_2O$

$H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$

$2e^- + 2H^+ + H_2O_2 \longrightarrow 2H_2O$

$6e^- + 14H^+ + Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$

$2e^- + 4H^+ + SO_4^{2-} \longrightarrow SO_2 + 2H_2O$

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^-

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a) Oxi $Mg \rightarrow Mg^{2+} + \cancel{2e^-}$

Red $\cancel{2e^-} + 2H^+ \rightarrow H_2$

$$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$$

b) Oxi $6\,(Fe^{2+} \rightarrow Fe^{3+} + \cancel{e^-}$

Red $Cr_2O_7^{2-} + 14H^+ + \cancel{6e^-} \rightarrow 2Cr^{3+} + 7H_2O$

$$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$$

c) Red $2\,(MnO_4^- + 8H^+ + \cancel{5e^-} \rightarrow Mn^{2+} + 4H_2O$

Oxi $5\,(H_2O_2 \rightarrow O_2 + 2H^+ + \cancel{2e^-}$

$$2MnO_4^- + \underset{6H^+}{\cancel{16H^+}} + 5H_2O_2 \rightarrow 2Mn^{2+} + 4H_2O + 5O_2 + \cancel{10H^+}$$

d) Oxi $5\,(C_2O_4^{2-} \rightarrow 2CO_2 + \cancel{2e^-}$

Red $2\,(MnO_4^- + 8H^+ + \cancel{5e^-} \rightarrow Mn^{2+} + 4H_2O$

$$5C_2O_4^{2-} + 2MnO_4^- + 16H^+ \rightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$$

e)

Red $\cancel{2e^-} + I_2 \rightarrow 2I^-$

$$\overset{+2}{} \qquad \overset{+2.5}{}$$

Oxi $H_2O + S_2O_3^{2-} \rightarrow S_2O_4^{2-} + 2H^+ + \cancel{2e^-}$

$$H_2O + S_2O_3^{2-} + I_2 \rightarrow 2I^- + S_2O_4^{2-} + 2H^+$$

f)

Red $Cr_2O_7^{2-} + 14H^+ + \cancel{6e^-} \rightarrow 2Cr^{3+} + 7H_2O$

Oxi $3(2I^- \rightarrow I_2 + \cancel{2e^-}$

$$Cr_2O_7^{2-} + 14H^+ + 6I^- \rightarrow 3I_2 + 2Cr^{3+} + 7H_2O$$