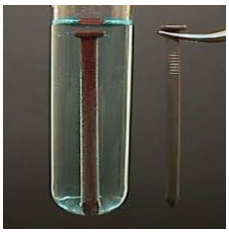
13.1 - Oxidation and Reduction

Demo: Iron is replaced by copper:

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$



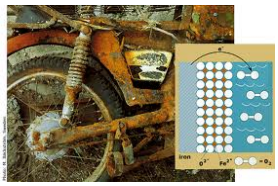
Oxidation

Colloquially, 'oxidation' means reacting with O_2.
In chemistry, oxidation = loss of electrons (symbol = e^-).
In the demo, iron lost e^-:

Half-reaction*: $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2\ e^-$

Thus, it was oxidized.

* A half-reaction shows a reactant losing or gaining electrons.



Oxidation

Reduction

Reduction: gaining e^-.

The oxidation number (charge) is reduced, hence the term reduction.

In our demo, e^- were gained by copper:

Half-reaction: $Cu^{2+}_{(aq)} + 2\ e^- \rightarrow Cu_{(s)}$

Since copper's charge lowered from 2^+ to zero: it underwent reduction.

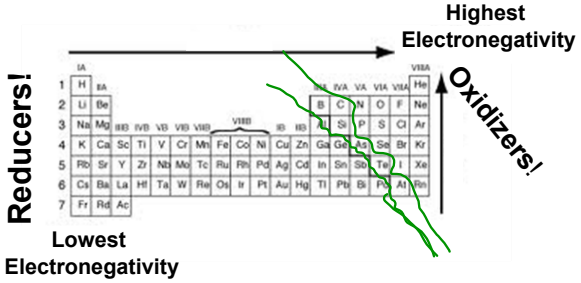
Oxidation-Reduction (Redox) Reactions

Def: Reaction involving e^- transfer.
No oxidation w/o reduction!
Oxidizing Agent (oxidizer): Oxidizes another chemical. It accepts e^-, so is reduced.

Reducing Agent (reducer): Reduces oxidizer's charge by giving it e^-. The reducer is oxidized.

Redox and Electronegativity

Electronegativity drives redox chemistry.
Metals = good reducers, Non-metals = good oxidizers.



6 Charge Rules:

1. Pure elements (polyatomic too: N_2, O_3) = 0.
2. Oxide (O^{2-}) always = -2; Fluoride (F^-) always = -1.
3. 1st group, including hydrogen = +1. Ex: $Na^+ = +1$.
   2nd group = +2. Ex: $Ca^{2+} = +2$.
4. Total charge in neutral compounds is 0.
   Ex: H_2SO_4 has net charge of 0.
5. Total charge of all ions = superscript.
   Cl^- = -1 charge. Ex: SO_4^{2-} has net charge of -2.
6. Use Ions List for ionic compounds.
   Ex: zinc sulfate ($ZnSO_4$): Zn = +2, SO_4 = -2

Chem Unit 13.1 Notes - Oxidation and Reduction

1. Accountant Method Guided Example: HNO_2

Accountants use The Bottom Line; this technique determines charges in compounds. Do HNO_2.

1. Draw 'Bottom Line': HNO_2
2. Put overall charge at end of Line: $HNO_2 = 0$
3. Place rule-based charges above elements: $\overset{+1\ ?\ -2}{HNO_2} = 0$
4. Multiply charges by subscripts: $\overset{+1\ ?\ -2}{HNO_2}$ → $+1+?-4 = 0$
5. Solve for missing information: $\overset{+1\ +3\ -2}{HNO_2}$ → $+1+3-4 = 0$

2. Oxidation Number Example

Determine each elements' oxidation number:

$\overset{3+}{Au^{3+}}$ $\quad \overset{+4\ -2}{CO_2}$ $\quad \overset{0}{Rb}$

$\overset{+5\ -2}{N_2O_5}$ $\quad \overset{+5\ -2}{ClO_3^-}$ $\quad \overset{0}{S_8}$

3. Example

$2\ K + Cl_2 \rightarrow 2\ KCl$

A. What is each element's charge?
Reactants are pure elements: charges = 0.
In KCl, K = +1 (rule 4), and Cl = -1 (Bottom Line).

$\overset{0\quad 0\quad +1\ -1}{2\ K_{(s)} + Cl_{2(g)} \rightarrow 2\ KCl_{(s)}}$

B. How many total e^- were involved?
Charges determine e^- transfer per molecule:
each K lost one e^- to Cl; total moved $e^- = 2$.

3. Example

$\overset{0\quad 0\quad +1\ -1}{2\ K_{(s)} + Cl_{2(g)} \rightarrow 2\ KCl_{(s)}}$

C. Which of the reactants was oxidized? Reduced?

K was oxidized: charge increased.

Cl went from zero to -1: reduced.

4. Robust Example

Assign charges: $\overset{0\quad 0\quad -3\ +1}{N_{2(g)} + 3H_{2(g)} \rightarrow 2\ NH_{3(g)}}$

Which reactant is the oxidizer?

Nitrogen gains 3 e^- : change = -3.
Hydrogen loses 1 e^-: change = +1.

Nitrogen = oxidizer - it was reduced (absorbed e^-).

5. New Format Example

Assign charges: $\overset{+4\ -2\quad 0\quad 0\quad +3\ -2}{3\ MnO_2 + 4\ Al \rightarrow 3\ Mn + 2\ Al_2O_3}$

Which element is reduced? Which is oxidized?

Mn gains 4 e^- (reduced); Al loses 3 e^- (oxidized).

Oxygen = no change: was spectator.

Pair changed elements:

$$\overset{+4\quad 0\quad 0\quad +3}{3\ MnO_2 + 4\ Al \rightarrow 3\ Mn + 2\ Al_2O}$$

(-4 connects Mn; +3 connects Al)

How many electrons were transferred?
The total number of transferred e^- was 12.

Homework
13.1 Problems
Due: Next Class