# *The Periodic Table*

## History

Doberiner (1817) – noticed that Ca, Sr, and Ba were similar; suggested elements should be arranged in triads.

Newlands (1866) – observed a repetition of elements in groups of eight, called octaves (as for music).

Mendeleev (1869) – listed elements in order of increasing atomic mass, arranged in columns by similar properties; found that some elements were missing and some heavier elements didn't fit the table.

Modern Periodic Law: the properties of elements recur periodically when the elements are arranged in increasing order by atomic number

## Families (or Groups)

<u>Alkali Metals</u> (Group 1)

| | | |
|---|---|---|
| Li: | $2s^1$ | – very reactive |
| Na: | $3s^1$ | – reactivity increases [down] |
| K: | $4s^1$ | – all have 1 $e^-$ in the highest s orbital |
| Rb: | $5s^1$ | – tend to form 1+ ions by losing $1e^-$ |
| Cs: | $6s^1$ | |
| Fr: | $7s^1$ | |

<u>Alkali Earth Metals</u> (Group 2)

| | | |
|---|---|---|
| Be: | $2s^2$ | – reactive, less so than alkali metals |
| Mg: | $3s^2$ | – reactivity increases [down] |
| Ca: | $4s^2$ | – all have 2 $e^-$ in the highest s orbital |
| Sr: | $5s^2$ | – tend to form 2+ ions by losing $2e^-$ |
| Ba: | $6s^2$ | |
| Ra: | $7s^2$ | |

<u>Transition Metals</u> (Groups 3-12)

| | | |
|---|---|---|
| Cr: | [Ar] $4s^1 3d^5$ | – Fill d orbitals last (except exceptions) |
| Cu: | [Ar] $4s^1 3d^{10}$ | – multiple possible ions (multi-valent) |
| Fe: | [Ar] $4s^2 3d^6$ | – tend to form + ions by losing electrons |
| Mo: | [Kr] $5s^1 4d^5$ | |

Group 13 elements

| | | |
|---|---|---|
| Al: | [Ne] $3s^23p^1$ | – Tend to form +3 ions |
| In: | [Kr] $5s^25p^1$ | – usually metals or semi-metals |

Group 14 elements

| | | |
|---|---|---|
| C: | $1s^22s^22p^2$ | – Tend to form +4 ions (except carbon) |
| Sn: | [Kr] $4s^24p^2$ | – Carbon often forms multi-branched covalent compounds |

Group 15 elements

| | | |
|---|---|---|
| N: | $1s^22s^22p^3$ | – Tend to form -3 ions, gains 3 $e^-$ |
| P: | [Ne] $3s^23p^3$ | – reactivity increases [up] |

Group 16 elements

| | | |
|---|---|---|
| O: | $1s^22s^22p^4$ | – Tend to form -2 ions, gains 2 $e^-$ |
| S: | [Ne] $3s^23p^4$ | – reactivity increases [up] |

Halogens (Group 17)

| | | |
|---|---|---|
| F: | $1s^22s^22p^5$ | – form -1 ions, gain 1 electron |
| Cl: | [Ne] $3s^23p^5$ | – reactivity increases [up] |
| Br: | [Ar] $4s^23d^{10}4p^5$ | – F, C, Br, I are diatomic molecules |
| I: | [Kr] $5s^24d^{10}5p^5$ | |
| At: | [Xe] $6s^24f^{14}5d^{10}6p^5$ | |

Noble Gases (Group 18)

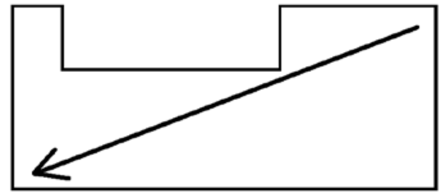
| | | |
|---|---|---|
| He: | $1s^2$ | – Very stable (not-reactive) |
| Ne: | $1s^22s^22p^6$ | – some reactivity of heavier atoms |
| Ar: | [Ne] $3s^23p^6$ | – completely filled outermost orbitals |
| Kr: | [Ar] $4s^23d^{10}4p^6$ | – other elements form ions isoelectronic to the nearest noble gas |
| Xe: | [Kr] $5s^24d^{10}5p^6$ | |
| Rn: | [Xe] $6s^24f^{14}5d^{10}6p^6$ | |

f- block transition metals

| | | |
|---|---|---|
| La – lanthanides | 4f block | La(#57) – Lu(#71) |
| Ac – actinides | 5f block | Ac(#89) – Lr(#103) |

# *Periodic Trends*

## Atomic radius



Recall that the size of an atom is mostly dictated by its electrons. So a large atomic radius (the distance from the centre of the nucleus to the furthest electron(s)) would have one of two characteristics; i) a lot of electrons, ii) a small effective nuclear charge ($z_{eff}$) compared to the number of electron shells (orbitals). This results from inner rows of electrons “shielding” the outer electrons from attraction to the positive nucleus.

Because of this, atomic radii (NON-IONS) decrease from left to right across the periodic table, and increases from top to bottom. Cesium is the largest, and fluorine is the smallest.

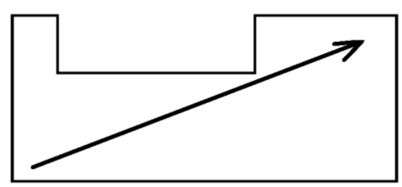
## Ionic radius

If we look at ions that are all isoelectronic (same number of electrons, with same electron configuration) we find that the more positive ions have smaller radii.

**Ex.** $N^{3-}$, $O^{2-}$, $F^{-}$, $Na^{+}$, $Mg^{2+}$, $Al^{3+}$ - are all isoelectronic, and have 10 electrons, but have different nuclear charge

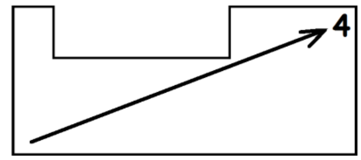
What is the order of the ions radius (from smallest to largest)?

## Ionization energy



The energy required to completely remove an electron from an atom. The factors affecting $E_{Ionization}$ (or IE) are again; i) atomic radius (how close the electrons are to the nucleus) and ii) the effective nuclear charge ($z_{eff}$). A low IE, would mean that it is easy to remove electrons from the atom, a high IE means it’s difficult to remove. The lowest IE’s are in the bottom left hand corner (Cs, Ba, Fr etc..) and the highest IE’s are in the top right hand corner (highest being helium at ~ 2300 kj/mol)

## Electronegativity




Recall, electronegativity is the ability of bonded elements/atoms to attract electrons. If we anthropomorphize elements, then it's their "desire" or "greed" to get more electrons. It arises from two factors; i) nuclear charge relative to atomic radius, and ii) relative stability in ionic form. If an atom is smaller, has a larger $z_{eff}$ to electron shells ratio, and only requires a small number of electrons to become more stable (think stable octet) then it will have a high electronegativity

Electronegativity increases as we go across the periodic table (left to right) and as we go up the periodic table, bottom to top. Fluorine has the largest electronegativity with a value of 4.0. This trend does NOT include the noble gases. We'll be dealing with electronegativity a lot in the coming sections.

## Metallic character

A measure of "metallic-ness" is a measure of how easily an atom loses its electrons. The most metallic element is technically francium (Fr) but it's not a naturally occurring element, so cesium (Cs) would be the most metallic element on earth.