**General Rule of Thumb:** Spontaneous oxidation-reduction reactions convert the stronger of a pair of oxidizing agents and the stronger of a pair of reducing agents into a weaker oxidizing agent and a weaker reducing agent.

**Strongest reducing agents are at one end and (RIG)**

**\*\*\*\***Good reducing agents include the active metals;

**\*\*\*\***Metal hydrides, which formally contain the H- ion

**In your periodic Table**:

\*\*\*\*\*\*As we go down the groups, the elements get better as reducing agents. This would be expected because as we have seen the ionization energy decreases as we go down the group which means that the metals lose electrons more easily.

The reducing powers of some species decreases in the order

**Na > Zn > Pb > H2 > Cu > Ag > Br- > Mn2+ > F-**

Thus, sodium metal is the strongest reducing agent and F- is the weakest reducing agent.

**\*\*\***Good reducing agents tend to consist of atoms with a low electronegativity, the ability of an atom or molecule to attract bonding electrons, and

**\*\*\***species with relatively small ionization energies serve as good reducing agents.

**The strongest oxidizing agents are at the other. (OIL)**

**\*\*\*\***Atoms, ions, and molecules that have an unusually large affinity for electrons tend to be good oxidizing agents.

**\*\*\*\***good oxidizing agents is among compounds with unusually large oxidation states

**\*\*\*\***oxidizing agents tend to posses an atom that is initially in a significantly positive oxidation state. Common examples include **KMnO4, CrO3, H2Cr2O7 and OsO4.**

For example: In KMnO4 an ionic compound composed of K+ and [MnO4]-, the oxidation state on Mn is + 7.

|  |  |
| --- | --- |
| **Oxidizing agent** | **Symbol** |
| Oxygen | O2 |
| Fluorine | F2 |
| Ozone | O3 |
| Chlorine | Cl2 |
| Iodine | I2 |
| Bromine | Br2 |
| Nitric Acid | HNO3 |
| Chlorate | ClO3- |
| Hypochlorite | OCl- |
| Dichromate | Cr2O72- |
| Chromate | CrO42- |
| Chromium Trioxide | CrO3 |
| Permanganate | MnO4- |
| Manganate | MnO42- |

**Exceptions to the Rule:**

**\*\*\*\***Some compounds can act as either oxidizing agents or reducing agents. One example is hydrogen gas, which acts as an oxidizing agent when it combines with metals and as a reducing agent when it reacts with nonmetals.

The common oxidation-reduction half-reactions have been organized into a table:

Furthermore, by convention**, *the strongest reducing agents are usually found at the top*** of the table.

***The Relative Strengths of Common Oxidizing Agents and Reducing Agents***

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
|  |  | K+ + e- <---->K | |  | | Best |
|  |  | Ba2+ + 2 e- <---->Ba | |  | | reducing |
|  |  | Ca2+ + 2 e- <---->Ca | |  | | agents |
|  |  | Na+ + e- <---->Na |  | | **RIG** (**R**eduction **I**s **L**oss) | |
|  |  | Mg2+ + 2 e- <---->Mg | |  | |  |
|  |  | H2 + 2 e- <---->2 H- | |  | |  |
|  |  | Al3+ + 3 e- <---->Al | |  | |  |
|  |  | Mn2+ + 2 e- <---->Mn | |  | |  |
|  |  | Zn2+ + 2 e- <---->Zn | |  | |  |
|  |  | Cr3+ + 3 e- <---->Cr | |  | |  |
|  |  | S + 2 e- <---->S2- | |  | |  |
|  |  | 2 CO2 + 2 H+ + 2 e- <---->H2C2O4 | |  | |  |
|  |  | Cr3+ + e- <---->Cr2+ | |  | |  |
|  |  | Fe2+ + 2 e- <---->Fe | |  | |  |
|  |  | Co2+ + 2 e- <---->Co | |  | |  |
|  |  | Ni2+ + 2 e- <---->Ni | |  | |  |
|  |  | Sn2+ + 2 e- <---->Sn | |  | |  |
|  |  | Pb2+ + 2 e- <---->Pb | |  | |  |
|  |  | Fe3+ + 3 e- <---->Fe | |  | |  |
|  |  | 2 H+ + 2 e- <---->H2 | |  | |  |
|  |  | S4O62- + 2 e- <---->2 S2O32- | |  | |  |
|  |  | Sn4+ + 2 e- <---->Sn2+ | |  | |  |
|  |  | Cu2+ + e- <---->Cu+ | |  | |  |
|  |  | O2 + 2 H2O + 4 e- <---->4 OH- | |  | |  |
|  |  | Cu+ + e- <---->Cu | |  | |  |
|  |  | I2 + 2 e- <---->2 I- | |  | |  |
| oxidizing |  | MnO4- + 2 H2O + 3 e- <---->MnO2 + 4 OH- | |  | |  |
| power |  | O2 + 2 H+ + 2 e- <---->H2O2 | |  | | Reducing |
| increases |  | Fe3+ + e- <---->Fe2+ | |  | | power |
|  |  | Hg22+ + 2 e- <---->2 Hg | |  | | increases |
|  |  | Ag+ + e- <---->Ag | |  | |  |
|  |  | Hg2+ + 2 e- <---->Hg | |  | |  |
|  |  | H2O2 + 2 e- <---->2 OH- | |  | |  |
|  |  | HNO3 + 3 H+ + 3 e- <---->NO + 2 H2O | |  | |  |
|  |  | Br2(*aq*) + 2 e- <---->2 Br- | |  | |  |
|  |  | 2 IO3- + 12 H+ + 10 e- <---->I2 + 6 H2O | |  | |  |
|  |  | CrO42- + 8 H+ + 3 e- <---->Cr3+ + 4 H2O | |  | |  |
|  |  | Pt2+ + 2 e- <---->Pt | |  | |  |
|  |  | MnO2 + 4 H+ + 2 e- <---->Mn2+ + 2 H2O | |  | |  |
|  |  | O2 + 4 H+ + 4 e- <---->2 H2O | |  | |  |
|  |  | Cr2O72- + 14 H+ + 6 e- <---->2 Cr3+ + 7 H2O | |  | |  |
|  |  | Cl2(*g*) + 2 e- <---->2 Cl- | |  | |  |
|  |  | PbO2 + 4 H+ + 2 e- <---->Pb2+ + 2 H2O | |  | |  |
|  |  | MnO4- + 8 H+ + 5 e-<----> Mn2+ + 4 H2O | |  | |  |
|  |  | Au+ + e- <---->Au | |  | |  |
| **OIL** (**O**xidation **I**s **L**oss) |  | H2O2 + 2 H+ + 2 e- <---->2 H2O | |  | |  |
|  |  | Co3+ + e- <---->Co2+ | |  | |  |
| Best |  | S2O82- + 2 e- <---->2 SO42- | |  | |  |
| oxidizing |  | O3(*g*) + 2 H+ + 2 e- <---->O2(*g*) + H2O | |  | |  |
| Agents |  | F2(*g*) + 2 H+ + 2 e- <---->2 HF(*aq*) | |  | |  |

**DON’T GET CONFUSED:** The species causing the oxidation (electron loss) is the oxidizing agent and the species causing the reduction (electron gain) is the reducing agent.

Some times oxidizing agents are simply referred to as oxidizers, and reducing agents may also be called reducers.

“As the redox reaction proceeds the oxidizing agent is reduced and the reducing agent is oxidized.”