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**

<table>
<thead>
<tr>
<th>Type</th>
<th>Element</th>
<th>Oxidation Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>atoms</td>
<td>Na</td>
<td>0</td>
</tr>
<tr>
<td>cations</td>
<td>Na⁺</td>
<td>+1</td>
</tr>
<tr>
<td>anions</td>
<td>Cl⁻</td>
<td>−1</td>
</tr>
</tbody>
</table>

*The number of electrons which must be added or removed to become neutral*

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

- a) N
- b) Fe³⁺
- c) S²⁻
- d) Cu
- e) Cu²⁺
- f) Cu⁺

**Molecules**

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

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<th>Type</th>
<th>Element</th>
<th>Oxidation Numbers</th>
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<tbody>
<tr>
<td>Elements</td>
<td>H in H₂</td>
<td>0</td>
</tr>
<tr>
<td>Compounds</td>
<td>C in CO₂</td>
<td>+4 and O = -2 +4 and 2(-2) = 0</td>
</tr>
</tbody>
</table>

- CO₂ 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) SO₂
- b) SO₃
- c) NO
- d) NO₂
- e) N₂O
- f) MnO₂
- g) P₄O₁₀
- h) Cl₂O₇
Complex ions ‘The sum of the oxidation numbers adds up to the charge on the ion’

in $\text{SO}_4^{2-}$  $S = +6$, $O = -2$  [i.e.  $+6 + 4(-2) = -2$ ] the ion has a 2- charge

Example What is the oxidation number (O.N.) of Mn in $\text{MnO}_4^-$?

• 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 $\text{MnO}_4^-$ = +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$_2$) |
| HYDROGEN (-1) | except -1 hydride ion, H$^-$ [in sodium hydride, NaH] |

| OXYGEN (-2) | except 0 atom (O) and molecule (O$_2$) |
| OXYGEN (+2) | except +2 in hydrogen peroxide, H$_2$O$_2$ |

| FLUORINE (-1) | except 0 atom (F) and molecule (F$_2$) |

| Metals |
|• have positive values in compounds |
|• value is usually that of the Group Number  $\text{Al}$ is +3 |
|• values can go no higher than the Group No.  $\text{Mn}$ can be  +2,+4,+6,+7 |

| Non metals |
|• mostly negative based on their usual ion  $\text{Cl}$ is usually -1 |
|• can have values up to their Group No.  $\text{Cl}$ 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 $\text{MnO}_2$ |
| sulphur(VI) oxide for $\text{SO}_3$ |
| dichromate(VI) for $\text{Cr}_2\text{O}_7^{2-}$ |
| phosphorus(V) chloride for $\text{PCl}_5$. |
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

- $\text{SO}_2$
- $\text{NH}_3$
- $\text{NO}_2$
- $\text{NH}_4^+$
- $\text{IF}_7$
- $\text{Cl}_2\text{O}_7$
- $\text{MnO}_4^{2-}$
- $\text{NO}_5^-$
- $\text{NO}_2^-$
- $\text{SO}_4^{2-}$
- $\text{S}_2\text{O}_3^{2-}$
- $\text{S}_4\text{O}_6^{2-}$

What is odd about the value of the oxidation state of S in $\text{S}_4\text{O}_6^{2-}$? 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?

- $\text{CH}_4$
- $\text{PCl}_3$
- $\text{P}$
- $\text{NCl}_3$
- $\text{N}$
- $\text{H}$
- $\text{Cl}$
- $\text{H}_2\text{SO}_4$
- $\text{MgCO}_3$
- $\text{Mg}$
- $\text{SOCl}_2$
- $\text{S}$
- $\text{Cl}$
- $\text{O}$
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)

<table>
<thead>
<tr>
<th>OIL RIG</th>
<th>Oxidation Is the Loss</th>
</tr>
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<tr>
<td></td>
<td>Reduction Is the Gain of electrons</td>
</tr>
</tbody>
</table>

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

| a) Mg   ——> Mg\(^{2+}\) | b) O\(^2-\) ——> O |
| c) Al\(^{3+}\) ——> Al | d) Fe\(^{3+}\) ——> Fe\(^{2+}\) |
| e) Ti\(^{3+}\) ——> Ti\(^{4+}\) | f) 2Q ——> 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) NO\(_3^-\) ——> NO | b) HNO\(_3\) ——> N\(_2\)O |
| c) CH\(_4\) ——> CO | d) Cr\(_2\)O\(_7^2-\) ——> Cr\(^{3+}\) |
| e) SO\(_3^{2-}\) ——> SO\(_4^{2-}\) | f) Cr\(_2\)O\(_7^2-\) ——> CrO\(_4^{2-}\) |
| g) H\(_2\)O\(_2\) ——> H\(_2\)O | h) H\(_2\)O\(_2\) ——> O\(_2\) |
How to balance redox half equations

Step 1 Work out the formula of the species before and after the change;
Step 2 If different numbers of the relevant species are on both sides, balance them
Step 3 Work out the oxidation number of the element before and after the change
Step 4 Add electrons to one side of the equation so the oxidation numbers balance
Step 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
Step 6 If the equation still doesn’t balance, add sufficient water molecules to one side

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

Steps 1/2 $Fe^{2+} \rightarrow Fe^{3+}$
Step 3 +2 +3
Step 4 $Fe^{2+} \rightarrow Fe^{3+} + e^-$ now balanced

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

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

Q.8 Balance the following half equations

$I_2 \rightarrow I^-$
$C_2O_4^{2-} \rightarrow 2CO_2$
$H_2O_2 \rightarrow O_2$
$H_2O_2 \rightarrow H_2O$
$Cr_2O_7^{2-} \rightarrow Cr^{3+}$
$SO_4^{2-} \rightarrow SO_2$
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**
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**

\[
\begin{align*}
\text{Oxidation: } & \quad Fe^{2+} \rightarrow Fe^{3+} + e^- \\
\text{Reduction: } & \quad MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O
\end{align*}
\]

**Step 2**

\[
\begin{align*}
\text{multiplied by 5: } & \quad 5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^- \\
\text{multiplied by 1: } & \quad MnO_4^- + 5e^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O
\end{align*}
\]

**Step 3**

\[
\begin{align*}
MnO_4^- + 5e^- + 8H^+ + 5Fe^{2+} & \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^- \\
MnO_4^- + 5e^- + 8H^+ + 5Fe^{2+} & \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^-
\end{align*}
\]

gives

\[
MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}
\]

**Q.9**

Construct balanced redox equations for the reactions between

a) Mg and H⁺
b) Cr₂O₇²⁻ and Fe²⁺
c) H₂O₂ and MnO₄⁻
d) C₂O₄²⁻ and MnO₄⁻
e) S₂O₃²⁻ and I₂
f) Cr₂O₇²⁻ and I⁻