Chapter 14 - The Elements: The First Four Main Groups

- Periodic Trends
- Hydrogen
- Group 1/I: Alkali Metals
- Group 2/II: The Alkaline Earth Metals
- Group 13/III: The Boron Family
- Group 14/IV: The Carbon Family
Effective Nuclear Charge: \( Z_{\text{eff}} \) The net nuclear charge after taking into account the shielding caused by other electrons in the atom.

Why: Going across the periodic table does not add more orbital's it only allows for electrons to enter into a preexisting orbital. The electrons in an orbital are spread out due to electron electron repulsion therefore the electrons that enter into the orbital do not shield the nucleus adequately and the effective nuclear charge goes up across a row. However going down a group adds more electron orbits which shield the nucleus more effectively and the effective nuclear charge goes down.

Example:
Which has the bigger effective nuclear charge? Li or F F or I
**Atomic Radii**: Half the distance between the centers of neighboring atoms in a solid or a homonuclear molecule.

**Why**: Going across a period the effective nuclear charge increases therefore the pull on the electrons increases and the atomic radii decrease. Going down a group the effective nuclear charge decreases therefore the atomic radii increases.

**Example**: Which has the bigger atomic radii? Li or F  F or I
**First Ionization Energy:** The minimum energy required to remove the first electron from the ground state of a gaseous atom, molecule, or ion.

**Why:** Going across a period the effective nuclear charge increases therefore it is harder to remove an electron and the first ionization energy increases. However, going down a group the effective nuclear charge decreases causing the first ionization energy to also decrease.

**Example:**
Which has the bigger first ionization energy? Li or F? F or I?
**Electron Affinity**: \( E_{\text{ea}} \) The energy released when an electron is added to a gas-phase atom.

**Why**: Going across a period the effective nuclear charge increases therefore the atom has a larger positive charge and releases more energy when an electron is added to the atom. Going down a group the effective nuclear charge decreases and therefore the atom has a smaller positive charge and the electron affinity decreases.

**Example:**
Which has the bigger electron affinity? Li or F? F or I?

**Note**: This trend has the most atoms that do not obey the trend.
**Electronegativity**: \((\chi)\) The ability of an atom to attract electrons to itself when it is part of a compound.

**Why**: Going across a period the effective nuclear charge increases therefore the atom has a larger positive charge and attracts more electrons to itself in a compound causing the electronegativity to increase. Going down a group the effective nuclear charge decreases and therefore the atom has a smaller positive charge causing the electronegativity to decrease.

**Example:**
Which has the bigger electronegativity? Li or F

**Note**: Bonds between things of similar electronegativities tend to be covalent.
**Polarizability**: $\alpha$ The ease with which the electron cloud of a molecule can be distorted.

**Why**: Going across a period the effective nuclear charge increases therefore the electrons are held tighter to the nucleus and are unable to deform when bonded with other atoms. However, going down a group the effective nuclear charge decreases and the electrons are not held as tightly to the nucleus and therefore, deform more easily when bonding with other atoms.

**Example:**
Which is easier to polarize? Li or F  F or I
Student Question:
Which atom is larger?
A) Thallium (Tl)  B) Lead (Pb)

Which atom is more electropositive?
A) Potassium (K)  B) Rubidium (Rb)

Which atom has the greatest electron affinity?
A) Arsenic (As)  B) Fluorine (F)  C) Sulfur (S)
Most main group elements form the same number of bonds as the oxidation number.

Elements in period three and higher have access to the empty $d$ orbitals and can use them to expand their valence shells past the usual octet of $e^-$ and therefore do not always follow this rule.

<table>
<thead>
<tr>
<th>Element</th>
<th>Number of Valence e\textsuperscript{-}</th>
<th>Typical Oxidation Number</th>
<th>Typical Number of Bonds Formed</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>1</td>
<td>+1</td>
<td>1</td>
<td>NaCl(s)</td>
</tr>
<tr>
<td>O</td>
<td>6</td>
<td>-2</td>
<td>2</td>
<td>H\textsubscript{2}O(l)</td>
</tr>
<tr>
<td>F</td>
<td>7</td>
<td>-1</td>
<td>1</td>
<td>HF(l)</td>
</tr>
</tbody>
</table>
The smaller the size of the atom the fewer the other atoms that can bond with it.

In general only period 2 elements form multiple bonds with themselves or other elements in the same period because only they are small enough for their p orbitals to have substantial $\pi$ overlap.
**Periodic Trends**

**Bonding Trends (Hydrides)**

**Hydrides**: Compounds that contain hydrogen.

Most elements in the main group form binary compounds with hydrogen that reflect their location on the periodic table.

\[
\text{NH}_3 \quad \text{CH}_4 \quad \text{H}_2\text{O} \quad \text{HF}
\]

The nature of the binary hydride is related to the characteristics of the element bonded to the hydrogen.

Three different classifications of binary hydrides:
- Saline Hydrides
- Metallic Hydrides
- Molecular Hydrides
Saline hydrides are formed by the members of the $s$ block when they are heated in the presence of $H_2(g)$.

**Example:**

$$2K(s) + H_2(g) \xrightarrow{\Delta} 2KH(s)$$

**Properties:**
- White
- High melting point
- Crystal structure similar to halides (rock salt structure)
- Ionic Bonds
Metallic hydrides are formed by heating certain $d$ block metals in the presence of $H_2(g)$

Example:

$$2Cu(s) + H_2(g) \xrightarrow{\Delta} 2CuH(s)$$

Properties:
- Black
- Powdery
- Electrically conductive solids

Metal hydrides release their H when heated or treated with acid therefore they are being investigated for storing and transporting hydrogen.
Molecular hydrides are formed when nonmetals form covalent bonds with hydrogen.

Example:
HF, HCl, HBr

Properties:
- Volatile
- Many are Bronsted acids
All main group elements except for the noble gases react with oxygen to form oxides.

<table>
<thead>
<tr>
<th>Oxides formed from atoms on the…</th>
<th>Left side of the periodic table (s and most of the d block)</th>
<th>Left side of the p block</th>
<th>Right side of the periodic table</th>
</tr>
</thead>
<tbody>
<tr>
<td>Soluble in H₂O(l)</td>
<td>Insoluble</td>
<td>Low melting point</td>
<td></td>
</tr>
<tr>
<td>Ionic</td>
<td>High melting points</td>
<td>Often gaseous (NO(g))</td>
<td></td>
</tr>
<tr>
<td>Tend to be strong bases</td>
<td>Tend to be amphoteric</td>
<td>Tend to be Lewis acids</td>
<td></td>
</tr>
</tbody>
</table>
**Periodic Trends**

**Bonding Trends (Anhydrides)**

**Acid Anhydrides:** A compound that forms an oxoacid (an acid that contains oxygen) when it reacts with water.

**Example:**

\[ \text{N}_2\text{O}_5(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow 2\text{HNO}_3(\text{l}) \]

\[ \text{N}_2\text{O}_5(\text{l}) \text{ is the anhydride} \]

\[ \text{SO}_3(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{SO}_4(\text{l}) \]

\[ \text{SO}_3(\text{l}) \text{ is the anhydride} \]

**Formula Anhydrides:** A compound that has the formula of an acid minus the elements of water but does not react with water to produce the acid.

**Example:**

\[ \text{CO(}g\text{)} + \text{H}_2\text{O(}l\text{)} \text{ does not go to } \text{HCOOH(}l\text{)} \]

\[ \text{CO(}g\text{)} \text{ is the formula anhydride} \]
Hydrogen

The Element

Properties:
- Colorless
- Odorless
- Tasteless

- Electron configuration is $1s^1$ (similar to the electron configurations of group 1 elements)
- Classified as a non metal
- Therefore it doesn’t fit into any group
The elemental form of H is H$_2$

H$_2$ is small and nonpolar so the H atoms can only attract each other through weak London forces.

**London Forces:** The force of attraction that arises from the interaction between instantaneous electric dipoles on neighboring molecules.
Most H is made up of only two particles (an electron and a proton)
- H is the most abundant element in the universe and accounts for 89% of all atoms.
- Little free H on earth
- H₂ gas is so light that it moves very fast and can escape the earth's gravitational pull
- Need heavier planets to confine H₂
Most commercial $H_2(g)$ is obtained as a by product of petroleum refining.

**Example:**

\[ CH_4(g) + H_2O(g) \xrightarrow{NiCatalyst} CO(g) + H_2(g) \]
Can’t Separate CO from $H_2$

\[ CO(g) + H_2O(g) \xrightarrow{Fe/CuCatalyst} CO_2(g) + H_2(g) \]
Membranes that can separate $CO_2$ from $H_2$

**Hydrocarbon:** Compounds that contain $H$ and $C$

1/3 of the $H$ produced is used for hydrometallurgical extractions of copper and other materials since $H$ is a strong reducing agent.
Hydrogen

The Element

Hydrogen as a fuel source

- Light (low density)
- Clean Burning
- Plenty of abundant H in H₂O

Example:

\[ 2\text{H}_2\text{O(l)} \xrightarrow{\text{light}} \text{O}_2(\text{g}) + 2\text{H}_2(\text{g}) \] (electrolysis of water)

Problem: The electrolysis of water requires a lot of energy in the form of electricity
Hydrogen can form both cations ($\text{H}^+$) and anions ($\text{H}^-$).

- Hydrogen has an intermediate electronegativity.
- Forms covalent bonds with both nonmetals and metalloids.

**Anion**

The hydride ion is very large.

The large radius makes it highly polarizable in compounds since it is hard for the single proton to control the two electrons.
Hydrogen Compounds of H (Hydrogen Bonding)

- Small elements
- Between H and highly electronegative atoms (ex: N, O, and F)
- 5% as strong as covalent bonds (between the same atoms)
- Coulombic interactions between the partially positive charge on a hydrogen atom and the partially negative charge of another atom form the H bond

\[ \text{O}^{\delta^-} \cdots \delta^+ \text{H-O} \]
Group 1: The Alkali Metals

Electron configuration is ns¹ (n is the period number)

Classified as metals

Properties:
- Soft
- Lustrous metals
Group 1: The Alkali Metals

The Elements

- Properties are dominated by the fact that they lose their $e^-$ easily
- Most Violently reactive of all the metals
- React strongly with $H_2O(l)$ the vigor of the reaction increase down the group (ex: $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$)
- The alkali metals are all too easily oxidized to be found in their free state in nature
- Great reducing agents
Group 1: The Alkali Metals

The Elements

- Not easily extracted from their ores. Have to use electrolysis of their molten salts (ex. Na in the Downs process)
- Low melting points
- Low boiling points
- Low densities
- Most form ionic compounds in nature
- Alkali metals react directly with almost all nonmetals (except the noble gasses)

Example:

\[ 6\text{Li(s)} + \text{N}_2(g) \rightarrow 2\text{Li}_3\text{N(s)} \]
Group 1: The Alkali Metals

Compounds of Lithium

Lithium differs slightly from the other element in the group

- Small size if the Li⁺ cation
  - Strong polarizing power
  - Forms bonds with highly covalent character

Uses:
  - Ceramics
  - Lubricants
  - Medicine (lithium carbonate (treatment for bipolar disorder))
Group 1: The Alkali Metals

Compounds of Sodium

**Importance of Sodium compound:**
- Low cost
- High solubility in water

**NaCl (Sodium Chloride commonly know as table salt)**

**Methods of Obtaining NaCl**
- Mined as rock salt which is a deposit of sodium chloride left as ancient oceans evaporated
- Obtained from the evaporation of brine (sea water)

**Uses:**
Electrolytic production of chlorine and sodium hydroxide from brine

\[ 2\text{Na}^+(aq) + 2\text{Cl}^-(aq) + 2\text{H}_2\text{O}(l) \rightarrow \text{Cl}_2(g) + 2\text{OH}^-(aq) + \text{H}_2(g) + 2\text{Na}^+(aq) \]
Group 1: The Alkali Metals

Compounds of Sodium

NaOH (Sodium Hydroxide commonly know as lye)

**Properties:**
- Soft
- Waxy
- White
- Corrosive solid

Methods of Obtaining NaOH
- Electrolysis of brine

**Uses:**
- Inexpensive starting material for the production of other sodium salts
Group 1: The Alkali Metals

Compounds of Sodium

NaHCO₃ (Sodium Hydrogen Carbonate (Sodium Bicarbonate) commonly called baking soda)

How baking soda works
The hydrogen carbonate reacts with a weak acid (HA) that is present in the batter (sour milk, buttermilk, lemon juice, vinegar…..)

HCO₃⁻(aq) + HA(aq) → A⁻(g) + H₂O(l) + CO₂(g)

The CO₂(g) produced causes the batter to rise

Baking powder contains a solid weak acid as well as the hydrogen carbonate therefore CO₂(g) is released when water is added
Group 1: The Alkali Metals

Common Reactions

**Reaction with Halogens**

$2M + X_2 \rightarrow 2MX$  \hspace{1cm} $X_2$ is any group 9 molecule

**Reactions with Oxygen**

$4Li + O_2 \rightarrow 2Li_2O$  \hspace{1cm} Need excess Oxygen

$2Na + O_2 \rightarrow Na_2O_2$

$M + O_2 \rightarrow MO_2$  \hspace{1cm} $M = K, Rb, or Cs$

**Reaction with H**

$2M + H_2 \rightarrow 2MH$

**Reaction with N**

$6Li + N_2 \rightarrow 2Li_3N$  \hspace{1cm} Li only

**Reaction with Water**

$2M + 2H_2O \rightarrow 2MOH + H_2$

**Reaction with Ions**

$2M + 2H^+ \rightarrow 2M^+ + H_2$
Group 1: The Alkali Metals

Compounds of Potassium

- More expensive than Na compounds
- Similar properties to Na compounds
- Less hygroscopic (water absorbing than corresponding Na compounds)
- Principle mineral source of K is $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$ and $\text{KCl}$

$\text{KNO}_3$ (Potassium Nitrate)

Used to facilitate the ignition of matches by releasing $\text{O}_2$ when heated

$$2\text{KNO}_3(s) \xrightarrow{\Delta} 2\text{KNO}_2(s) + \text{O}_2(g)$$
Group 2: The Alkali Earth Metals

The Elements

- Electron configuration is ns² (n is the period number)
- Classified as metals
- All Group 2 elements are too reactive to occur in the uncombined state in nature
- Usually found as doubly charged cations
- All Group 2 elements except for beryllium react with water and the vigor of the reaction increases going down the group
Group 2: The Alkali Earth Metals

The Elements (Beryllium)

- Has some non metal tendencies
- Mainly found in the form $3\text{BeO} \cdot \text{Al}_2\text{O}_3 \cdot 6\text{SiO}_2$ (these crystals can weigh several tons)
- The gemstone emerald contains Be but its green color is caused by $\text{Cr}^{3+}$ ions
- Used as windows for x-ray tubes (thin sheets of the metal are transparent to x-rays)
- Obtained by the electrolytic reduction of molten beryllium chloride
Group 2: The Alkali Earth Metals

The Elements (Magnesium)

- Occurs in sea water as the mineral dolomite $\text{CaCO}_3 \cdot \text{MgCO}_3$
- Mg is present in the chlorophyll molecule therefore enables photosynthesis
- Protective oxide forms which protects Mg from extensive oxidation from air
- Used in the manufacturing of airplanes due to the fact that it is a light and tough
- Obtained by either chemical or electrolytic reduction of its compounds
Group 2: The Alkali Earth Metals

The Elements (Calcium)

- Also found in sea water
- Ca is the element of rigidity and construction (bones, shells, concrete, mortar, limestone (buildings)…)
- Obtained by electrolysis or by reduction with aluminum in a version of the thermite process (same for strontium and barium)

\[
3\text{CaO}(s) + 2\text{Al}(s) \overset{∆}{→} \text{Al}_2\text{O}_3(s) + 3\text{Ca}(s)
\]
Alkali earth metals can often be identified by the color they give off in a flame.

<table>
<thead>
<tr>
<th>Group 2 Element</th>
<th>Flame Color</th>
</tr>
</thead>
<tbody>
<tr>
<td>Calcium</td>
<td>Orange-Red</td>
</tr>
<tr>
<td>Strontium</td>
<td>Crimson</td>
</tr>
<tr>
<td>Barium</td>
<td>Yellow-Green</td>
</tr>
</tbody>
</table>

Due to their colors, salts of these elements are often used in fireworks.
Group 2: The Alkali Earth Metals

Compound of Beryllium

- Beryllium compounds are very toxic
- Dominated by the highly polarizing character of the Be\(^{2+}\) ion and small size
- Highly polarizing character causes the formation of bonds that have strong covalent character
- Small size limits the number of groups that can attach to it (4)
- The structural unit is commonly tetrahedral

\[
\text{BeO(s) + C(s) + Cl}_2\text{(s)} \xrightarrow{600–800^\circ\text{C}} \text{BeCl}_2\text{(g)} + \text{CO(g)}
\]

The Be atom in the BeCl\(_2\) act as Lewis acids and accept electrons pairs form the Cl atoms of the neighboring BeCl\(_2\) groups forming a chain of tetrahedral BeCl\(_4\) units in the solid.
Compounds primarily have ionic bonds but still have some covalent character.

**Mg(OH)\(_2\)** (*Magnesium Hydroxide commonly called Milk of Magnesia)*

Because Mg(OH)\(_2\) is relatively insoluble in water. It is not absorbed into the stomach and stays in the stomach a long time to react with whatever acid is present.

**Properties:**
- Not very soluble in H\(_2\)O
- Liquid
- Collide suspension

When Mg(OH)\(_2\) neutralizes stomach acid it produces MgCl\(_2\) which is a laxative therefore Mg(OH)\(_2\) should be used sparingly.
Group 2: The Alkali Earth Metals

Compound of Magnesium

\[ \text{MgSO}_4 \, (\text{Magnesium Sulfate commonly called Epsom Salts}) \]

Is another laxative. The magnesium ions inhibit the absorption of water in the intestines thereby increasing the flow of water through the intestines.

**Chlorophyll**

- Captures light from the sun and channels its energy into photosynthesis
- The role of the \( \text{Mg}^{2+} \) ion is to keep the ring rigid thereby insuring that the photon is not lost as heat before the chemical reaction occurs
CaCO$_3$ (Calcium Carbonate)

- Most common calcium compound
- Occurs naturally in chalk and limestone

CaCO$_3$(s) decomposes into CaO(s) (lime or quicklime) when heated

CaCO$_3$(s) $\xrightarrow{\Delta}$ CaO(s) + CO$_2$(g)

CaO is called quicklime because the reaction with water is fast and extremely exothermic.

CaO(s) + H$_2$O(l) $\rightarrow$ Ca$^{2+}$(aq) + 2OH$^-$(g)

Ca(OH)$_2$ is known as slacked lime because the thirst of lime for water has been quenched or slacked
Group 2: The Alkali Earth Metals

Compound of Calcium

Uses of CaO:
Used in iron making. CaO is an Lewis base and reacts with silica in the iron ore to transform it into liquid slag

\[
\text{CaO}(s) + \text{SiO}_2(s) \xrightarrow{\Delta} \text{CaSiO}_3(l)
\]

About 50 kg of lime is needed to produce 1 ton of iron

Uses of Ca(OH)\(_2\):
Used as an inexpensive base in industry as well as to adjust the pH of soils in agriculture and to remove Ca\(^{2+}\) from hard water containing Ca(HCO\(_3\))\(_2\)

\[
\text{HCO}_3^{-}(aq) + \text{OH}^{-}(aq) \rightarrow \text{CO}_3^{2-}(aq) + \text{H}_2\text{O}(l)
\]

\[
\text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{CaCO}_3(s)
\]
Concrete

Strong building material made from bonder and a filler

Filler:
Usually gravel and sand

Bonder:
Made by heating calcium oxide, calcium silicates, and calcium aluminum silicates.
The pellets are ground together with gypsum (CaSO$_4$·2H$_2$O)

When the filler, bonder, and water are mixed together the water reacts to form hydrates (compounds containing H$_2$O) and hydroxides (compounds containing OH$^-$) which bind the salts together in a three dimensional network
Group 2: The Alkali Earth Metals

Common Reactions

**Reaction with Halogens**
\[ M + X_2 \rightarrow MX_2 \]
\( X_2 \) is any group 9 molecule

**Reaction with Oxygen**
\[ 2M + O_2 \rightarrow 2MO \]

**Reaction with H**
\[ M + H_2 \rightarrow MH_2 \]

**Reaction with N**
\[ 3M + N_2 \rightarrow M_3N_2 \]
High temperatures

**Reaction with Water**
\[ M + 2H_2O \rightarrow M(OH)_2 + H_2 \]

**Reaction with Ions**
\[ M + 2H^+ \rightarrow M^{2+} + H_2 \]
<table>
<thead>
<tr>
<th>Elements</th>
<th>Electron configuration is ns²np¹ (n is the period number)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Boron and aluminum almost always have an oxidation number of +3</td>
</tr>
<tr>
<td></td>
<td>The heavier elements of the group are more likely to keep their s electrons and can have oxidation numbers of +1 or +3</td>
</tr>
</tbody>
</table>
Group 13: The Boron Family

Elements (Boron)

- High ionization energy
- Metalloid
- Forms covalent bonds
- Small atomic radius
- Tends to form compounds that have incomplete octets or are electron deficient
- Mined as borax and kernite \((Na_2B_4O_7 \cdot xH_2O \ \ x = 10 \ or \ 4)\)
- Elemental boron exists in variety of different structures, one of the more common ones being \(B_{12}\)
- Because of the three dimensional network formed by the bonds, boron is very hard and when incorporated in plastics, forms a material that is stiffer than steel yet lighter than aluminum
Group 13: The Boron Family

Elements (Aluminum)

- Most abundant metallic element in the Earth’s crust
- Low density
- Strong metal
- Amphoteric
- Excellent electrical conductor
- Commercial source of aluminum is bauxite \((\text{Al}_2\text{O}_3 \cdot x\text{H}_2\text{O})\) where \(x\) ranges from 1 to 3)
- Bauxite ore is turned into alumina \((\text{Al}_2\text{O}_3)\) using the Bayer process
Hall discovered that if you add Na₃AlF₆ to alumina (Al₂O₃) that the melting temperature decreased from 2050°C to 950°C.

An electrochemical cell can then be used to extract the Al(s)

Cathode Reaction: \[ \text{Al}^{3+} (\text{melt}) + 3e^- \rightarrow \text{Al} (l) \]
Anode Reaction: \[ 2\text{O}^{2-} (\text{melt}) + \text{C} (gr) \rightarrow \text{CO}_2 (g) + 4e^- \]
Overall: \[ 4\text{Al}^{3+} (\text{melt}) + 6\text{O}^{2-} (\text{melt}) + 3\text{C} (gr) \rightarrow 4\text{Al} (l) + 3\text{CO}_2 (g) \]

A current of 1 A must flow for 80 h to produce 27 g of Al about enough for 2 soft drink cans.

If recycled Al is used then the energy consumption drops to less than 5% of the original energy need to extract Al from bauxite.

The energy that we are throwing away when we discard an aluminum can is equivalent to burning the amount of gasoline that would fill half the can.
Group 13: The Boron Family

Compounds of Boron

B(OH)₃ (Boronic Acid)

- Toxic to bacteria, insects, and humans
- Used as a mild antiseptic and pesticide
- Has an incomplete octet so forms bonds by accepting lone pairs of electrons
- Forms an acid anhydride with water

\[
\text{(OH)}_3\text{B} + \text{:OH}_2 \rightarrow \text{(OH)}_3\text{B-OH}_2
\]
Al$_2$O$_3$ \textit{(Aluminum oxide commonly known as alumina)}

- Variety of crystal structures
- Many forms are important ceramic materials
- Some impure forms of alumina are ruby (Cr$^{3+}$), sapphire (Fe$^{3+}$ and Ti$^{4+}$), and topaz (Fe$^{3+}$)
- Amphoteric
Group 13: The Boron Family

Common Reactions

Reaction with Halogens
2M + 3X₂ → 2MX₃  
X₂ is group 9 molecule, Tl gives TlX as well but no TlI₃

Reactions with O
4M + 3O₂ → 2M₂O₃

Reactions with N
2M + N₂ → 2MN

Reactions with the ions
2M + 6H⁺ → 2M³⁺ + 3H₂
2M + 2OH⁻ + 6H₂O → 2M(OH)₄⁻ + 3H₂