Balanced chemical reactions are the *math* of chemistry.

They show the *relationship* between the reactants and the products.

We will use thermodynamics later on to calculate the *feasibility* of reactions and to understand how equilibrium is established.

The concept of *equilibrium* allows us to understand chemical processes such as ionic speciation, oxidation state distributions, gas solubility, the carbonate system ……
A. Oxidation state or number

Oxidation: Loss of electrons from an element.
Oxidation number increases

Reduction: Gain of electrons by an element.
Oxidation number decreases

If we want to determine whether reaction is oxidation or reduction

Need to know oxidation number of the element and how it changes
<table>
<thead>
<tr>
<th>Element</th>
<th>Electron Configuration</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>1s¹</td>
</tr>
<tr>
<td>Helium</td>
<td>1s²</td>
</tr>
<tr>
<td>Lithium</td>
<td>1s²2s²</td>
</tr>
<tr>
<td>Beryllium</td>
<td>1s²2p²</td>
</tr>
<tr>
<td>Boron</td>
<td>1s²2p²2p¹</td>
</tr>
<tr>
<td>Carbon</td>
<td>1s²2s²2p²</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>1s²2s²2p²2p³</td>
</tr>
<tr>
<td>Oxygen</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>Fluorine</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>Neon</td>
<td>1s²2s²2p⁶</td>
</tr>
<tr>
<td>Sodium</td>
<td>1s²2s²2p⁶2p³¹</td>
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<tr>
<td>Magnesium</td>
<td>1s²2s²2p⁶2p³¹2p²</td>
</tr>
<tr>
<td>Aluminium</td>
<td>1s²2s²2p⁶2p³¹3s¹</td>
</tr>
<tr>
<td>Silicon</td>
<td>1s²2s²2p⁶2p³¹3s²</td>
</tr>
<tr>
<td>Phosphorus</td>
<td>1s²2s²2p⁶2p³¹3s²3p³¹</td>
</tr>
<tr>
<td>Sulfur</td>
<td>1s²2s²2p⁶2p³¹3s²3p³¹3p⁴</td>
</tr>
<tr>
<td>Chlorine</td>
<td>1s²2s²2p⁶2p³¹3s²3p³¹3p⁴3p⁵</td>
</tr>
<tr>
<td>Argon</td>
<td>1s²2s²2p⁶2p³¹3s²3p³¹3p⁴3p⁵</td>
</tr>
</tbody>
</table>
Rules for determining oxidation number of an element

1. **Oxidation state of an element in its elementary state = 0**
   e.g. Cl₂, Na, P,….etc.

2. **Oxidation state of an element in a monatomic (only one atom) ion is equal to the charge on the ion**
   e.g. Na⁺ = +1; Cl⁻ = -1; Fe³⁺ = +3

3. **Oxidation state of certain elements is the same in all, or almost all of their compounds**
   e.g. Group 1A elements: Li, Na, K, Rb, Cs =+1
   Group 2A elements: Be, Mg, Ca, Sr, Ba, Ra = +2
   Group VII b elements: F, Cl, Br, I, At = -1 in binary compounds
   Oxygen is almost always -2 (Except: when bonded to O or F)
   H is almost always +1; Except with a metal, e.g. NaH, CaH₂, is -1

4. **The sum of the oxidation states in a neutral species is = 0;**
   In a charged ion it is equal to the charge on the ion
   e.g. Na₂Se: Na = +1x2 = 2, thus Se = -2
   MnO₄⁻: O= -2x4 = -8, thus Mn =8-1 = 7
5. Fractional oxidation numbers are possible. E.g., in Na$_2$S$_4$O$_6$ (sodium tetrathionate), S has an oxidation number of +10/4:

O: 6(-2) = -12  
Na: 2(+1) = 2  
Residual = -10, which must be balance by S:  
S: 4(+10/4) = +10

6. The oxidation number is designated by:
- Arabic number below the atom, or
- Roman numeral or Arabic number after the atom (in parentheses)

What is the oxidation state of the elements in KNO$_3$?

K= ?; O= ?; N=?

K$_2$Cr$_2$O$_7$? K= ?; O= ?, Cr= ?

Oxidation is an increase in oxidation state  
e.g. Cl$^-$ to Cl$_2$ is -1 to 0

Reduction is a decrease in oxidation state  
SO$_4^{2-}$ to H$_2$S is a reduction S VI to S -II

Recognising oxidation/reduction  
KMnO$_4$ to MnO$_2$ oxidation or reduction?
B. Balancing oxidation-reduction reactions

Conventionally always put the oxidised species on the left, the reduced species on the right

\[ \text{e.g. } \text{MnO}_4^- \text{(aq)} + \text{Cl}^- \text{(aq)} = \text{Mn}^{2+} \text{(aq)} + \text{Cl}_2 \text{(g)} \]

1. Separate the reaction into a reduction and oxidation part

\[ \text{MnO}_4^- \text{(aq)} = \text{Mn}^{2+} \text{(aq)} \quad \text{reduction} \]
\[ \text{Cl}^- \text{(aq)} = + \text{Cl}_2 \text{(g)} \quad \text{oxidation} \]

2. Balance each 1/2 reaction with respect to mass then with respect to charge. Use e\(^-\), H\(^+\), H\(_2\)O or OH\(^-\)

\[ 2\text{Cl}^- \text{(aq)} = \text{Cl}_2 \text{(g)} \quad \text{mass} \]
\[ 2\text{Cl}^- \text{(aq)} = 2\text{Cl}_2 \text{(g)} + 2\text{e}^- \quad \text{mass + charge} \quad A \]

\[ \text{MnO}_4^- \text{(aq)} = \text{Mn}^{2+} \text{(aq)} + 4\text{H}_2\text{O} \text{ (mass oxygen)} \]
\[ \text{MnO}_4^- \text{(aq)} + 8\text{H}^+ = \text{Mn}^{2+} \text{(aq)} + 4\text{H}_2\text{O} \text{ (mass oxygen+ hydrog)en} \]
\[ \text{MnO}_4^- \text{(aq)} + 8\text{H}^+ + 5\text{e}^- = \text{Mn}^{2+} \text{(aq)} + 4\text{H}_2\text{O} \text{ (mass + charge)} \quad B \]

3. Combine half reactions so electron gain equals loss

\[ 5\ A = 10 \text{ e}^-; \ 2\ B = 10 \text{ e}^- \ i.e. \ 5\ A +2*B \]

\[ 10 \text{Cl}^- \text{(aq)} + 2\text{MnO}_4^- + 16\text{H}^+ = 5\text{Cl}_2 \text{(g)} + 2\text{Mn}^{2+} \text{(aq)} + 8\text{H}_2\text{O} \]

4. Check for atom and charge balance
C. Oxidation of organic matter

\[ \text{CH}_2\text{O} + \text{SO}_4^{2-} + \text{H}_2\text{O} = \text{CO}_2 + \text{H}_2\text{S} \]

0 \quad \text{IV} \quad \text{ox state of C}

How did we get the oxidation state of the C in CH\(_2\)O?

Separate into oxidation and reduction half reactions

\[ \text{SO}_4^{2-} + 8\text{e}^- + 10\text{H}^+ = \text{H}_2\text{S} + 4\text{H}_2\text{O} \quad \text{A} \]

\[ \text{CH}_2\text{O} + \text{H}_2\text{O} = \text{CO}_2 + 4\text{H}^+ + 4\text{e}^- \quad \text{B} \]

Combine so electrons balance:

\[ \text{A} + \text{B} \times 2 \]

\[ 2 \text{CH}_2\text{O} + \text{SO}_4^{2-} + 2\text{H}_2\text{O} + 10\text{H}^+ = \text{H}_2\text{S} + 4\text{H}_2\text{O} + 2\text{CO}_2 + 8\text{H}^+ \]

Simplify by subtracting 8 H\(^+\) and 2 H\(_2\)O from each side.

\[ 2 \text{CH}_2\text{O} + \text{SO}_4^{2-} + 2\text{H}^+ = \text{H}_2\text{S} + 2\text{H}_2\text{O} + 2\text{CO}_2 \]

This reaction is the oxidation of organic matter through the reduction of sulphate, you will see this reaction later in reducing sediments.
D. An example in basic solution:

\[ I^- + MnO_4^- = I_2 + MnO_2 \]

Oxidation: \( 2I^- = I_2 + 2e^- \)  \( \text{A} \)

Reduction: \( MnO_4^- + 4H^+ + 3e^- = MnO_2 + 2H_2O \)

remove \( H^+ \) by adding \( OH^- \) to each side \(* 4 \) \( (4H^+ + 4OH^- = 4H_2O) \)

\[ MnO_4^- + 4H_2O^+ + 3e^- = MnO_2 + 2H_2O + 4OH^- \]

simplify: \( MnO_4^- + 2H_2O + 3e^- = MnO_2 + 4OH^- \)  \( \text{B} \)

combine so electrons balance: \( A \ast 3 + B \ast 2 \)

\[ 6I^- + 2MnO_4^- + 4H_2O = 3I_2 + 2MnO_2 + 8OH^- \]

E. A weathering reaction.

\( Fe_2SiO_4 + O_2 = Fe_2O_3 + FeSiO_3 \)

II  \underline{\text{III}}  II ox state of Fe i.e. Fe is oxidised

Iron Olivine  =  Haematite + Ferrosilite ( Fe pyroxene)

Separate and balance for mass and charge

\[ O_2 + 4e^- = 2O^{2-} \]  \( \text{reduction A} \)

\[ 2Fe_2SiO_4 + H_2O = Fe_2O_3 + 2e^- + 2FeSiO_3 + 2H^+ \]  \( \text{oxidation B} \)

Note have added \( H_2O \) on LH side

Combine eqns balancing \( e^- \) \( A \ast 2 \ast B \)

\[ 4Fe_2SiO_4 + 2H_2O + O_2 + 4e^- = 2Fe_2O_3 + 4e^- + 4FeSiO_3 + 4H^+ + 2O^{2-} \]

cancel \( 4e^- \) on each side then cancel as \( 2H_2O = 4H^+ + 2O^{2-} \)

\[ 4Fe_2SiO_4 + O_2 = 2Fe_2O_3 + 4FeSiO_3 \]

Removal of oxygen by oxidation of reduced iron compounds
Example where the same compound is being oxidised and reduced:

\[ \text{Cl}_2 + \text{H}_2\text{O} = \text{HOCl} + \text{H}^+ + \text{Cl}^- \]
\[ \text{Cl}_2 + 2\text{e}^- = 2\text{Cl}^- \text{ (reduction of Cl)} \]
\[ \text{Cl}_2 + 2\text{H}_2\text{O} = 2 \text{ HOCl} + 2\text{H}^+ + 2\text{e}^- \text{ (oxidation of Cl)} \]
\[ 2\text{Cl}_2 + 2\text{H}_2\text{O} = 2 \text{ HOCl} + 2\text{H}^+ + 2 \text{ Cl}^- \]

This may have been what happened to Cl\(_2\) released from the degassing of the early Earth.