CHAPTER 20

"Oxidation-Reduction Reactions"

LEO SAYS GER

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
  - Town in Pennsylvania has methane running through their water

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*
- $Mg + S \rightarrow Mg^{2+} + S^{2-}$ **(MgS)**

- 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:

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

LEO says GER :

Lose Electrons = Oxidation

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

Sodium is oxidized

Gain Electrons = Reduction

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

Chlorine is reduced

LEO says GER :

- 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 the reducing agent

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

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

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 <u>no change in oxidation number</u> are **NOT** redox reactions.

Examples:

$$\overset{+1}{Ag}\,\overset{+5}{N}\,\overset{-2}{O_3}(aq) + \overset{+1}{Na}\,\overset{-1}{Cl}(aq) \rightarrow \overset{+1}{Ag}\,\overset{-1}{Cl}(s) + \overset{+1}{Na}\,\overset{+5}{N}\,\overset{-2}{O_3}(aq)$$

$$2\,\overset{+1}{Na}\,\overset{-2}{O}\,\overset{+1}{H}(aq) + \overset{+1}{H_2}\,\overset{+6}{S}\,\overset{-2}{O_4}(aq) \rightarrow + \overset{+1}{Na_2}\,\overset{+6}{S}\,\overset{-2}{O_4}(aq) + \overset{+1}{H_2}\,\overset{-2}{O}(l)$$

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.**

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

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

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

(+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.**

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

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

thus X = +5

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

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

thus X = +6

Rules for Assigning Oxidation Numbers

7) The oxidation number of many elements corresponds to the element's position in the periodic table.

(a) elements in Group 1A are always +1

(b) elements in group 2A are always +2

Reducing Agents and Oxidizing Agents

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

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

Sodium is <u>oxidized</u> – it is the <u>reducing agent</u>

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

Chlorine is <u>reduced</u> – it is the <u>oxidizing</u> 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

Conceptual Problem 20.3, page 643

Technology & Society – page 644

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

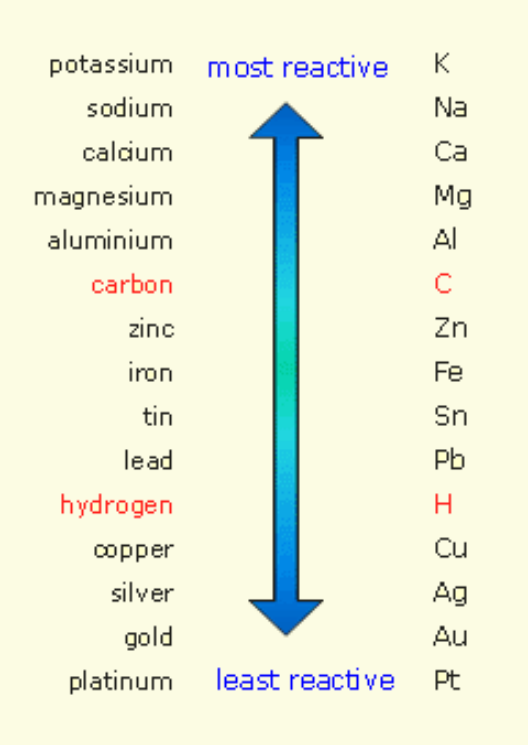
Single Replacement Reaction

- **Single Replacement reaction** – one ion switches with another ion
- Can be a redox reaction
- Ex: $Cu + 2\ AgNO_3 \rightarrow Cu(NO_3)_2 + 2\ Ag$

  Oxi # 0 -1 -1 +2 -1 0

Single Displacement → Reactivity Series

- For a single replacement reaction, only metals above will have a reaction
- Ex: $2\ Na + MgCl_2 \rightarrow$ NaCl
- Ex: Fe + KOH → no r



Single Displacement → Reactivity Series

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

Halogen Displacement

- The halogens also have a unique reactivity series
  - Reactivity decreases down a column
  - $F_2 > Cl_2 > Br_2 > I_2$
- Ex: $Cl_2(g) + 2\ KBr(aq) \rightarrow$???
  - Cl_2 is more reactive…
  - So… $Cl_2 + 2\ KBr \rightarrow 2\ KCl + Br_2\ (l)$

Identifying Redox Equations

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

$$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$$

- Is this a redox reaction? ⟶ **YES!**
  - 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.