Redox Reactions

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Reduction

Transformation of matter from one kind into another occurs through the various types of reactions. One important type of reaction is redox reaction

Oxidation

What is oxidation?

Oxidation: **When a molecule/ion loses electrons (becomes more positive)**

What is reduction?

Reduction: **When a molecule/ion gains electrons (becomes more negative)**

Thus we can differentiate oxidation and reduction as

Oxidation:

- Gain of oxygen
- Loss of electrons

Increase in oxidation number

Reduction:

- Loss of oxygen
- Gain of electrons

Decrease in oxidation number

For example
If we do 2 Experiments

1. **Burning of magnesium**
2. **Copper in silver nitrate solution**

Oxidised – gains oxygen

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Must be a redox!

Oxidised – loss of e^-

$Mg \rightarrow Mg^{2+}$ $+2e^-$

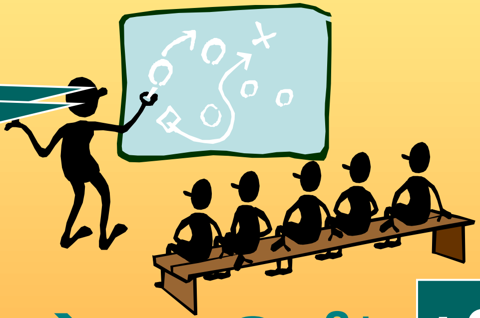
Put the e^- in.

Reduced – gain of e^-

O $+2e^-$ $\rightarrow O^{2-}$

$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

Oxidised?
Reduced?



Oxidised – loss of e^- $Cu \rightarrow Cu^{2+}$ $+2e^-$

Reduced – gain of e^- Ag^+ $+e^-$ $\rightarrow Ag$

Oxidation and Reduction (Redox)

Early chemists saw "oxidation" reactions only as the combination of a material with oxygen to produce an oxide.

- **For example, when methane burns in air, it oxidizes and forms oxides of carbon and hydrogen.**

Oxidation and Reduction (Redox)

- But, not all oxidation processes that use oxygen involve burning:
  - Elemental iron slowly oxidizes to compounds such as iron (III) oxide, commonly called **"rust"**
  - Bleaching stains in fabrics
  - Hydrogen peroxide also releases oxygen when it decomposes

Oxidation and Reduction (Redox)

A process called **"reduction"** is the opposite of oxidation, and originally meant the loss of oxygen from a compound

- Oxidation and reduction **always occur simultaneously**
- The substance gaining oxygen (or losing electrons) is oxidized, while the substance losing oxygen (or gaining electrons) is reduced.

Oxidation and Reduction (Redox)

Today, many of these reactions may not even involve oxygen
Redox currently says that electrons are transferred between reactants

- $Mg \quad + \quad S \rightarrow \quad Mg^{2+} \quad + \quad S^{2-}$

•**The magnesium atom (which has zero charge) changes to a magnesium ion by losing 2 electrons, and is *oxidized* to Mg^{2+}**

•**The sulfur atom (which has no charge) is changed to a sulfide ion by gaining 2 electrons, and is *reduced* to S^{2-}**

Oxidation and Reduction (Redox)

$$2\overset{0}{Na} + \overset{0}{Cl_2} \rightarrow 2\overset{+1}{Na}\overset{-1}{Cl}$$

Each sodium atom loses one electron:

$$\overset{0}{Na} \rightarrow \overset{+1}{Na} + e^-$$

Each chlorine atom gains one electron:

$$\overset{0}{Cl} + e^- \rightarrow \overset{-1}{Cl}$$

Lose of Electrons = Oxidation

$$\overset{0}{Na} \rightarrow \overset{+1}{Na} + e^{-}$$

Sodium is oxidized

Gain of Electrons = Reduction

$$\overset{0}{Cl} + e^{-} \rightarrow \overset{-1}{Cl}$$

Chlorine is reduced

- Losing electrons is oxidation, and the substance that loses the electrons is called the **reducing agent.**
- Gaining electrons is reduction, and the substance that gains the electrons is called the **oxidizing agent.**

Mg is oxidized: loses e^-, becomes a Mg^{2+} ion

Mg is the reducing agent

$$Mg_{(s)} + S_{(s)} \rightarrow MgS_{(s)}$$

S is the oxidizing agent

S is reduced: gains e^- = S^{2-} ion

Oxidation and Reduction (Redox)

It is easy to see the loss and gain of electrons in *ionic* compounds, but what about *covalent* compounds?

- **In water, we learned that oxygen is highly *electronegative*, So:The oxygen gains electrons (is reduced and is the oxidizing agent), and the hydrogen loses electrons (is oxidized and is the reducing agent)**

Not All Reactions are Redox Reactions

- **Reactions in which there has been no change in oxidation number are not redox reactions.**

Examples:

$$\overset{+1}{Ag}\,\overset{+5}{N}\,\overset{-2}{O}_3(aq)+\overset{+1}{Na}\,\overset{-1}{Cl}(aq)\rightarrow\overset{+1}{Ag}\,\overset{-1}{Cl}(s)+\overset{+1}{Na}\,\overset{+5}{N}\,\overset{-2}{O}_3(aq)$$

$$2\,\overset{+1}{Na}\,\overset{-2}{O}\,\overset{+1}{H}(aq)+\overset{+1}{H}_2\,\overset{+6}{S}\,\overset{-2}{O}_4(aq)\rightarrow+\overset{+1}{Na}_2\,\overset{+6}{S}\,\overset{-2}{O}_4(aq)+\overset{+1}{H}_2\,\overset{-2}{O}(l)$$

Identifying Redox Equations

In general, all chemical reactions can be assigned to one of two classes:

1) **Oxidation-reduction, in which electrons are transferred:**
   - **Single-replacement, combination, decomposition, and combustion**
2) **This second class has no electron transfer, and includes all others:**
   - **Double-replacement and acid-base reactions**

Identifying Redox Equations

In an electrical storm, nitrogen and oxygen react to form nitrogen monoxide:

$$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)} \longrightarrow \textbf{YES!}$$

Is this a redox reaction?

If the oxidation number of an element in a reacting species changes, then that element has undergone either oxidation or reduction; therefore, the reaction as a whole must be a redox.

Oxidation Numbers

- **The oxidation number of an atom in an element is zero.**
  E.g. Mg in Mg, O in O_2.

Assigning Oxidation Numbers

- An "*oxidation number*" is a positive or negative number assigned to an atom to indicate its degree of oxidation or reduction.
- Generally, a bonded atom's oxidation number is the charge it would have if the electrons in the bond were assigned to the atom of the more electronegative element

Rules for Assigning Oxidation Numbers

1) The oxidation number of any uncombined element is zero.

2) The oxidation number of a monatomic ion equals its charge.

$$2\overset{0}{Na} + \overset{0}{Cl_2} \rightarrow 2\overset{+1}{Na}\overset{-1}{Cl}$$

Rules for Assigning Oxidation Numbers

3) The oxidation number of oxygen in compounds is -2, except in peroxides, such as H_2O_2 where it is -1.

4) The oxidation number of hydrogen in compounds is +1, except in metal hydrides, like NaH, where it is -1.

$$\overset{+1}{H}_2\overset{-2}{O}$$

Rules for Assigning Oxidation Numbers

5) The **sum** of the oxidation numbers of the atoms in the compound must equal 0.

$$\overset{+1}{H_2}\overset{-2}{O} \qquad \overset{+2}{Ca}(\overset{-2}{O}\overset{+1}{H})_2$$

2(+1) + (-2) = 0
H O

(+2) + 2(-2) + 2(+1) = 0
Ca O H

Rules for Assigning Oxidation Numbers

6) The sum of the oxidation numbers in the formula of a polyatomic ion is equal to its ionic charge.

$$\overset{?}{N}\overset{-2}{O}_3{}^{-}$$

$$\overset{?}{S}\overset{-2}{O}_4{}^{2-}$$

X + 3(-2) = -1
N O

thus X = +5

X + 4(-2) = -2
S O

thus X = +6

Reducing Agents and Oxidizing Agents

- **An increase in oxidation number = oxidation**
- **A decrease in oxidation number = reduction**

$$\overset{0}{Na} \uparrow \overset{+1}{Na} + e^-$$

Sodium is oxidized – it is the reducing agent

$$\overset{0}{Cl} + e^- \rightarrow \overset{-1}{Cl}$$

Chlorine is reduced – it is the oxidizing agent

Trends in Oxidation and Reduction

Active metals:

- Lose electrons easily
- Are easily oxidized
- Are strong reducing agents

Active nonmetals:

- Gain electrons easily
- Are easily reduced
- Are strong oxidizing agents

Oxidation Numbers and names

- To avoid any confusion when an element can have several oxidation numbers, the oxidation number is usually mentioned in the compound's name.
  In names like "elementate(X)", the number refers to "element" and not the associated oxygens.
- So if we look at some examples , we get the following names:-

$KMnO_4$	potassium manganate(VII)
$NaClO_3$	sodium chlorate(V)
$POCl_2F$	phosphorus(V) oxydichlorofluoride
NaH_2PO_3	sodium dihydrogenphosphate(III)
$K_2Cr_2O_7$	potassium dichromate(VI)

Balancing Redox Equations

It is essential to write a correctly balanced equation that represents what happens in a chemical reaction

Fortunately, two systematic methods are available, and are based on the fact that the total electrons gained in reduction equals the total lost in oxidation. The two methods:

1) Use oxidation-number changes
2) Use half-reactions

Using Oxidation-Number Changes

First, you compare the increase and decrease in oxidation numbers.

- **start with the skeleton equation**
- **Step 1: assign oxidation numbers to all atoms; write above their symbols**
- **Step 2: identify which are oxidized/reduced**
- **Step 3: use bracket lines to connect them**
- **Step 4: use coefficients to equalize**
- **Step 5: make sure they are balanced for *both* atoms and charge**

Using half-reactions

A half-reaction is an equation showing just the oxidation or just the reduction that takes place

- **they are then balanced separately, and finally combined**

  Step 1: write unbalanced equation in ionic form

  Step 2: write separate half-reaction equations for oxidation and reduction

  Step 3: balance the atoms in the half-reactions

(More steps on the next screen.)

Using half-reactions

continued

- **Step 4**: add enough electrons to one side of each half-reaction to balance the charges
- **Step 5**: multiply each half-reaction by a number to make the electrons equal in both
- **Step 6**: add the balanced half-reactions to show an overall equation
- **Step 7**: add the spectator ions and balance the equation

Choosing a Balancing Method

1) The oxidation number change method works well if the oxidized and reduced species appear only once on each side of the equation, and there are no acids or bases.
2) The half-reaction method works best for reactions taking place in acidic or alkaline solution.

Need to Balance Redox Equations

In redox equations, something will be oxidized and something will be reduced.

Sometimes the number of electrons that was lost in the oxidation process does not equal the number of electrons gained in the reduction process.

Therefore, we have to balance the redox equation

Balancing Redox Equations

The oxidation of Fe^{2+} to Fe^{3+} by $Cr_2O_7^{2-}$ in acid solution?

1. **Write the unbalanced equation for the reaction in ionic form.**

$$Fe^{2+} + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + Cr^{3+}$$

2. **Separate the equation into two half-reactions.**

Oxidation: $\overset{+2}{Fe^{2+}} \longrightarrow \overset{+3}{Fe^{3+}}$

Reduction: $\overset{+6}{Cr_2O_7^{2-}} \longrightarrow \overset{+3}{Cr^{3+}}$

3. **Balance the atoms other than O and H in each half-reaction.**

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$$

Balancing Redox Equations

4. For reactions in acid, add H_2O to balance O atoms and H^+ to balance H atoms.

$$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+} + 7H_2O$$

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

5. Add electrons to one side of each half-reaction to balance the charges on the half-reaction.

$$Fe^{2+} \longrightarrow Fe^{3+} + 1e^-$$

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

6. If necessary, equalize the number of electrons in the two half-reactions by multiplying the half-reactions by appropriate coefficients.

$$6Fe^{2+} \longrightarrow 6Fe^{3+} + 6e^-$$

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

Balancing Redox Equations

7. Add the two half-reactions together and balance the final equation by inspection. **The number of electrons on both sides must cancel.**

Oxidation: $6Fe^{2+} \longrightarrow 6Fe^{3+} + \cancel{6e^-}$

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

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

8. Verify that the number of atoms and the charges are balanced.

$$14x1 – 2 + 6x2 = 24 = 6x3 + 2x3$$

9. For reactions in basic solutions, add OH^- to **both sides** of the equation for every H^+ that appears in the final equation.

Test Your Skill

Balance the following redox equations

$$Fe(s) + Ag^{+} \rightarrow Ag(s) + Fe^{2+}$$

$$MnO_4^{-}(aq) + Fe^{2+} \rightarrow Fe^{3+}(s) + Mn^{2+}$$

$$H_2O_2(aq) + N_2H_4(aq) \rightarrow N_2(g) + H_2O(l)$$

In basic conditions

Check your knowledge

Q. The oxidation state of Cr in $[Cr(NH_3)_4Cl_2]^+$ is

(a) 0 (b) +1

(c) +2 (d) +3

Ans. (d)

Q. The oxidation state of chromium in the final product formed by the reaction between KI and acidified potassium dichromate solution is

(a)+3 (b) +2

(c) +6 (d) +4

Ans. (a)

Check your knowledge

Q. **Oxidation number of Cl in $CaOCl_2$ (bleaching powder) is**

(a)zero, since it contains Cl_2

(b) -1, since it contains Cl^-

(c) +1, since it contains ClO^-

(d) +1 and -1, since it contains ClO^- and Cl^-

Ans. (d)

Q. **MnO_4^- is a good oxidising agent in different medium changing to**

$$MnO_4^- \quad \uparrow Mn^{2+} \quad \uparrow MnO_4^{2-}$$

$$\uparrow MnO_2 \quad \uparrow Mn_2O_3$$

Check your knowledge

Changes in oxidation number respectively, are
(a)1,3,4,5
(b)5,4,3,2
(c)5,1,3,4
(d)2,6,4,3
Ans. (C)

Check your knowledge

Q. Which of the following is a redox reaction ?

(a) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$

(b) $CaC_2O_4 + 2HCl \rightarrow CaCl_2 + H_2C_2O_4$

(c) $Ca(OH)_2 + 2NH_4Cl \rightarrow CaCl_2 + 2NH_3 + 2H_2O$

(d) $2K[Ag(CN)_2] + Zn \rightarrow 2Ag + K_2[Zn(CN)_4]$

Ans. (d)

Thank You…