Oxidation and Reduction HL

Name:

<table>
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<th>Atomic Structure</th>
<th>Objectives</th>
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| 5. Oxidation and Reduction | - define oxidation and reduction in terms of electron transfer  
- use simple examples, e.g. Na with Cl₂, Mg with O₂, Zn with Cu²⁺ to describe oxidation and reduction in terms of electron transfer  
- apply knowledge of oxidation and reduction to explain the rusting of iron  
- define oxidising agent and reducing agent  
- carry out an experiment to show that halogens act as oxidising agents (reactions with bromides, iodides, Fe²⁺ and sulfites; half equations only required)  
- carry out an experiment to demonstrate the displacement reactions of metals (Zn with Cu²⁺, Mg with Cu²⁺) |

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<th>Chemical Bonding</th>
<th>6. Oxidation Numbers</th>
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|                   | - define oxidation number, oxidation state  
- define oxidation and reduction in terms of change of oxidation numbers  
- state the rules for oxidation numbers (exclude peroxides, except for hydrogen peroxide)  
- calculate oxidation numbers of transition metals in their compounds and of other elements  
- use oxidation numbers in nomenclature of transition metal compounds  
- give an example of an oxidising and a reducing bleach |

Oxidation and reduction can be described in four ways:

In terms of:

1. Addition/removal of oxygen.
3. Electron transfer.
4. Change in oxidation number.

1. **Addition/Removal of Oxygen:**

   *Def*: Oxidation is the addition of oxygen.

   E.g. \( \text{C} + \text{O}_2 \rightarrow \text{CO}_2 \) The carbon gains oxygen, therefore the carbon is oxidised.

   *Def*: Reduction is the removal of oxygen.

   E.g. \( \text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O} \) The copper loses oxygen, therefore the copper is reduced.

2. **Addition/Removal of Hydrogen:**

   *Def*: Oxidation is the removal of hydrogen.

   E.g. \( \text{H}_2\text{S} + \text{Cl}_2 \rightarrow \text{S} + 2\text{HCl} \) The sulphur loses hydrogen, therefore the S is oxidised.

   *Def*: Reduction is the addition of hydrogen.

   E.g. \( \text{CO} + 2\text{H}_2 \rightarrow \text{CH}_3\text{OH} \) The carbon monoxide gains hydrogen, therefore the CO is reduced.
3. **Electron Transfer:**

*Def*: Oxidation is the loss of electrons.

E.g. \( \text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu} \) The zinc loses 2e\(^-\), therefore the zinc is oxidised.

*Def*: Reduction is the gain of electrons.

E.g. \( \text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu} \) The copper gains 2e\(^-\), therefore the copper is reduced.

Remember, for electron transfer: **O**xidation I*s* Loss R*eduction I*s* G*ain*.

4. **Change in Oxidation Number:** (Finding oxidation numbers mentioned later in these notes)

*Def*: Oxidation is an increase in oxidation number.

E.g. \( \text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O} \) The O.N. of H increased from 0 to +1, therefore H is oxidised.

*Def*: Reduction is a decrease in oxidation number.

E.g. \( \text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O} \) The O.N. of O decreased from 0 to -2, therefore O is reduced.

**Oxidising/Reducing Agents:**

*Def*: An **Oxidising Agent** is a substance that brings about oxidation in other substances by being reduced.

*Def*: A **Reducing Agent** is a substance that brings about reduction in other substances by being oxidised.

E.g. \( \text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu} \)

**Assigning Oxidation Numbers:**

*Def*: **Oxidation Number** is the apparent charge an atom has when electrons are distributed according to certain rules.

**Rules:**

1. The oxidation number of any element that is not bonded to another different element is 0.
   E.g. O.N. of H in \( \text{H}_2 \) is 0. O.N. of Na is 0. O.N. of Cl in \( \text{Cl}_2 \) is 0.

2. The oxidation number of an ion of an element is the same as its charge.
   E.g. O.N. of H in \( \text{H}^+ \) is +1. O.N. of Ca in \( \text{Ca}^{2+} \) is +2. O.N. of Cl in \( \text{Na}^+\text{Cl}^- \) is -1.

3. The total of the oxidation numbers in a neutral compound must add to give 0.
   E.g. For \( \text{H}_2\text{O} \), each H is +1 and O is -2. This gives \(2(+1)+(-2)=0\).
4. Oxygen has an oxidation number of -2.
   Exceptions:
   (a) Peroxides like $\text{H}_2\text{O}_2$, as O.N of H = +1, each O must have an O.N. of -1 to give a total sum of 0.
   (b) In $\text{OF}_2$, as O.N. of F = -1, O must have an O.N. of +2.

5. Hydrogen has an oxidation number of +1.
   Exception:
   (a) In metal hydrides like NaH, CaH$_2$ (H after a metal). Here H has an O.N. of -1.

6. Halogens (Group VII elements) have an oxidation number of -1.
   Exception:
   (a) When bonded to a more electronegative element (O or F), halogens can have other oxidation numbers
   which are calculated using rule 3.

7. The total of the oxidation numbers in a complex ion must add to give the charge on the ion.
   E.g. For $\text{NO}_3^-$, each O has an O.N. of -2, giving 3(-2)=-6. The total must give -1 as this is the charge on the
   ion. Therefore N must have an O.N. of +5.

**Balancing Redox Equations:**

Using oxidation numbers, balance the following equation:

$$\text{MnO}_4^- + \text{Fe}^{2+} + \text{H}^+ \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+} + \text{H}_2\text{O}$$

1. Assign oxidation numbers:

   $\text{MnO}_4^- + \text{Fe}^{2+} + \text{H}^+ \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+} + \text{H}_2\text{O}$

2. Show the number of electrons gained/lost:

   $\text{MnO}_4^- + \text{Fe}^{2+} + \text{H}^+ \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+} + \text{H}_2\text{O}$

3. Write the half-equations for the oxidation/reduction of the two substances above:

   $$\text{MnO}_4^- + 5\text{e}^- \rightarrow \text{Mn}^{2+}$$
   $$\text{Fe}^{2+} - \text{1e}^- \rightarrow \text{Fe}^{3+}$$

Balance so that:
   (a) No. of Mn atoms is equal on both sides
   (b) No. of Fe atoms is equal on both sides
   (c) electrons gained = electrons lost:

   $$\text{MnO}_4^- + 5\text{e}^- \rightarrow \text{Mn}^{2+}$$
   $$5\text{Fe}^{2+} - 5\text{e}^- \rightarrow 5\text{Fe}^{3+}$$

4. Write the coefficients from step 3 into the original equation:

   $$\text{MnO}_4^- + 5\text{Fe}^{2+} + \text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + \text{H}_2\text{O}$$

5. Balance the rest of the equation by inspection (leave H until last):

   $$\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$$