# 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 +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 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

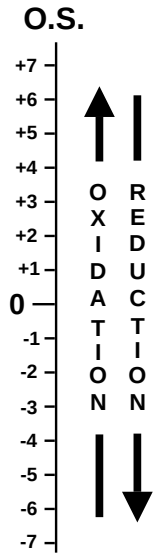
**Redox** When reduction and oxidation take place

**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
**R**eduction **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 (O), reduction (R) or neither (N).*

a) $\underline{N}O_3^-$ ——> $NO$

b) $H\underline{N}O_3$ ——> $N_2O$

c) $\underline{C}H_4$ ——> $CO$

d) $\underline{Cr}_2O_7^{2-}$ ——> $Cr^{3+}$

e) $\underline{S}O_3^{2-}$ ——> $SO_4^{2-}$

f) $\underline{Cr}_2O_7^{2-}$ ——> $CrO_4^{2-}$

g) $H_2\underline{O}_2$ ——> $H_2O$

h) $H_2\underline{O}_2$ ——> $O_2$

**How to balance redox half equations**

**Step 1** Work out the formula of the species before and after the change; balance if required

**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* $Fe^{2+}$ ———> $Fe^{3+}$

*Step 2* +2 +3

*Step 3* $Fe^{2+}$ ———> $Fe^{3+} + e^-$ *now balanced*

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

*Step 1* $MnO_4^-$ ———> $Mn^{2+}$

*Step 2* +7 +2

*Step 3* $MnO_4^- + 5e^-$ ———> $Mn^{2+}$

*Step 4* $MnO_4^- + 5e^- + 8H^+$ ———> $Mn^{2+}$

*Step 5* $MnO_4^- + 5e^- + 8H^+$ ———> $Mn^{2+} + 4H_2O$ *now balanced*

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

| | | |
|---|---|---|
| $I_2$ | —> | $I^-$ |
| $C_2O_4^{2-}$ | —> | $2CO_2$ |
| $H_2O_2$ | —> | $O_2$ |
| $H_2O_2$ | —> | $H_2O$ |
| $Cr_2O_7^{2-}$ | —> | $Cr^{3+}$ |
| $SO_4^{2-}$ | —> | $SO_2$ |

## 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*

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

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

*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 $I^-$*