RedOx

Chapter 18

Oxidation- Reduction Reactions

- Redox or oxidation-reduction reactions are reactions that involve a transfer of electrons.
- Oxidation is the **loss** of electrons.
- Reduction is the **gain** of electrons.
- (think of the charge, OIL RIG)
- So in the reaction
- $4\ K + O_2 \rightarrow 4\ K^+ + 2\ O^{2-}$
- Potassium get oxidized, oxygen get reduced

Of course…

- We would normally write this expression
- $4\ K + O_2 \rightarrow 4\ K^+ + 2\ O^{2-}$
- as…
- $4\ K + O_2 \rightarrow 2\ K_2O$
- It doesn't change anything
- Potassium still gets oxidized, oxygen still gets reduced.

Oxidation States

- $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$
- The above reaction also involves a transfer of electrons, so it is a redox reaction.
- To see this we need to assign oxidation numbers to each atom present.
- Oxidation states (numbers)- hypothetical charge an atom would have if all bonds were ionic.
- It is used as a method to keep track of electrons in oxidation reduction reactions.

Rules for assigning Oxidation States

1. The oxidation state of an uncombined atom is 0.
2. The oxidation state of a monoatomic ion is the same as its charge.
3. The sum of the oxidation states of a neutral compound must be 0.
4. The sum of the oxidation states of a polyatomic ion must equal its charge.
5. In binary compounds, the element with the greater electronegativity is assigned a negative oxidation state equal to its charge as an ion.

Special Elements

6. Diatomic elements are assigned an oxidation state of zero if they are not bonded to anything else.
7. Alkali metals when in a compound are +1
8. Alkaline Earth metals are almost always +2
9. Halogens are normally -1 unless bonded to a more electronegative halogen or oxygen
10. Except for the above rules, Oxygen is assigned an oxidation state of -2 unless it in a compound containing peroxide (O_2^{2-}), then it gets a -1.
11. Hydrogen gets a charge of +1when bonded to nonmetals, it if is bonded to a metal it is -1.

Determining Oxidation States

- C
- F_2
- AlF_3
- CO_2
- H_2O_2
- K_3PO_4
- $NaClO_4$
- SO_4^{2-}
- $CaCl_2$
- NH_3
- NaH
- $CaSiO_3$

Determining Oxidation States

- C
  - C 0
- F_2
  - F 0
- AlF_3
  - Al +3
  - F -1
- CO_2
  - C +4
  - O -2
- H_2O_2
  - H + 1
  - O -1
- K_3PO_4
  - K +1
  - P +5
  - O -2
- $NaClO_4$
  - Na +1
  - Cl +7
  - O -2
- SO_4^{2-}
  - S +6
  - O -2
- $CaCl_2$
  - Ca +2
  - Cl -1
- NH_3
  - N -3
  - H +1
- NaH
  - Na + 1
  - H - 1
- $CaSiO_3$
  - Ca +2
  - Si +4
  - O -2

Using oxidation states

- In the reaction…
- $2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$
- 0 +1 -2 +1 -2 +1 0
- Note the changes
- Sodium went from 0 to 1
- 2 of the hydrogen atoms went from +1 to 0 (the other two were unchanged)

So…

- Sodium must have lost 2 electrons
- $2\ Na \rightarrow 2Na^+ + 2\ e^-$
- And Hydrogen gained two electrons
- $2\ H_2O + 2\ e^- \rightarrow 2\ OH^- + H_2$
- Sodium is oxidized, hydrogen is reduced in this reaction
- Oxidation is an increase in oxidation state
- Reduction is a decrease in oxidation state

In the reaction from earlier

- $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$
- -4 +1 0 +4 -2 +1 -2

- Carbon went from -4 to +4, it was oxidized
- Oxygen went from 0 to -2 it was reduced

Balancing Redox Equations by Half Reactions Method

Balancing Equations

- Redox reactions don't follow normal rules for balancing equations because we also have to pay attention of the electron transfer.
- For example
- $Ce^{4+} + Sn^{2+} \rightarrow Ce^{3+} + Sn^{4+}$
- Is it balanced?
- No, look at the charges

Half reactions

- Half reactions are exactly what the sound like, half of the reaction.
- Half reactions also include electrons, e^-, as reactants or products.
- So for
- $Ce^{4+} + Sn^{2+} \rightarrow Ce^{3+} + Sn^{4+}$
- we have
- $Ce^{4+} \rightarrow Ce^{3+}$
- $Sn^{2+} \rightarrow Sn^{4+}$

Determining # of electrons

- To determine how many electrons are added or lost look at the charges.
- For $Ce^{4+} \rightarrow Ce^{3+}$
- I have to decrease my charge by one so I must add an electron.
- $Ce^{4+} + e^- \rightarrow Ce^{3+}$
- For $Sn^{2+} \rightarrow Sn^{4+}$
- My charge increases by two, so electrons were lost.
- $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

Simple Rule

- **Electrons lost must equal electrons gained!**
- So to make this work I have to double the cerium half reaction
- **2** Ce^{4+} **+2** $e^- \rightarrow$ **2** Ce^{3+}
- $Sn^{2+} \rightarrow 2e^- + Sn^{4+}$
- Put our two half reactions back together to make a "whole" reaction again.
- $2\ Ce^{4+} + Sn^{2+} \rightarrow 2\ Ce^{3+} + Sn^{4+}$