

Oxidation Numbers and Writing Redox Equations

Rules for assigning oxidation numbers

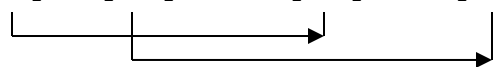
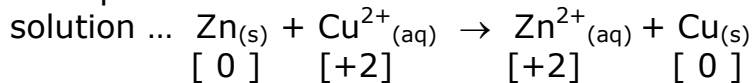
1. oxidation numbers apply to an element in a particular substance
2. for **elements, ON = 0**
example: in $\text{Cl}_{2(g)}$, which is an element ... ON for Cl = 0
3. for **H atoms, ON = +1** [unless in metal hydride eg. NaH ...where H = -1]
Example: in H_2SO_4 , a compound, ON for H = +1
4. for **O atoms, ON = -2** [unless in peroxide eg. Na_2O_2 ... where ON = -1]
Example: in $\text{K}_2\text{Cr}_2\text{O}_7$, a compound, ON for O = -2
5. for **simple ions or ionic compounds, ON = charge on ion**
Example: in the Ca^{2+} ion in CaCl_2 , ON for Ca = +2
6. for **molecular ions, $\Sigma\text{ON} = \text{charge on ion}$**
Example: in molecular ion, MnO_4^- ... $\Sigma\text{ON} = \text{charge on ion}$
 $\text{ON for Mn} + 4[\text{ON for O}] = -1$
 $\text{ON for Mn} + [-8] = -1$
 $\text{ON for Mn} = -1 + 8 = +7$
7. for **compounds, $\Sigma\text{ON} = 0$**
Example: in $\text{K}_2\text{Cr}_2\text{O}_7$, ... $\Sigma\text{ON} = 0$
 $2[\text{ON of K}^+] + 2[\text{ON of Cr}] + 7[\text{ON of O}] = 0$
 $2[+1] + 2[\text{ON of Cr}] + 7[-2] = 0$
 $+2 + 2[\text{ON of Cr}] -14 = 0$
 $2[\text{ON of Cr}] = 0 - 2 + 14 = +12$
 $\text{ON of Cr} = +12/2 = +6$

Using Oxidation Numbers

Oxidation numbers can be used to determine if a chemical reaction is redox and if a substance has been oxidised [ie. is a reductant] or reduced [ie. is an oxidant]

- if there has been any change in oxidation number, the reaction is a redox reaction

Example: Consider the chemical reaction between zinc metal and copper sulfate



the change in ON shows that the reaction must be redox

- if the oxidation number of an atom has increased, the substance that it is part of has been oxidised and is therefore the reductant
Example: in the reaction above, the ON of Zn has increased from 0 [in $\text{Zn}_{(s)}$] to +2 [in $\text{Zn}^{2+}_{(aq)}$] ... Zn has been oxidised, therefore Zn is the reductant

- if the oxidation number of an atom has been reduced, the substance that it is part of has been reduced and is therefore the oxidant
- Example: in the reaction above, the ON of Cu has reduced from +2 [in $\text{Cu}^{2+}_{(\text{aq})}$] to 0 [in $\text{Cu}_{(\text{s})}$] ... Cu^{2+} has been reduced, therefore Cu^{2+} is the oxidant

Writing Redox Equations

A redox reaction equation can be formed by adding together 2 redox half reactions. In the reaction above ... $\text{Zn}_{(\text{s})} + \text{Cu}^{2+}_{(\text{aq})} \rightarrow \text{Zn}^{2+}_{(\text{aq})} + \text{Cu}_{(\text{s})}$

- the oxidation half equation is $\text{Zn}_{(\text{s})} \rightarrow \text{Zn}^{2+}_{(\text{aq})} + 2\text{e}^{-}$
- the reduction half equation is $\text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^{-} \rightarrow \text{Cu}_{(\text{s})}$

Notice that **electrons** are ... **produced in the oxidation, used up in the reduction**, but don't appear in the overall redox equation. The electrons, on different sides of the half equations, cancel out when the half equations are added.

- When half equations are added, the number of electrons in them must be equalised.

Rules for Establishing Redox Half Reaction Equations [polyatomic ions]

For atoms or simple ions, redox half reaction equations are easily established ...

Example: oxidation of zinc atoms = $\text{Zn}_{(\text{s})} \rightarrow \text{Zn}^{2+}_{(\text{aq})} + 2\text{e}^{-}$

For polyatomic ions, a list of rules can be used. Consider the half reaction in which dichromate ions $[\text{Cr}_2\text{O}_7^{2-}_{(\text{aq})}]$ are reduced to chromium III ions $[\text{Cr}^{3+}_{(\text{aq})}]$...

1. balance all elements [except hydrogen and oxygen]

Example: $\text{Cr}_2\text{O}_7^{2-}_{(\text{aq})} \rightarrow 2\text{Cr}^{3+}_{(\text{aq})}$

2. balance the O atoms by adding water

Example: $\text{Cr}_2\text{O}_7^{2-}_{(\text{aq})} \rightarrow 2\text{Cr}^{3+}_{(\text{aq})} + 7\text{H}_2\text{O}_{(\text{l})}$

3. balance H atoms by adding H^{+}

Example: $\text{Cr}_2\text{O}_7^{2-}_{(\text{aq})} + 14\text{H}^{+}_{(\text{aq})} \rightarrow 2\text{Cr}^{3+}_{(\text{aq})} + 7\text{H}_2\text{O}_{(\text{l})}$

4. balance charge by adding electrons

Example: $\text{Cr}_2\text{O}_7^{2-}_{(\text{aq})} + 14\text{H}^{+}_{(\text{aq})} + 6\text{e}^{-} \rightarrow 2\text{Cr}^{3+}_{(\text{aq})} + 7\text{H}_2\text{O}_{(\text{l})}$

Using the Electrochemical Series [see p.361 textbook]

The electrochemical series is a list of redox half reaction equations, all written as reductions, organised in order of decreasing oxidising strength. When written in reverse, a reduction half reaction equation becomes an oxidation half reaction equation. So, by reference to the electrochemical series, the redox half reaction equations can simply be written from the list.

Oxidation Numbers and Redox Reaction Equations

1. When zinc metal is dropped into dilute hydrochloric acid, a reaction occurs that liberates hydrogen gas and forms a zinc chloride solution.

(a) write a balanced chemical equation for the reaction

(b) write a balanced ionic equation for the reaction

(c) establish the oxidation numbers of all reactants and products in the ionic equation for the reaction

(d) use these oxidation numbers to:-

(i) show that this reaction is redox

(ii) identify the oxidant and reductant in the reaction

(e) write redox half reaction equations for the reaction

2. When a solution of iron(II) sulfate [$\text{FeSO}_{4(\text{aq})}$] is reacted with an acidified solution of potassium permanganate [$\text{KMnO}_{4(\text{aq})}$], a solution containing $\text{Mn}^{2+}_{(\text{aq})}$ ions and $\text{Fe}^{3+}_{(\text{aq})}$ ions is formed.

(a) write the redox half reaction equation for the formation of $\text{Fe}^{3+}_{(\text{aq})}$ ions

(b) derive and write the redox half reaction equation for the conversion of $\text{MnO}_4^{-}_{(\text{aq})}$ ions to $\text{Mn}^{2+}_{(\text{aq})}$ ions

(c) combine the two redox half reaction equations to write the balanced ionic equation for the reaction between $\text{FeSO}_{4(\text{aq})}$ and $\text{KMnO}_{4(\text{aq})}$.

(d) identify the oxidant and reductant in the redox reaction between $\text{FeSO}_{4(\text{aq})}$ and $\text{KMnO}_{4(\text{aq})}$

(e) explain why the $\text{KMnO}_{4(\text{aq})}$ solution has to be acidified before any appreciable reaction occurs

(f) Use the electrochemical series [p.510 text] to verify the redox half reaction equations you wrote in answering 2(a) and 2(b).
