II. Groups and Chemical Reactivity

* Group # = # of outermost (valence) e⁻’s

**Group 1A:** Alkali Metals

- Tend to lose 1 e⁻

  \[ \text{eg } K \rightarrow K^+ + e^- \]

  \[ \uparrow \]

  so typical charge (aka oxidation number) is +1
  or oxidation state.

**Group 2A:** Alkaline Earth Elements

- Tend to lose 2 e⁻’s

  \[ \text{eg } \text{Ca} \rightarrow \text{Ca}^{2+} + 2e^- \]

  \[ \therefore \text{ typical ox. state is } +2 \]

**Group 8A:** Noble Gases

- Don’t do squat (a fact we simply accept!)

* Atoms in chemical reactions seek to have as many e⁻’s as (i.e. become isoelectronic with) the nearest noble gas.

**DEMO:** \( \text{Ca(s)} + \text{H}_2\text{O}_\text{(e)} \rightarrow \ldots \)
**DEMO**: \( K(s) + H_2O(e) \rightarrow \ldots \)

**Group 7A: Halogens**

- Tend to gain or share 1\(e^-\)

eg \( F + e^- \rightarrow F^- \)

\(-1\)ox state

eg \( F \) (7 valence \(e^-\)'s)

Lewis (Chap.9) - arrange \(e^-\)'s

in (up to) 4 groups

so \( \cdot \cdot \cdot :F::F::F:\) (1 short of Ne)

\( \Rightarrow F + F \rightarrow F_2 \)

\( \cdot \cdot \cdot :F::F::F:\)

or \( \cdot \cdot \cdot :F-\cdot: \)

ox. state of zero

"shared e- pair" or "covalent bond"

\( F_2 \) eg of a molecule:

1. finite group of atoms
2. fixed composition
3. held together by 1 or more covalent bonds
Group 6A: Chalcogens
- Tend to gain or share 2 e⁻'s

\[ \text{eg } \text{O} + 2\text{e}^- \rightarrow \text{O}^{2-} \]
-2 ox. state

\[ \text{O} + \text{O} \rightarrow \text{O}_2 \]
0 ox. state

(in Chap 9/10, learn why \( \text{O}^{2-} \rightarrow \text{O}_2 \))

Reactivity: \( F > \text{O}, \text{Cl} > S, \text{etc.} \) (why in Chap 8)

III. Compounds and Nomenclature

\[ \uparrow \]
- Contains more than 1 element
- Fixed composition

A. Binary (contains 2 elements)

- Metal + non-metal

\[ \downarrow \text{e}^- \text{ transfer} \]
ionic compound or salt

1. Infinite (in principle) group of atoms
2. Fixed composition
3. Held together by ionic bonds (\( \Theta \leftrightarrow \Theta \))

\[ \text{eg } 2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s) \]

O and all Group 7A elements are \underline{diatomic}
How do we know NaCl consists of ions?

**DEMO:** \( \text{NaCl(s)} \rightarrow \text{Na}^+ \text{(aq)} + \text{Cl}^- \text{(aq)} \)

\( \text{H}_2\text{O} \) solvates the \( \text{Na}^+ \) and \( \text{Cl}^- \), but doesn't react with them chemically.

**NOMENCLATURE**

1. Write cation, then anion
   - element name
   - "ide" suffix
2. Write charge if more than 1 ox state possible
   - eg \( \text{Fe}^{2+} \) iron(II), \( \text{Fe}^{3+} \) iron(III)
3. Overall charge of a formula must be neutral
4. Use a Greek prefix for # of waters of hydration

* non-metal + non-metal
   - \( \downarrow \) sharing of e-'
   - molecule or covalent compound
   - eg nitrogen oxides

**NOMENCLATURE**

1. Write 2nd element as if it were an anion
2. Use Greek prefixes to indicate number of each atom
   - (since it can vary!)
3. Leave off an initial "mono"
B. Compounds with More Than 2 Elements

If 1 element is a metal ⇒ ionic compound
If all elements are non-metallic ...

ionic compound
⇒ cation and/or anion is polyatomic
(more than 1 atom)

eg \( \text{NH}_4\text{CeO}_4 \text{ (s)} \)
\[ \downarrow \text{H}_2\text{O} \]
\( \text{NH}_4^+(\text{aq}) + \text{CeO}_4^{2-} \text{ (aq)} \)

[H\(_2\)O solution can break ionic bonds by ions ...]

\( \text{H} \quad \text{':O':} \)
\[ \text{H-N-H} \quad \text{':O-\text{Ce}-O':} \]
\[ \text{H} \quad \text{':O':} \]

[but H\(_2\)O can't break covalent bonds within polyatomic ions!]

... in terms of chemistry and nomenclature,
... treat polyatomic ions as a unit.