NEW JERSEY CENTER
FOR TEACHING & LEARNING

Progressive Science Initiative®

We, at the New Jersey Education Association, are proud founders and supporters of NJCTL, an independent non-profit organization with the mission of empowering teachers to lead school improvement for the benefit of all students.

njea

NEW JERSEY CENTER
FOR TEACHING & LEARNING

Chemistry

The Periodic Table

2015-11-16

www.njctl.org

Table of Contents: The Periodic Table

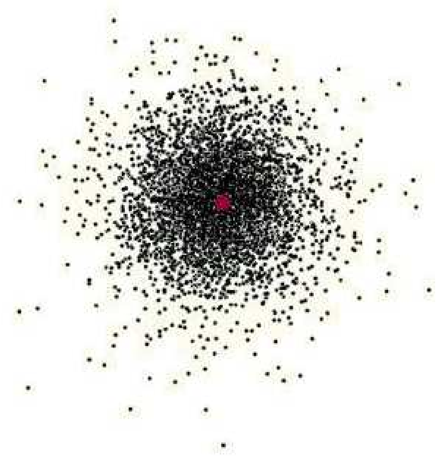
Click on the topic to go to that section

The Periodic Table

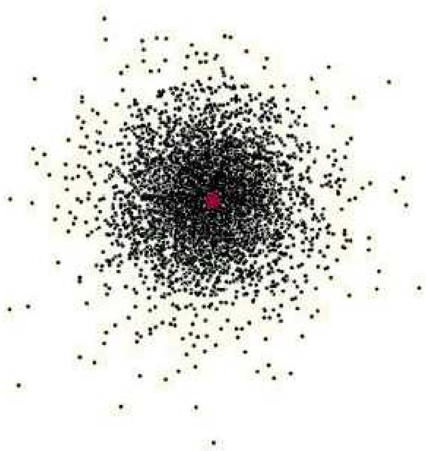
Return to Table of Contents

Identifying Properties of Atoms

Now that we know where (or approximately where) to find the parts of atoms, we can start to understand how these factors all come together to affect how we view the elements.



Identifying Properties of Atoms



We can look at them as individual yet interacting chemicals, and we are able to group them based, not only on the properties they present when in isolation, but also the properties they reveal when exposed to other elements or compounds.

"Periodic" Table of Elements

The Periodic Table of Elements contains physical and chemical information about every element that matter can be made of in the Universe.

The Pillars of Creation, part of the Eagle Nebula shown to the right, *is a cloud of interstellar gases 7,000 light years from Earth made up of the same gaseous elements found on the Periodic Table.

Courtesy of Hubble Telescope

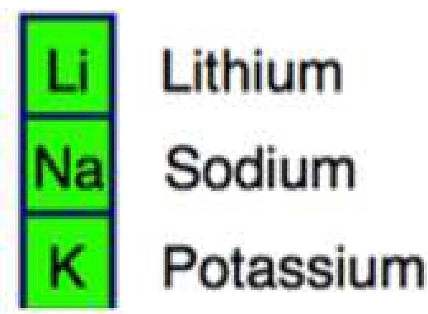
*NASA recently captured this image; however, the Pillars of Creation no longer exists. The Eagle Nebula was destroyed by a Supernova around 6000 years ago, but from our viewpoint, it will be visible for another 1000 years.

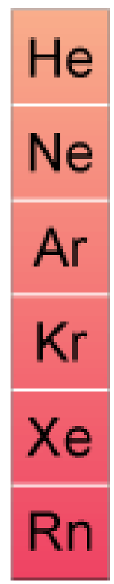
"Periodic" Table of Elements

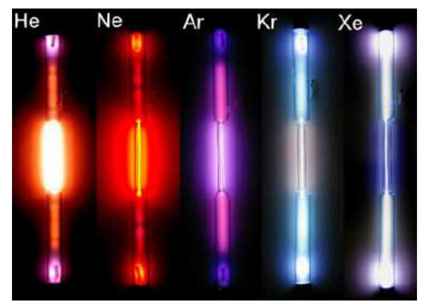
Why is one of the most useful tools ever created by humans called the "Periodic Table"?

When scientists were organizing the known elements, they noticed that certain patterns of chemical and physical behavior kept repeating themselves.

These elements are all shiny metals and react violently in water.







These elements are all very stable gases.

"Periodic" Table of Elements

These patterns were so predictable that Dmitri Mendeleev, the scientist who formulated the Periodic Law, was actually able to predict the existence of elements #31 and #32 and their approximate masses *before they were discovered* based on the existing patterns of known elements.

Gallium, $_{31}Ga$

Germanium, $_{32}Ge$

Mendeleev's work preceded the discovery of subatomic particles.

"Periodic" Table of Elements

THE PERIODIC LAW

Mendelejeff's First Periodic Table (March, 1869)

						Ti	50	Zr	90	?	100
						V	51	Nb	94	Ta	182
						Cr	52	Mo	96	W	186
						Mn	55	Rh	104·4	Pt	197·4
						Fe	56	Ru	104·4	Ir	198
						Ni=Co	59	Pd	106·6	Os	199
H	1					Cu	63·4	Ag	108	Hg	200
		Be	9·4	Mg	24	Zn	65·2	Cd	112		
		B	11	Al	27	?	68	U	116	Au	197?
		C	12	Si	28	?	70	Sn	118		
		N	14	P	31	As	75	Sb	122	Bi	210?
		O	16	S	32	Se	79·4	Te	128?		
		F	19	Cl	35·5	Br	80	I	127		
Li	7	Na	23	K	39	Rb	85·4	Cs	133	Tl	204
				Ca	40	Sr	87·6	Ba	137	Pb	207
				?	45	Ce	92				
				Er?	56	La	94				
				Yt?	60	Di	95				
				In	75·6?	Th	118?				

History of the Periodic Table

Mendeleev argued that elemental properties are periodic functions of their atomic weights.

We now know that element properties are periodic functions of their atomic number.

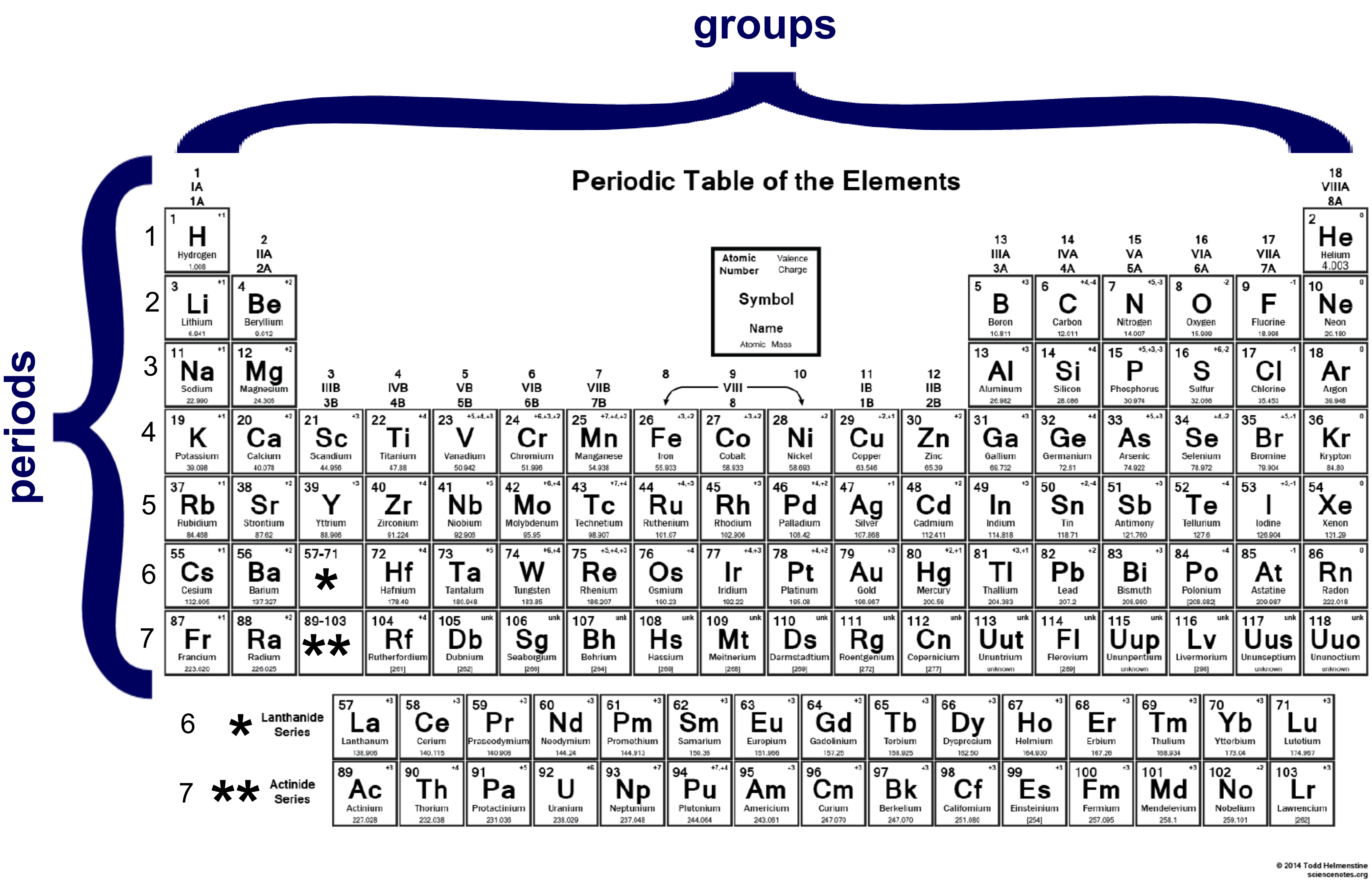
Atoms are listed on the periodic table in rows, based on number of protons.

Periodic Table

The periodic table is made of rows and columns:

Rows in the periodic table are called Periods.
Columns in the periodic table are called Groups.

Groups are sometimes referred to as *Families*, but "groups" is more traditional.


groups
periods
Periodic Table of the Elements
Atomic Number
Valence Charge
Symbol
Name
Atomic Mass
Lanthanide Series
Actinide Series
© 2014 Todd Helmenstine
sciencenotes.org

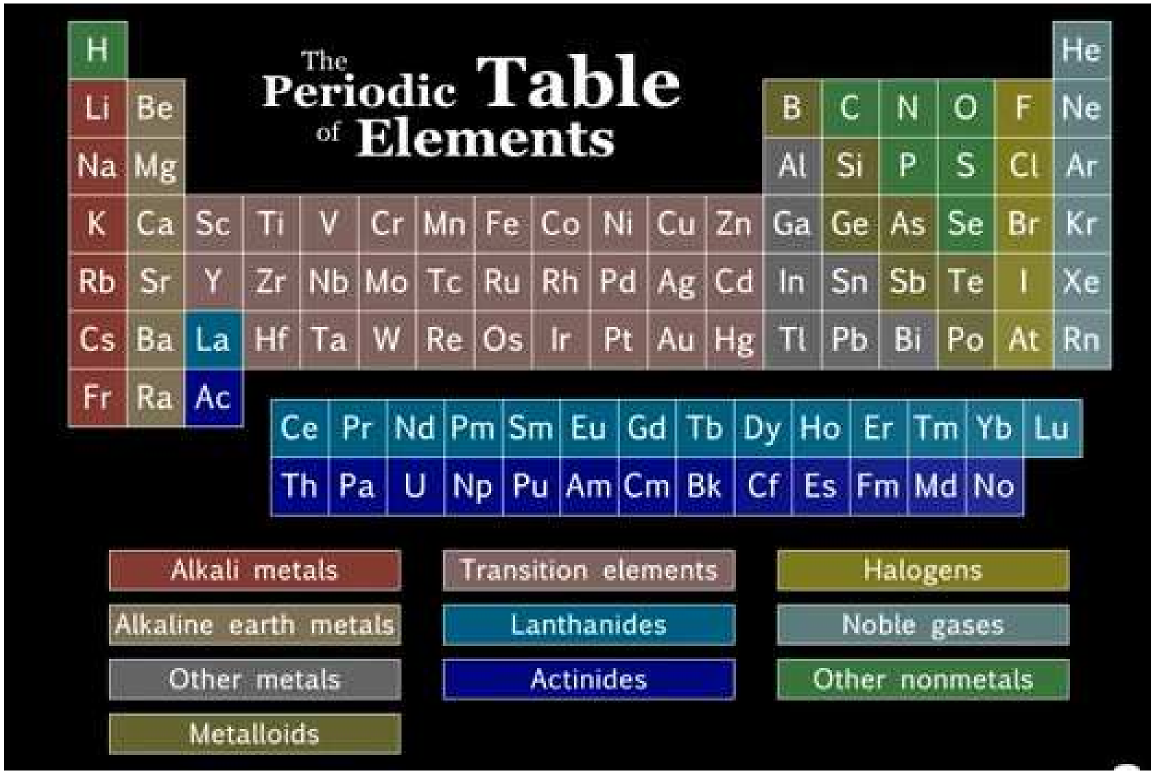
1 The elements in the Periodic Table are arranged from left to right in order of increasing ___.

- A mass
- B number of neutrons
- C number of protons
- D number of protons and electrons

2 What is the atomic number for the element in period 3, group 16?

3 What is the atomic number for the element in period 5, group 3?

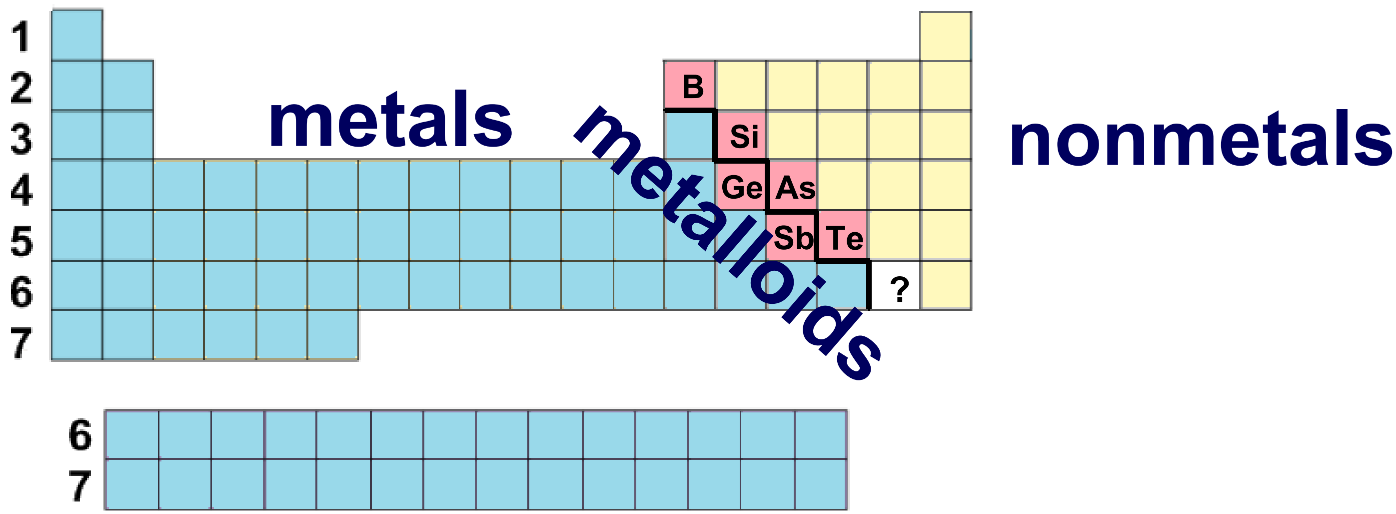
Groups of Elements



Enjoy Tom Lehrer's
Famous Element Song!

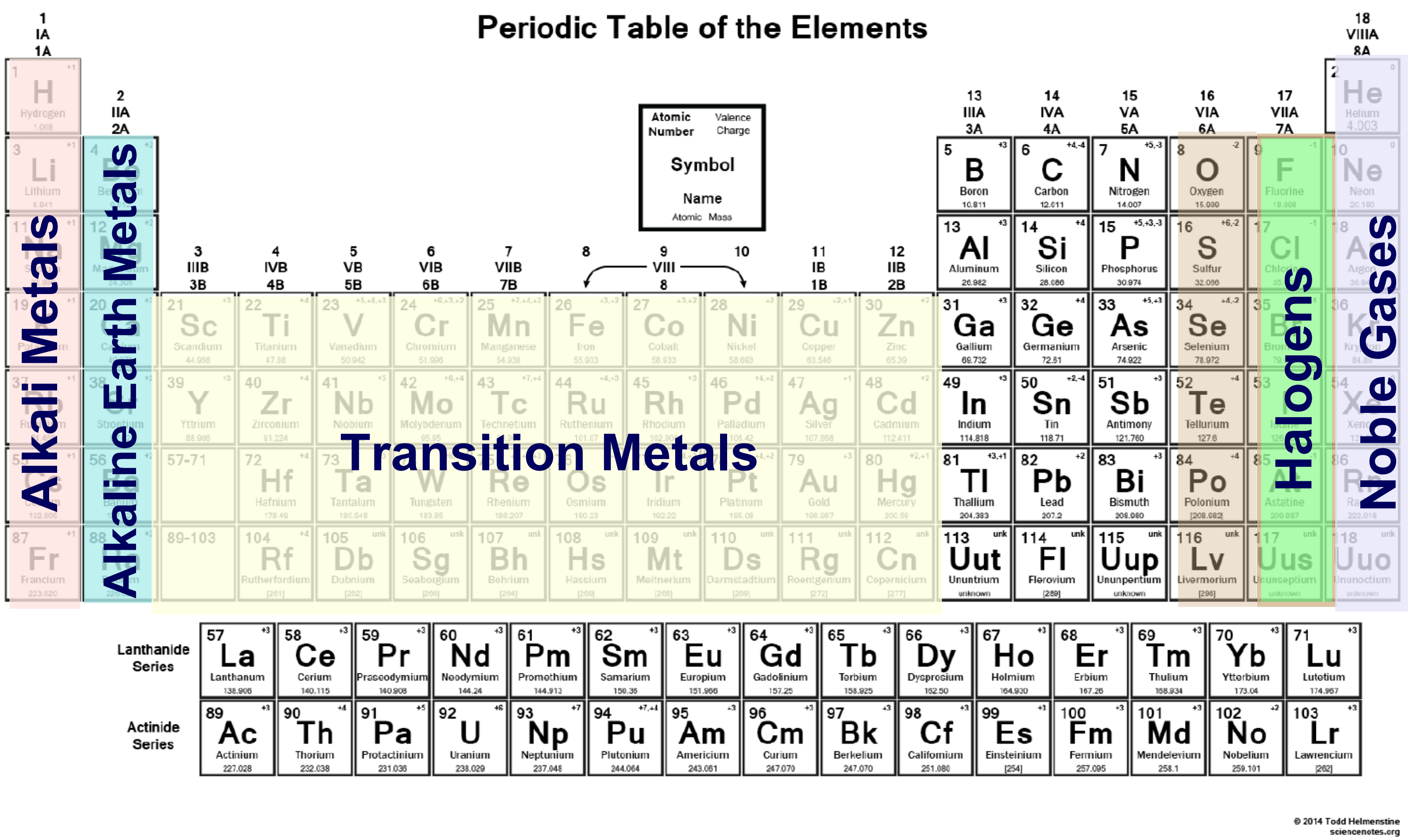
Metals, Nonmetals, and Metalloids

The periodic table can be divided into metals (blue) and nonmetals (yellow) . A few elements retain some ofthe properties of metals and nonmetals, they are called metalloids (pink).

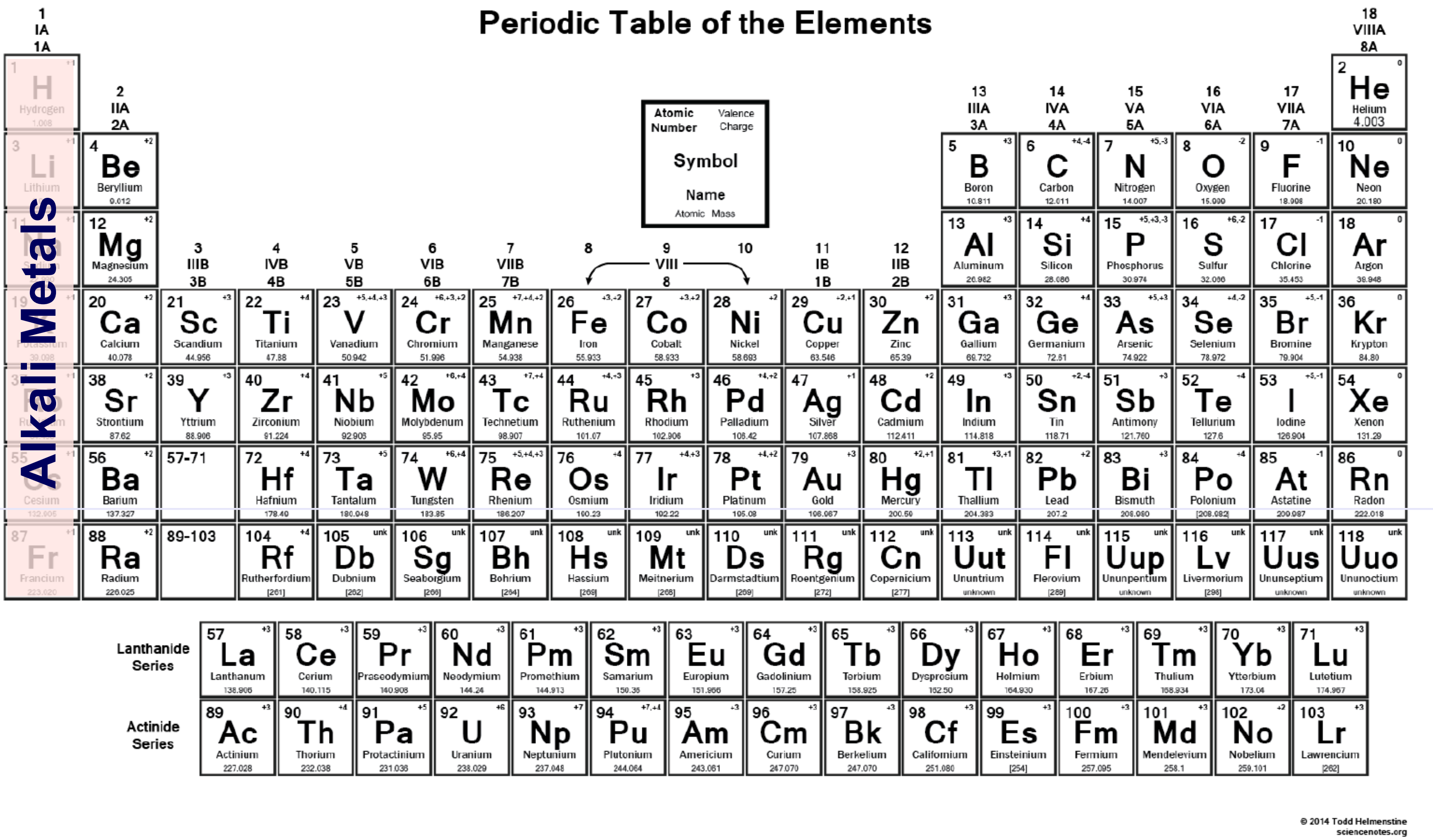


Special Groups

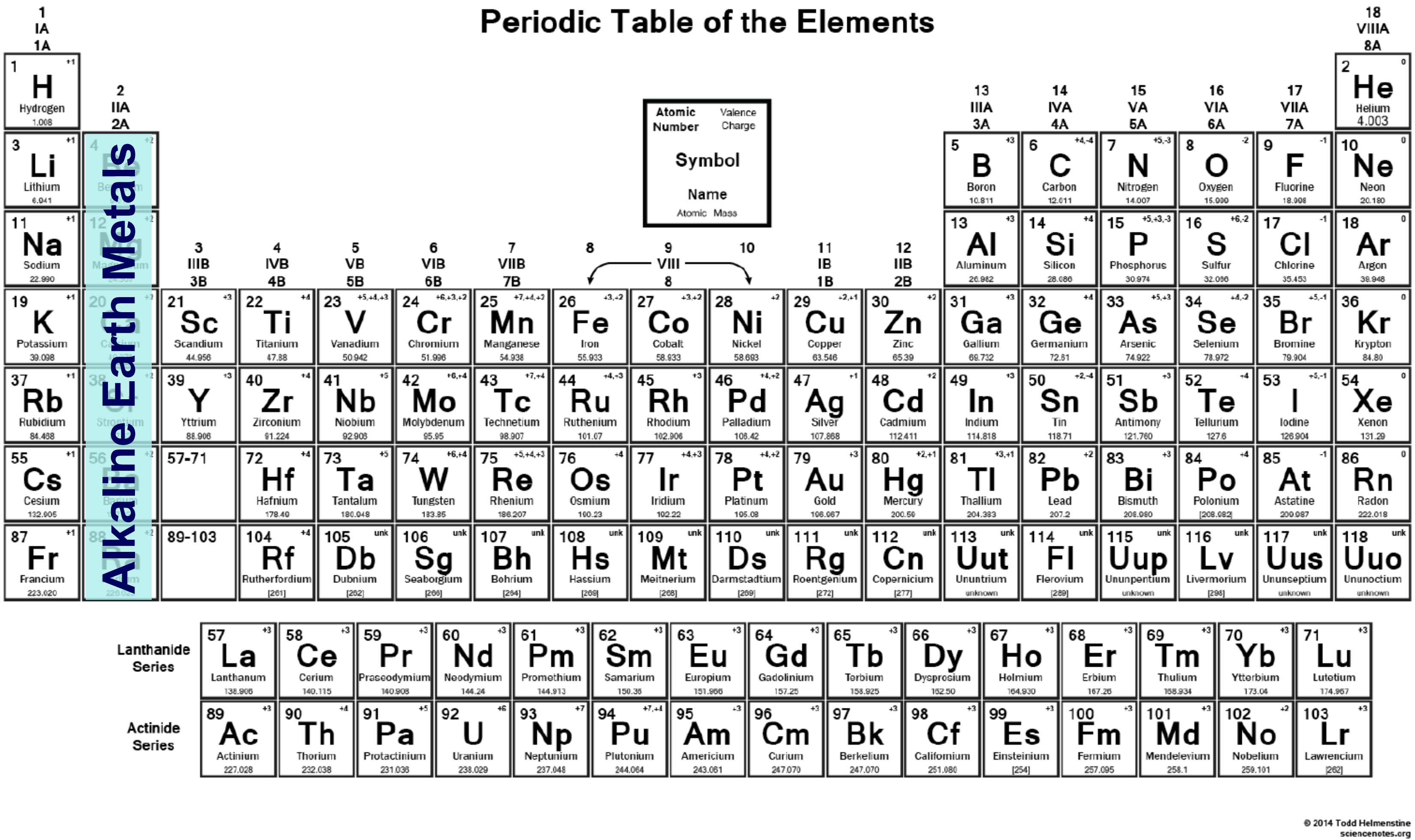
Some groups have distinctive properties and are given special names.



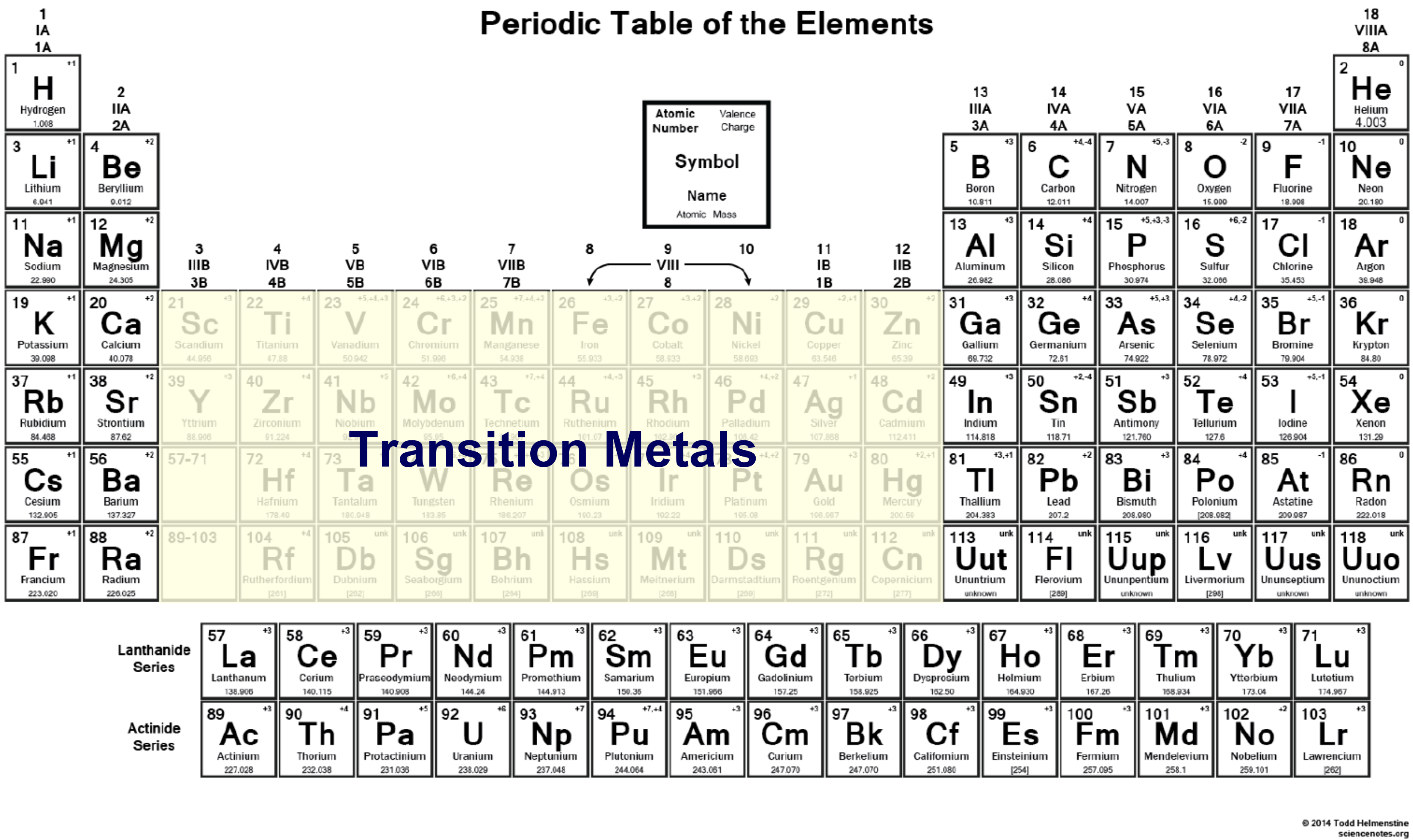
Group 1 *Alkali Metals* (very reactive metals)



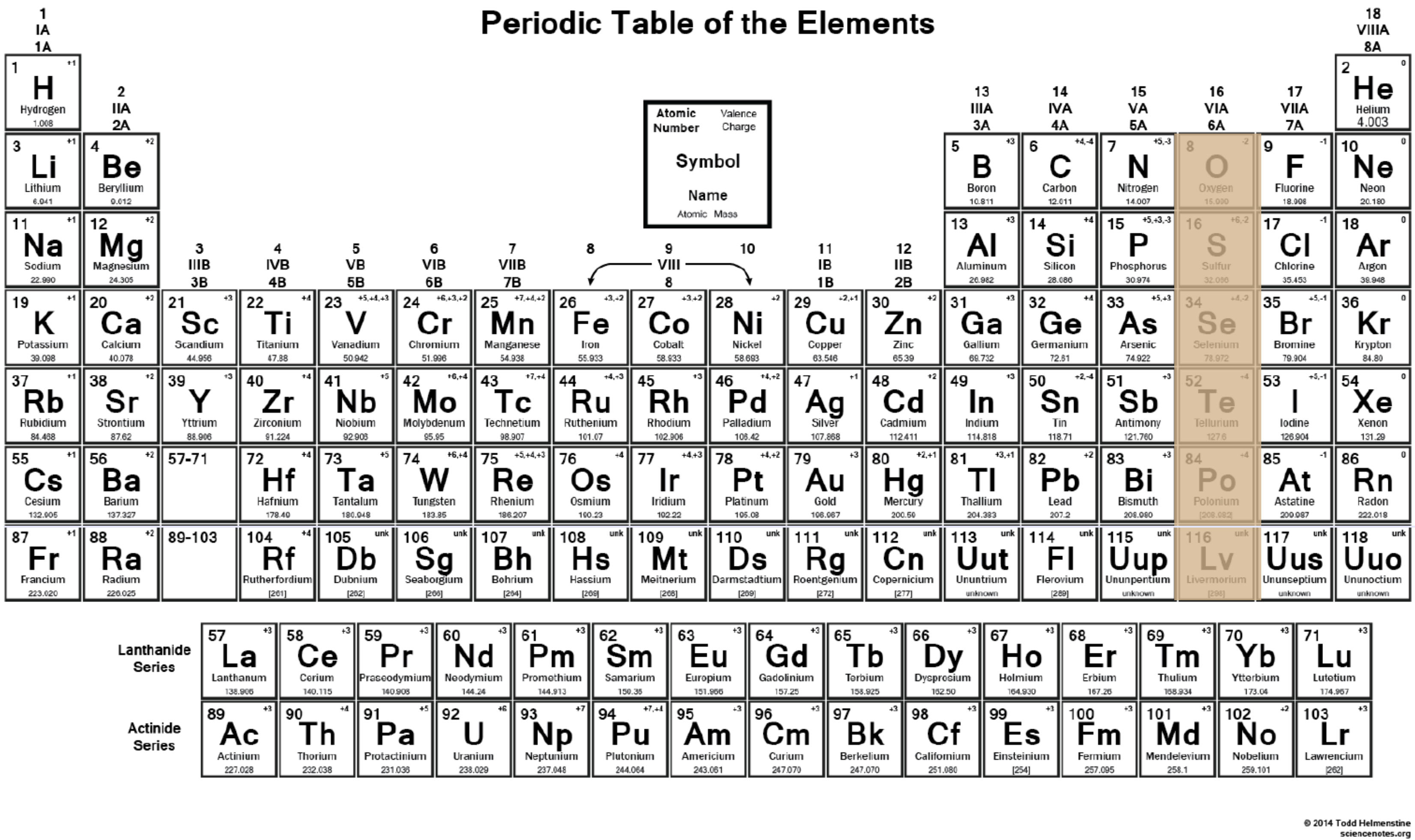
Group 2 *Alkaline Earth Metals* (reactive metals)



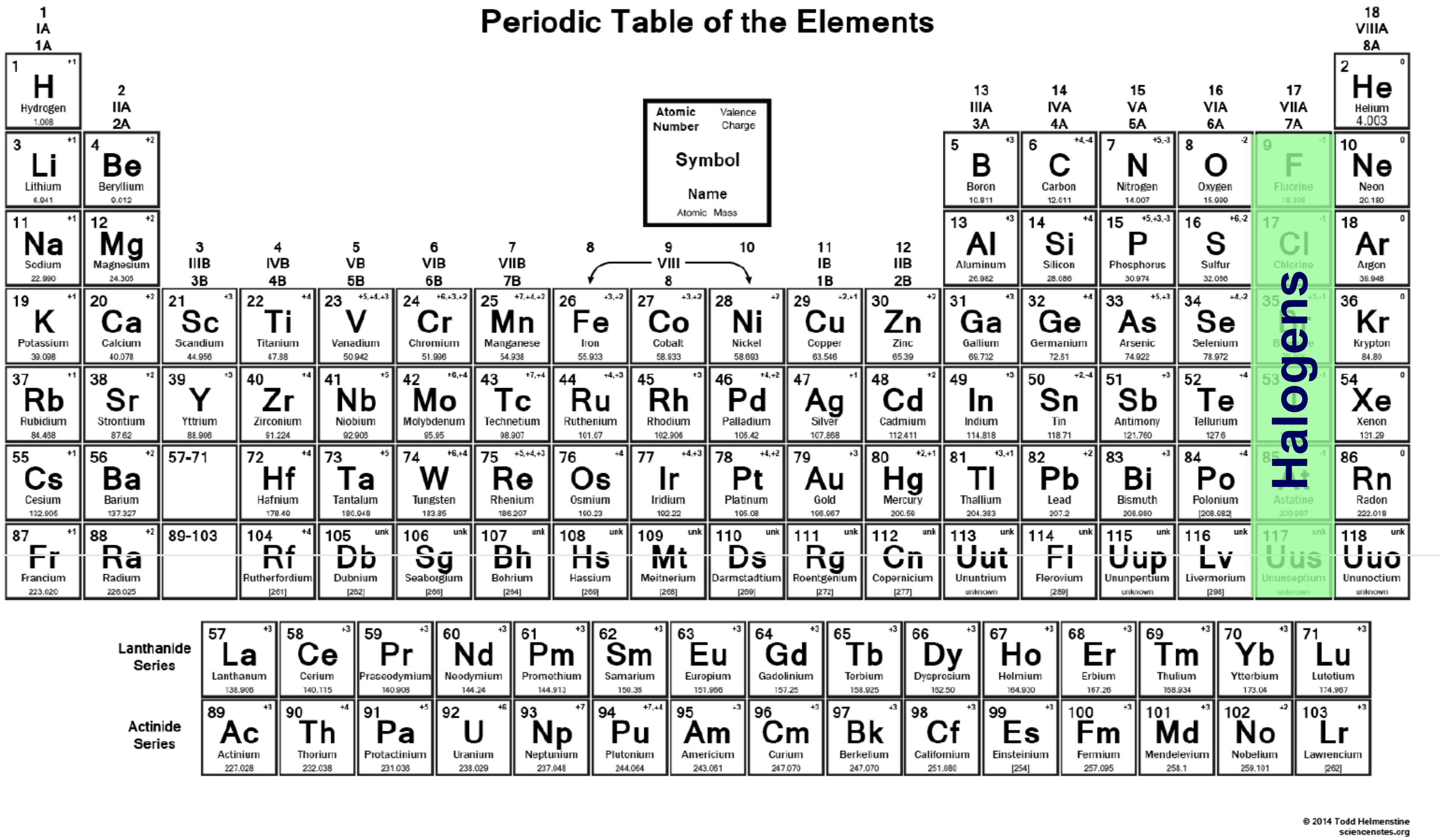
Groups 3 - 12 *Transition Metals* (low reactivity, typical metals)



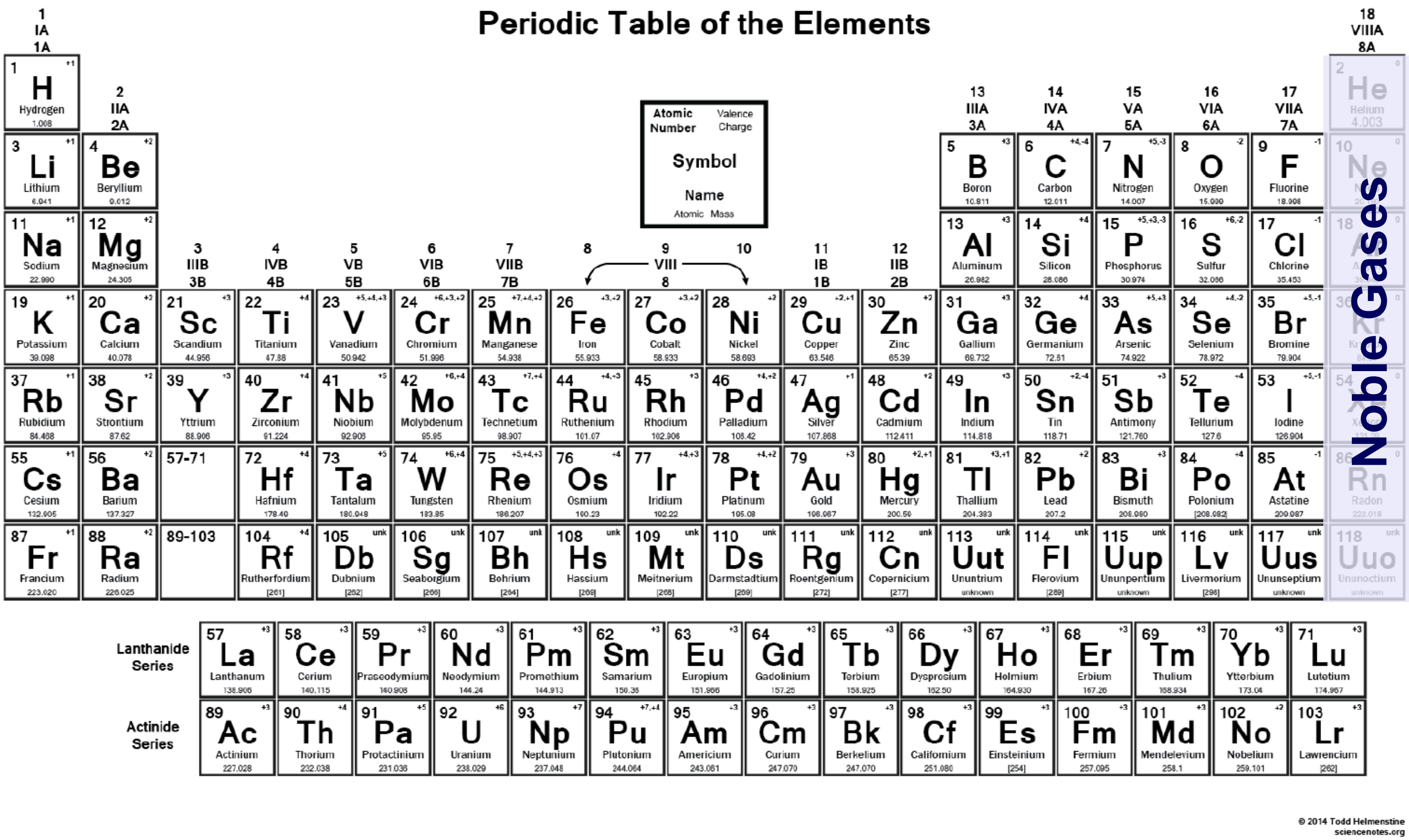
Group 16 Oxygen Family (elements of fire)



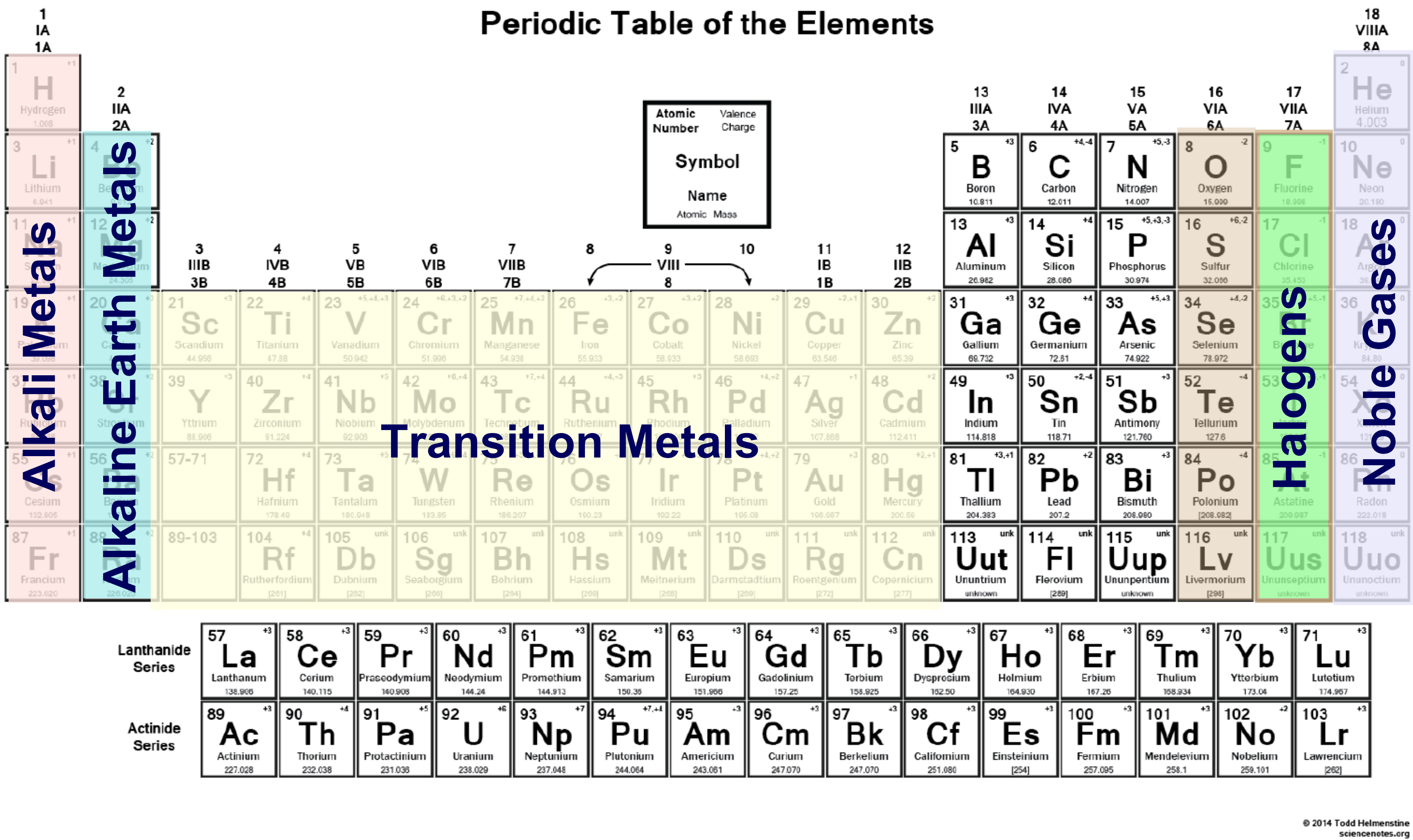
Group 17 *Halogens* (highly reactive, nonmetals)



Group 18 *Noble Gases* (nearly inert)



Major Groups of the Periodic Table



4 To which group on the periodic table does Iodine belong?

- A Noble Gases
- B Alkali Metals
- C Transition Metals
- D Halogens

5 To which group on the periodic table does Neon belong?

- A Alkali Metals
- B Transition Metals
- C Noble Gases
- D Alkaline Earth Metals

6 To which group on the periodic table does Fluorine belong?

- A Alkali Metals
- B Transition Metals
- C Noble Gases
- D Halogens

7 To which group on the periodic table does Iron belong?

- A Alkali Metals
- B Transition Metals
- C Halogens
- D Alkaline Earth Metals

8 To which group on the periodic table does Beryllium belong?

- A Alkali Metals
- B Transition Metals
- C Halogens
- D Alkaline Earth Metals

9 Two elements are studied. One with atomic number X and one with atomic number X+1. It is known that element X is a Noble Gas. Which group on the periodic table is X+1 in?

- A Transition Metals
- B Halogens
- C Alkali Metals
- D There is no way to tell

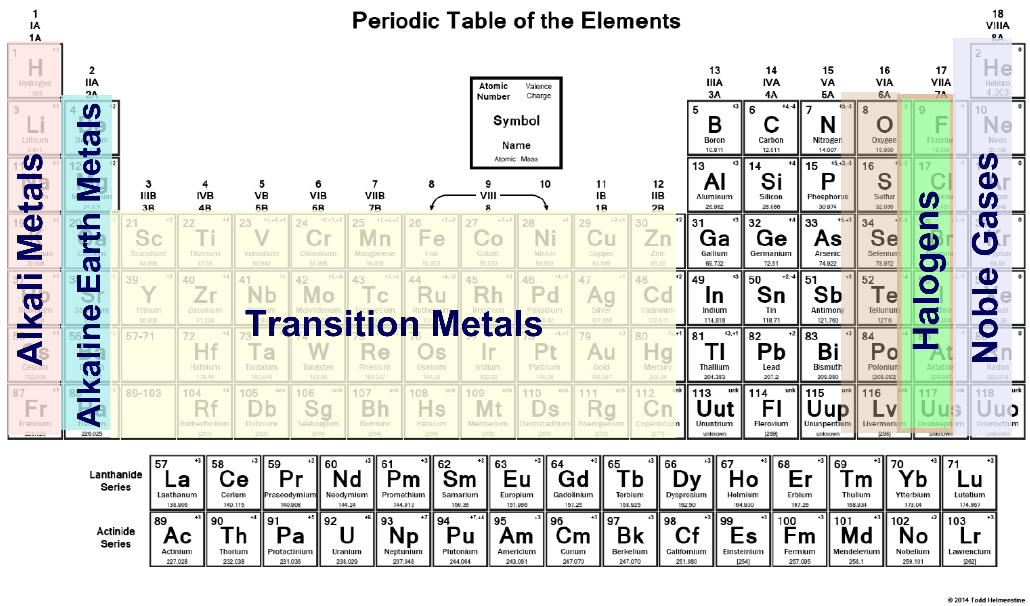
Periodic Table & Electron Configurations

Return to Table of Contents

Periodic Table & Electron Configuration

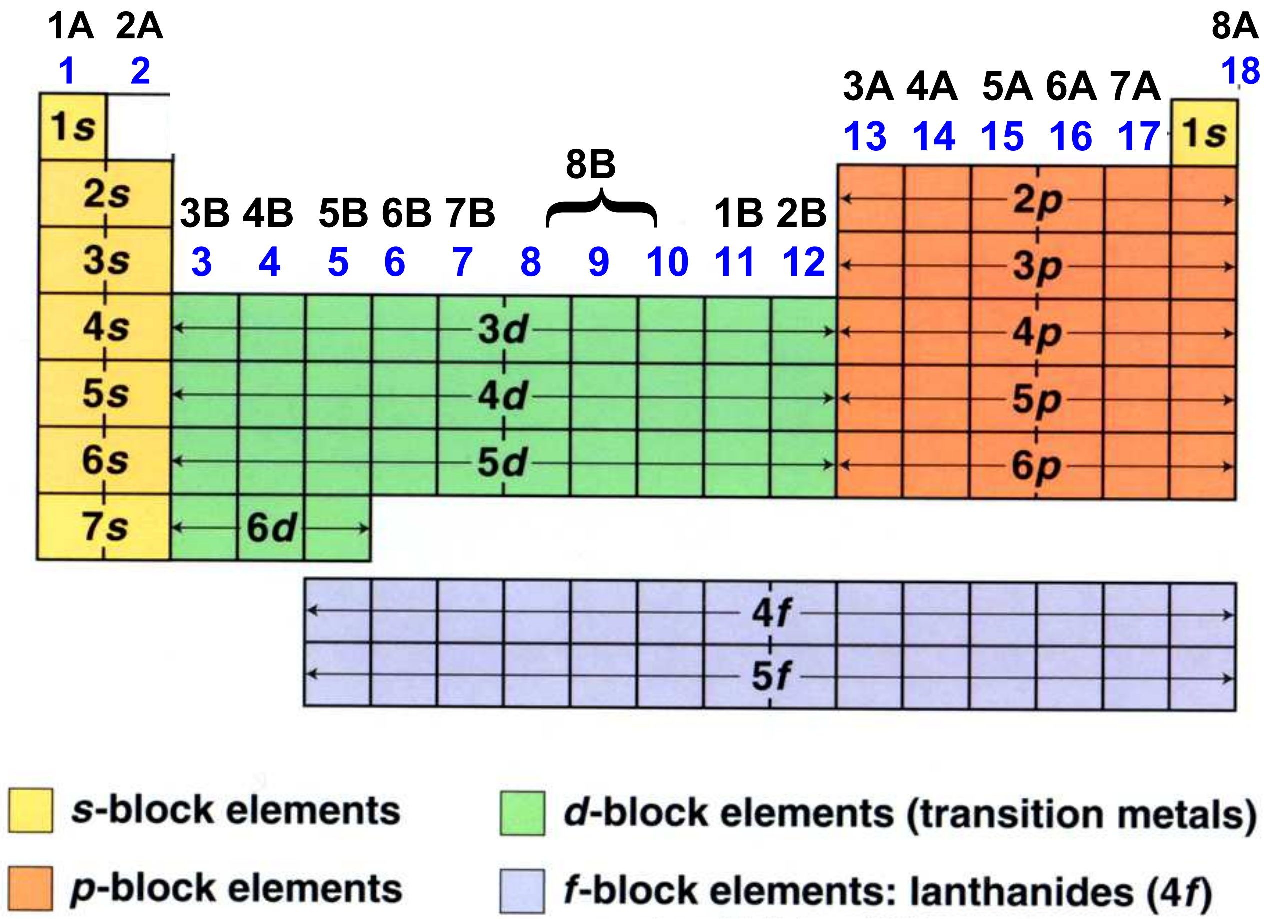
The elements are arranged by groups with similar reactivity.

How an element reacts depends on **how its electrons are arranged**. . .



. . . we now know that elements in the same groups, with the same chemical properties have very similar electron configurations.

Periodic Table & Electron Configuration



There are two methods for labeling the groups, the older method shown in black on the top and the newer method shown in blue on the bottom.

Periodic Table & Electron Configuration

Click here to view an Interactive Periodic Table that shows orbitals for each Element

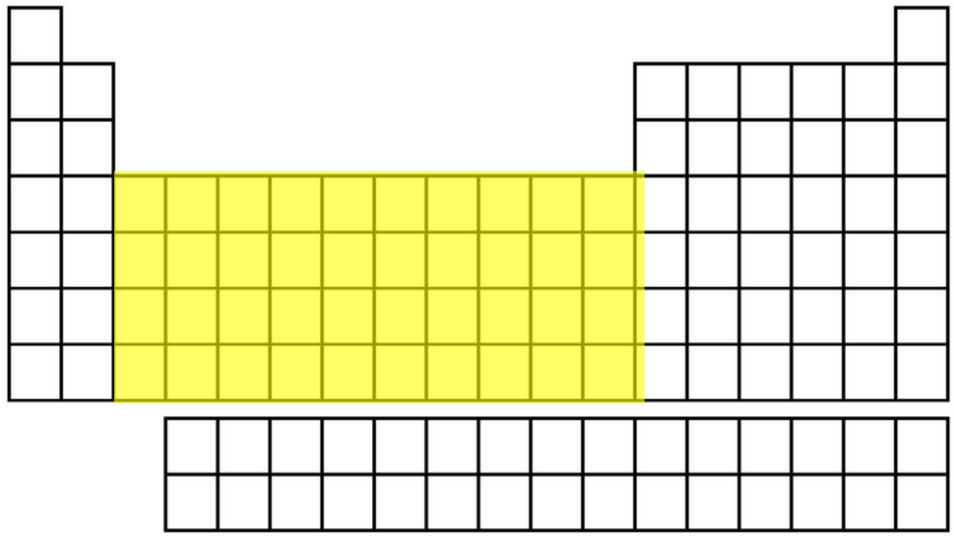
Click here for an electron orbital game.

Group Names

Group Name	Group #	Electron Configuration	Characteristic
Alkali Metals	1	s^1 ending	Very reactive
Alkaline Earth Metals	2	s^2 ending	Reactive
Transition Metals	3-12 (d block)	ns^2, (n-1)d ending	Somewhat reactive, typical metals
Inner Transition Metals	f block	ns^2, (n-2)f ending	Somewhat reactive, radioactive
Halogens	17	s^2p^5 ending	Highly reactive
Noble Gases	18	s^2p^6 ending	Nonreactive

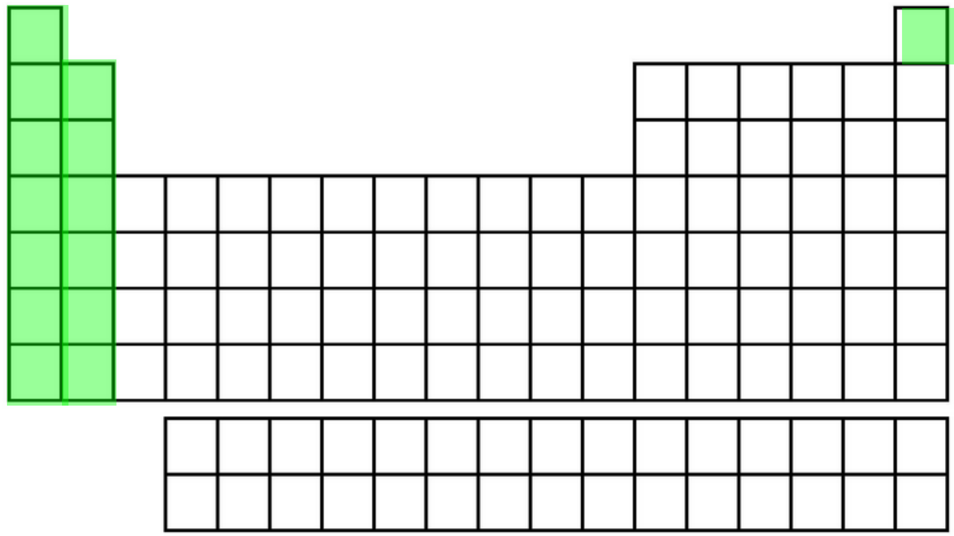
10 The highlighted elements below are in the ___.

- A s block
- B d block
- C p block
- D f block



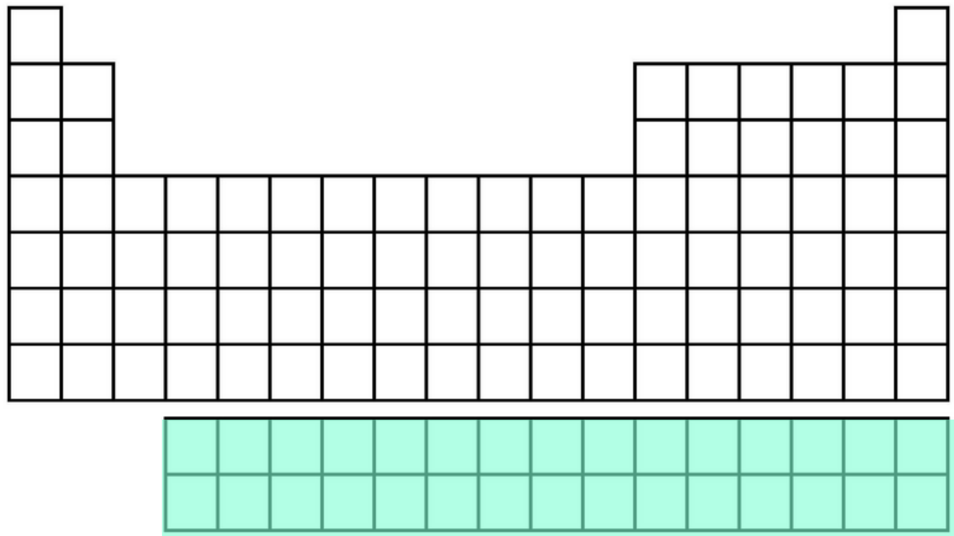
11 The highlighted elements below are in the ___.

- A s block
- B d block
- C p block
- D f block



12 The highlighted elements below are in the ___.

- A s block
- B d block
- C p block
- D f block



13 Elements in each group on the Periodic Table have similar ___.

- A mass
- B number of neutrons
- C number of protons and electrons
- D electron configurations

14 The electron configuration ending ns^2p^6 belongs in which group of the periodic table?

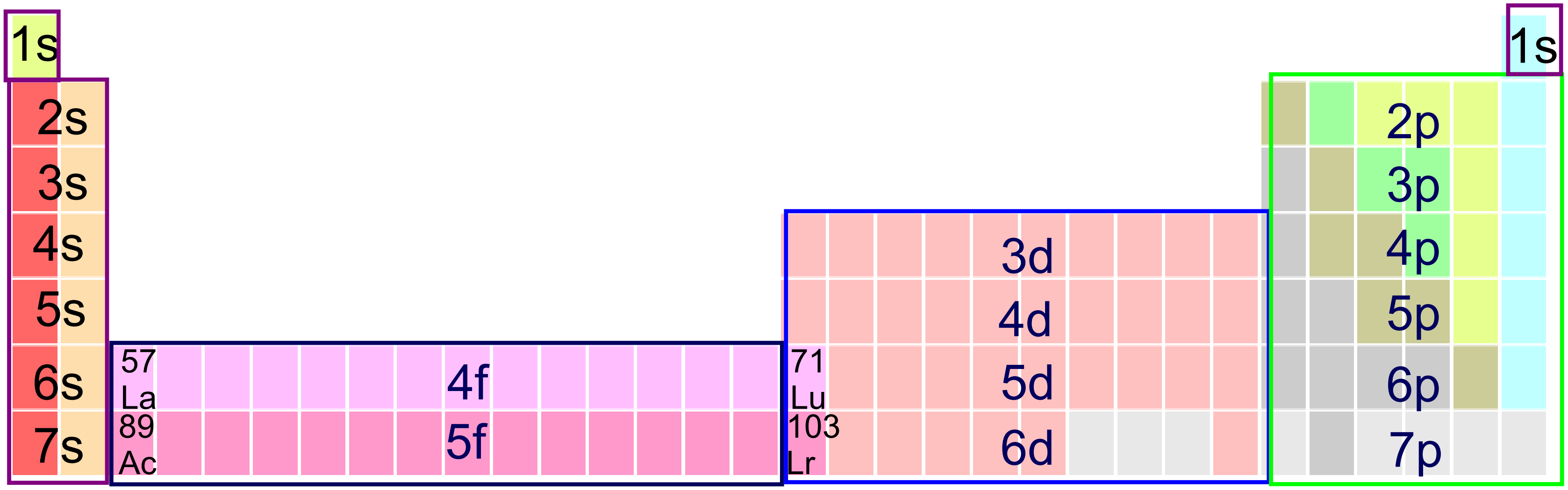
- A Alkali Metals
- B Alkaline Earth Metals
- C Halogens
- D Noble Gases

15 An unknown element has an electron configuration ending in s^2. It is most likely in which group?

- A Alkaline Earth Metals
- B Halogens
- C Alkali Metals
- D Transition Metals

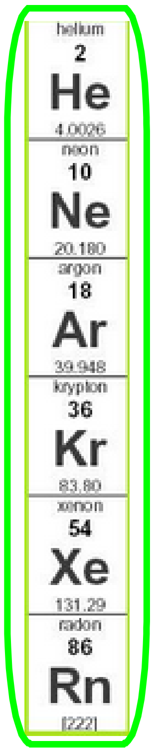
Periodic Table with f block in Place

Here is the Periodic Table with the f block in sequence.
Why isn't this the more commonly used version of the table?



Shorthand Configurations

Noble Gas elements are used to write shortened electron configurations.



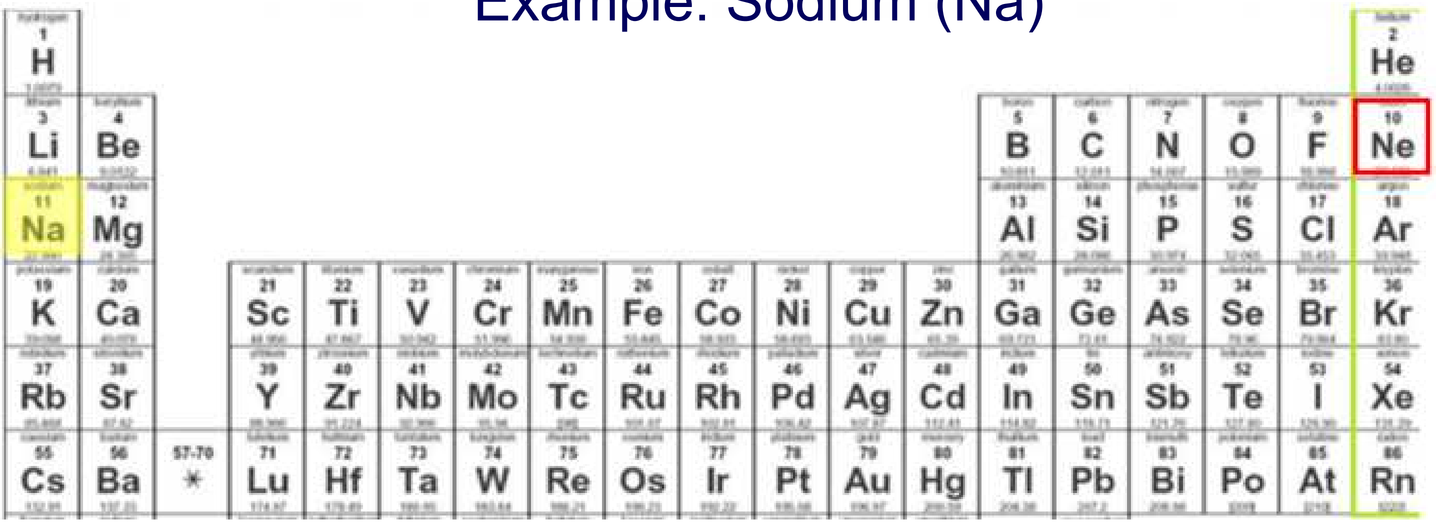
To write a Shorthand Configuration for an element:

(1) Write the Symbol of the Noble Gas element from the row before it in brackets [ ].

(2) Add the remaining electrons by starting at the s orbital of the row that the element is in until the configuration is complete.

Shorthand Configurations

Example: Sodium (Na)



Electron Configuration: $1s^2 2s^2 2p^6 3s^1$

Neon's electron configuration

Shorthand Configuration: $[Ne]\ 3s^1$

Fill in Shorthand Configurations

Element	Shorthand Configuration
Ca	Slide for Answers
V	
F	
Ag	
I	
Xe	
Fe	
Sg	

16 What would be the expected "shorthand" electron configuration for Sulfur (S)?

- A $[He]3s^2 3p^4$
- B $[Ar]3s^2 4p^4$
- C $[Ne]3s^2 3p^3$
- D $[Ne]3s^2 3p^4$

17 What would be the expected "shorthand" electron configuration for vanadium (V) ?

- A $[He]4s^2 3d^1$
- B $[Ar]4s^2 3d^{10} 4p^1$
- C $[Ar]4s^2 3d^3$
- D $[Kr]4s^2 3d^1$

18 Which of the following represents an electron configuration of a halogen?

- A $[He]2s^1$
- B $[Ne]3s^2 3p^5$
- C $[Ar]4s^2 3d^2$
- D $[Kr]5s^2 4d^{10} 5p^4$

19 The electron configuration $[Ar]4s^2 3d^5$ belongs in which group of the periodic table?

- A Alkali Metals
- B Alkaline Earth Metals
- C Transition Metals
- D Halogens

20 Which of the following represents an electron configuration of an alkaline earth metal?

- A $[He]2s^1$
- B $[Ne]3s^2 3p^6$
- C $[Ar]4s^2 3d^2$
- D $[Xe]6s^2$

21 The element iridium is found in a higher abundance in meteorites than in Earth's crust. One specific layer of Earth associated with the end of the Cretaceous Period has an abnormal abundance of iridium, which led scientists to hypothesize that the impact of a massive extraterrestrial object caused the extinction of the dinosaurs 66 million years ago. Using the Periodic Table, choose the correct electron configuration for iridium.

- A $[Xe]6s^2 5d^7$
- B $[Xe]6s^2 4f^{14} 5d^7$
- C $[Xe]6s^2 5f^{14} 5d^7$
- D $[Xe]6s^2 5f^{14} 6d^7$

22 The element tin has been known for a long and was even mentioned in the Old Testament of the Bible. During the Bronze Age, humans mixed tin and copper to make a malleable alloy called bronze. Tin's symbol is Sn, which comes from the Latin word "stannum." Which of the following is tin's correct electron configuration?

- A $[Xe]5s^2 5d^{10} 5p^2$
- B $[Kr]5s^2 4f^{14} 5d^{10} 5p^2$
- C $[Kr]5s^2 4d^{10} 5p^2$
- D $[Kr]5s^2 5d^{10} 5p^2$

23 Chemical elements with atomic numbers greater than 92 are called transuranic elements. They are all unstable and decay into other elements. All were discovered in the laboratory by using nuclear reactors or particle accelerators, although neptunium and plutonium were also discovered later in nature. Neptunium, number 93, and plutonium, number 94, were synthesized by bombarding uranium-238 with deuterons (a proton and neutron). What is plutonium's electron configuration?

- A $[Rn]7s^2 5d^{10} 6f^2$
- B $[Rn]7s^2 5f^{14} 6d^{10} 6p^2$
- C $[Rn]7s^2 6d^{10} 5f^6$
- D $[Rn]7s^2 5f^6$

Stability

When the elements were studied, scientists noticed that, when put in the same situation, some elements reacted while others did not.

The elements that did not react were labeled "stable" because they did not change easily. When these stableelements were grouped together, periodically, they formed a pattern.

Today we recognize that this difference in stability is due to electron configurations.

Based on your knowledge and the electron configurations of argon and zinc, can you predict which electron is more stable?

Argon	Zinc
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

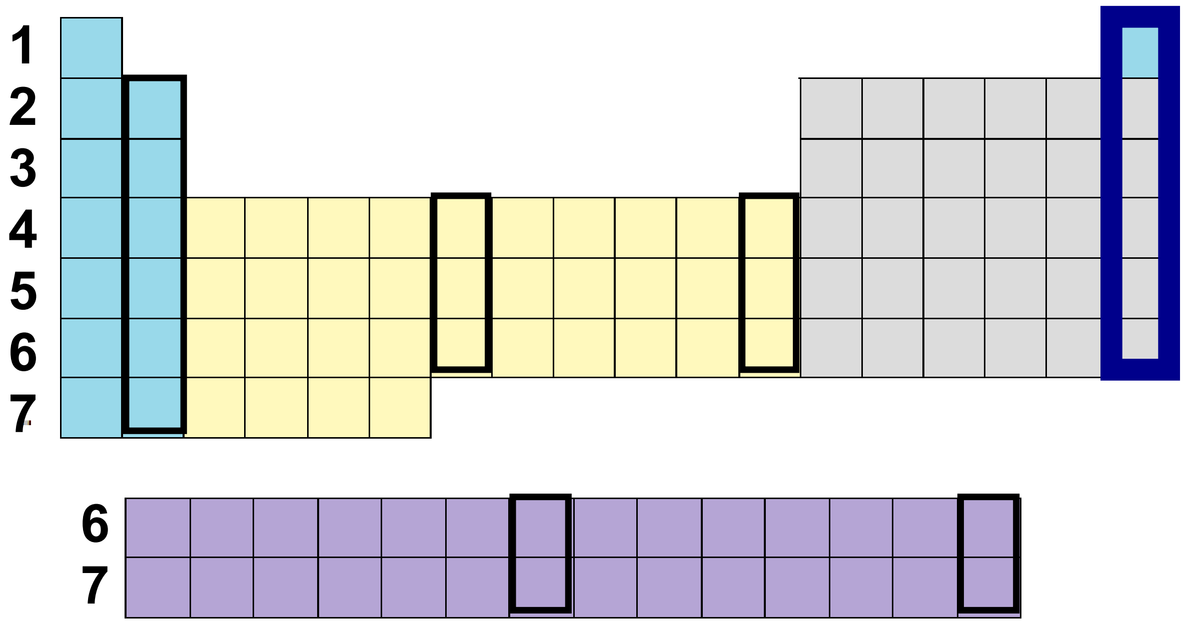
Stability

Elements of varying stability fall into one of 3 categories. The most stable atoms have completely full energy levels.

~Full Energy Level

~Full Sublevel (s, p, d, f)

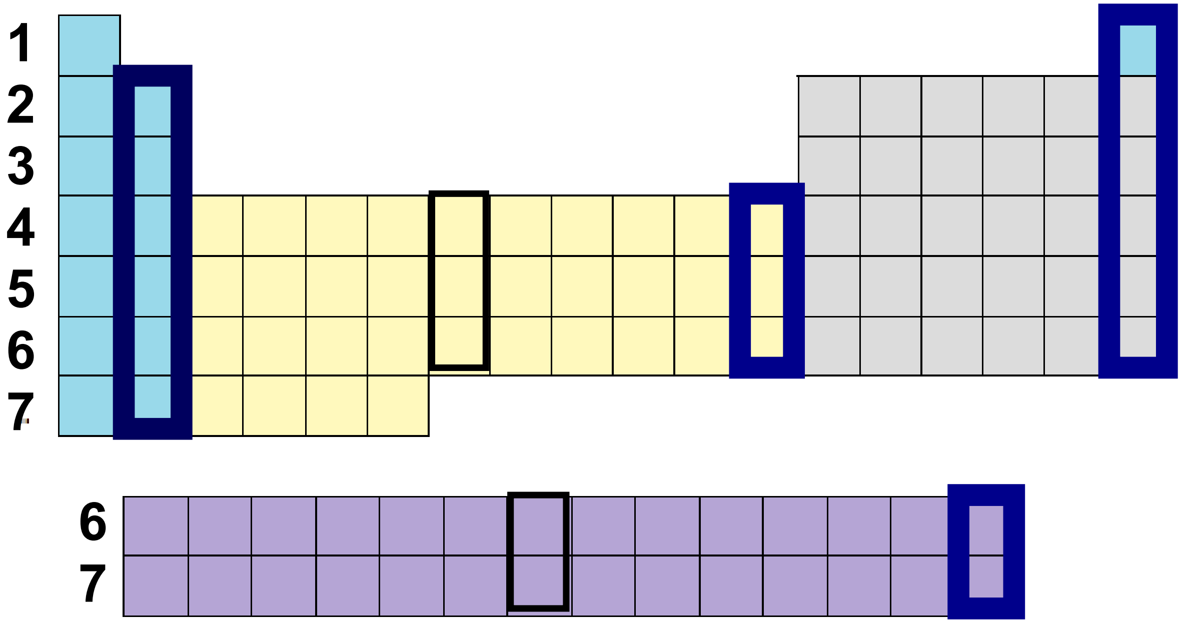
~Half Full Sublevel (d^5, f^7)



Stability

Next in order of stability are elements with full sublevels.

~Full Energy Level
~Full Sublevel (s, p, d, f)
~Half Full Sublevel (d^5, f^7)



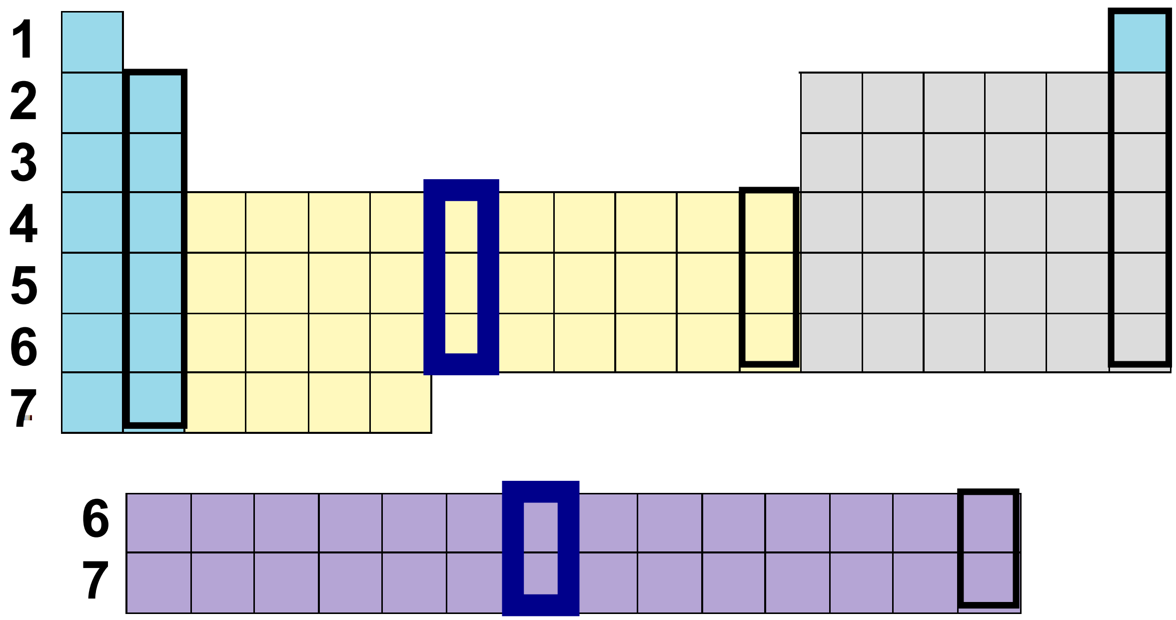
Stability

Finally, the elements with half full sublevels are also stable, but not as stable as elements with fully energy levels or sublevels.

~Full Energy Level

~Full Sublevel (s, p, d, f)

~Half Full Sublevel (d^5, f^7)



24 The elements in the periodic table that have completely filled shells or subshells are referred to as:

- A noble gases.
- B halogens.
- C alkali metals.
- D transition elements.

25 Alkaline earth metals are more stable than alkali metals because...

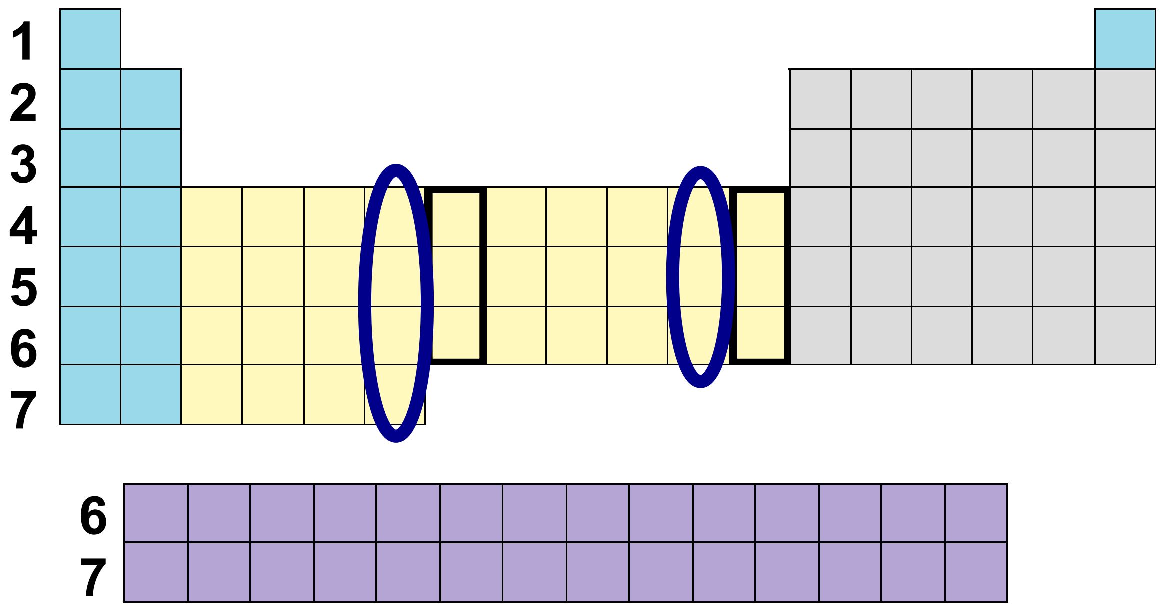
- A they have a full shell.
- B they have a full subshell.
- C they have a half-full subshell.
- D they contain no p orbitals.

26 The elements in the periodic table which lack one electron from a filled shell are referred to as ___.

- A noble gases
- B halogens
- C alkali metals
- D transition elements

Electron Configuration Exceptions

There are basic exceptions in electron configurations in the d- and f-sublevels These fall in the circled areas on the table below.

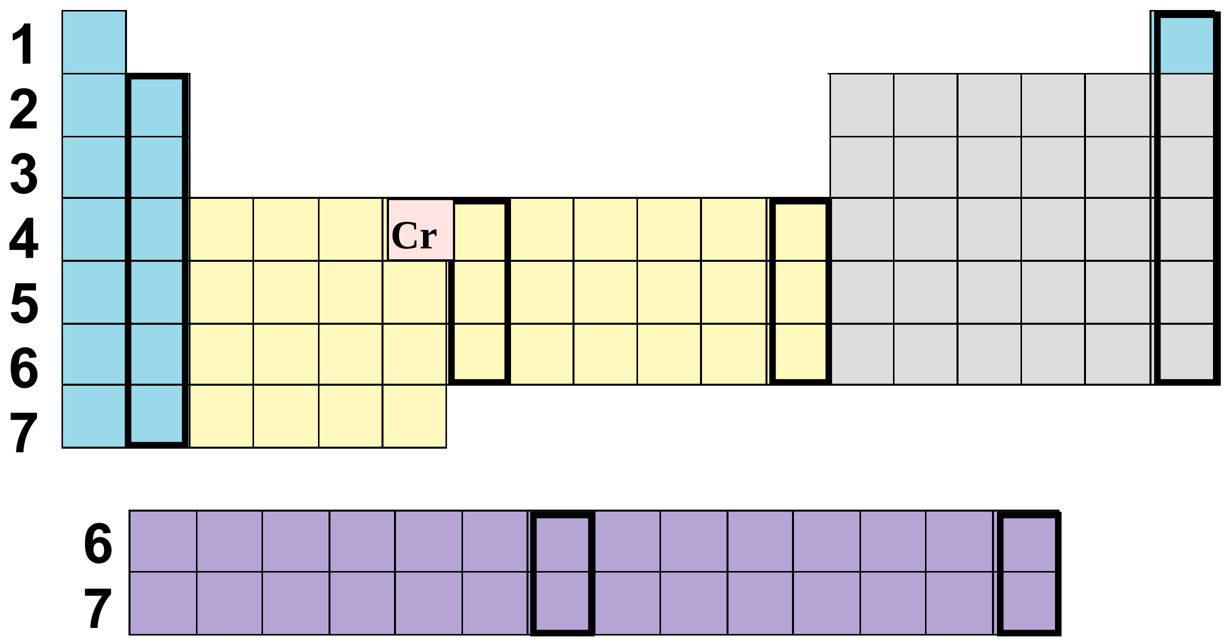


Electron Configuration Exceptions

Chromium

Expect: [Ar] $4s^2$ $3d^4$ Actually: [Ar] $4s^1$ $3d^5$

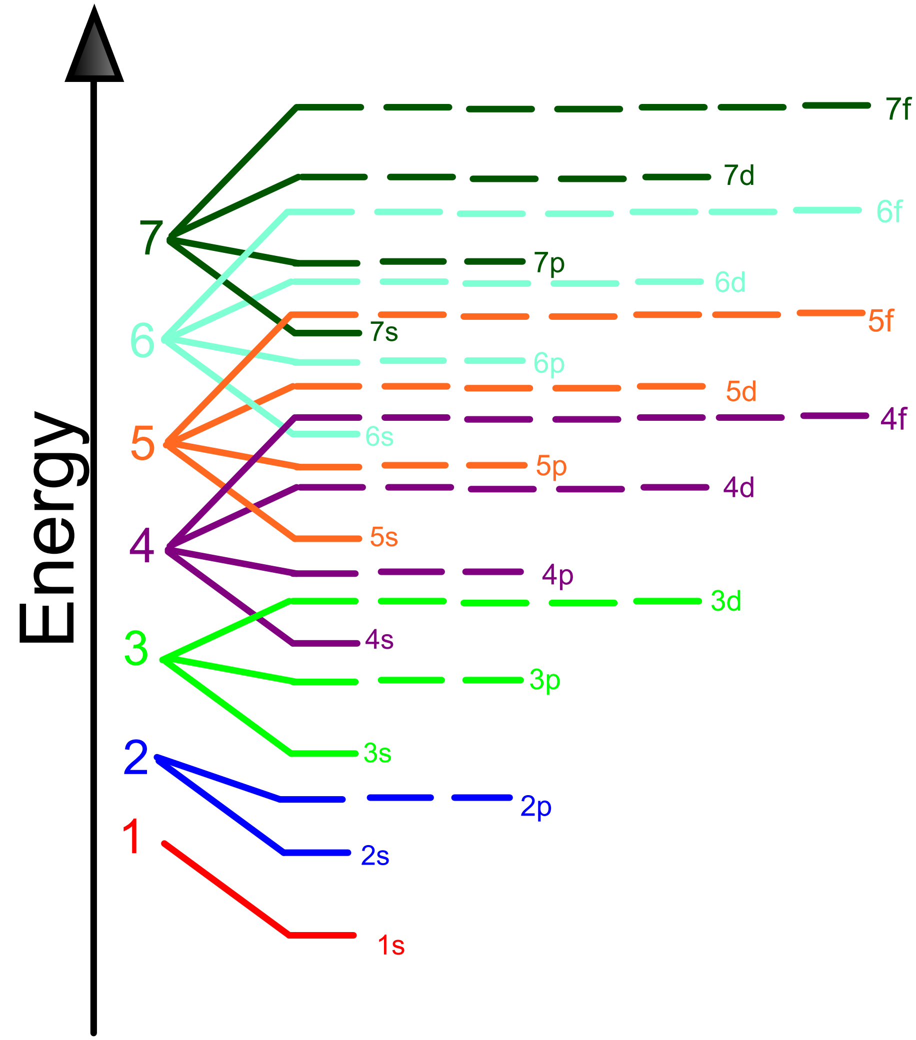
For some elements, in order to exist in a more stable state, electrons from an s sublevel will move to a d sublevel, thus providing the stability of a half-full sublevel. Tosee why this can happen we need to examine how "close" d and s sublevels are.



Energies of Orbitals

Because of how close the f and d orbitals are to the s orbitals, very little energy is required to move an electron from the s orbital (leaving it half full) to the f or d orbital, causing them to also be half full.

(It's kind of like borrowing a cup of sugar from a neighbor).

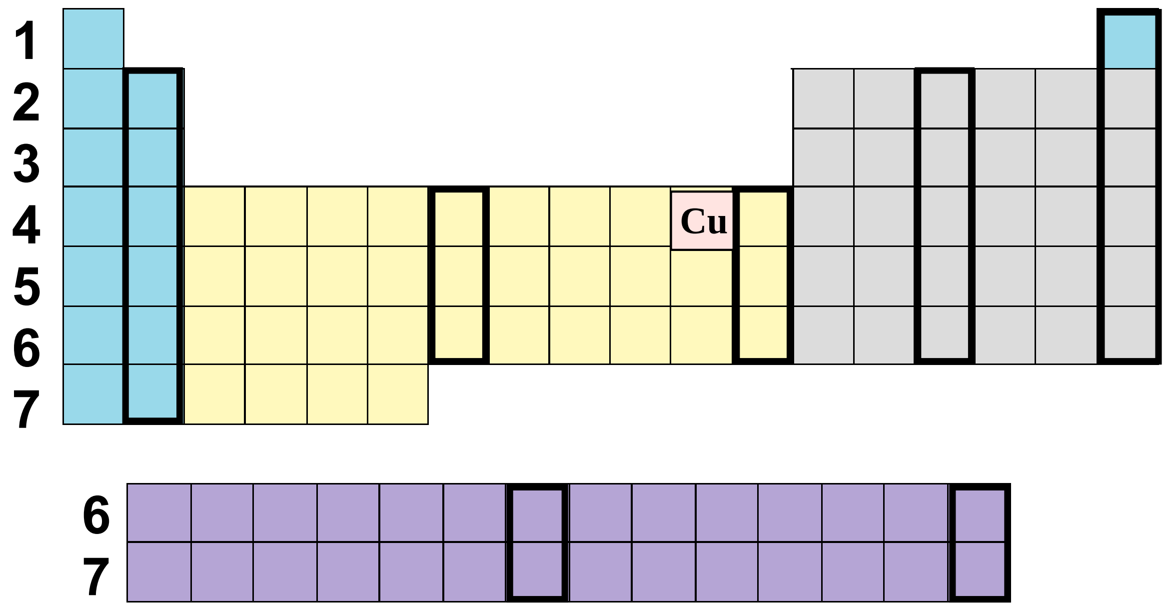


Electron Configuration Exceptions

Copper

Expected: [Ar] $4s^2$ $3d^9$ Actual: [Ar] $4s^1$ $3d^{10}$

Copper gains stability when an electron from the 4s orbital fills the 3d orbital.



27 The electron configuration for Copper (Cu) is

- A [Ar] $4s^2 4d^9$
- B [Ar] $4s^1 4d^9$
- C [Ar] $4s^2 3d^9$
- D [Ar] $4s^1 3d^{10}$

28 What would be the shorthand electron configuration for Silver (Ag)?

- A $[Kr]5s^2 5d^9$
- B $[Ar]5s^1 4d^{10}$
- C $[Kr]5s^2 4d^9$
- D $[Kr]5s^1 4d^{10}$

29 What would be the shorthand electron configuration for Molybdenum (Mb)?

- A $[Kr]5s^2 5d^4$
- B $[Ar]5s^2 4d^4$
- C $[Kr]5s^1 4d^5$
- D $[Kr]5s^2 4d^4$

Effective Nuclear Charge and Coulomb's Law

Return to Table of Contents

Periodic Trends

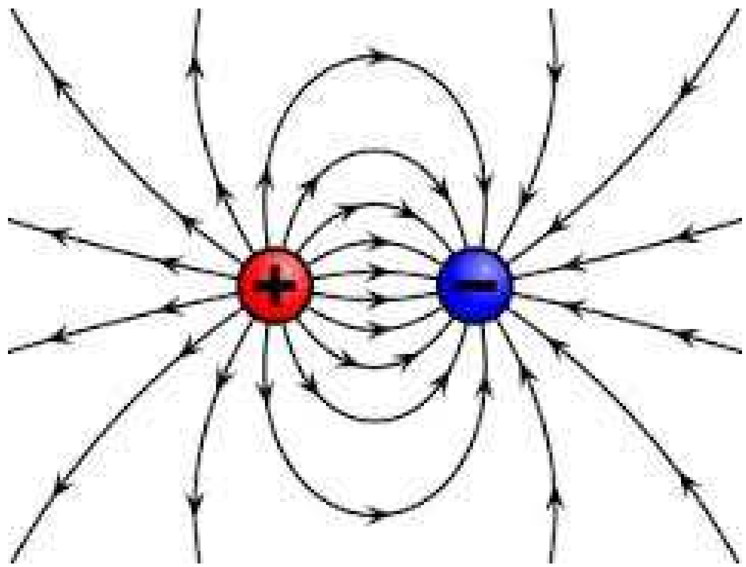
There are four main trends in the periodic table:

- Radius of atoms
- Electronegativity
- Ionizatioin Energy
- Metallic Character

These four periodic trends are all shaped by the interactions between the positive charge of the atomic nucleus and the negative charge of electrons. How do these charges interact with each other?

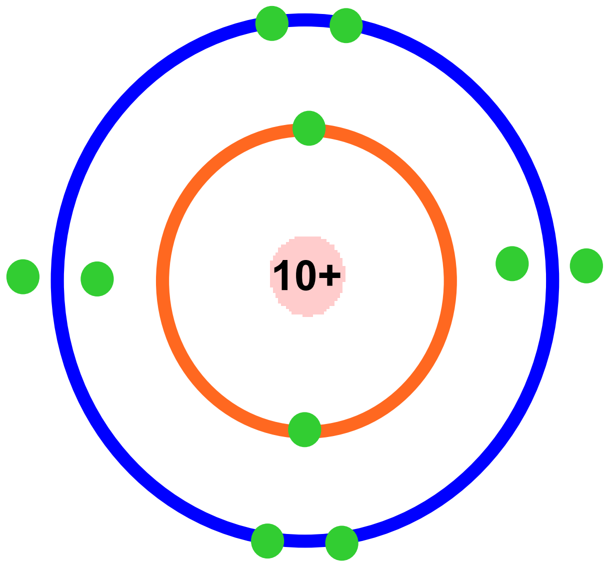
Periodic Trends

Remember that like charges repel and opposite charges attract. The positive protons are attracted to the negative electrons. The negative electrons, on the other hand, are repelled by neighboring electrons.



Atom Diagrams

Atoms of an element are often depicted showing total number of electrons in each energy level, like the diagram below:



For example, Neon's electron configuration:

$1s^2 2s^2 2p^6$

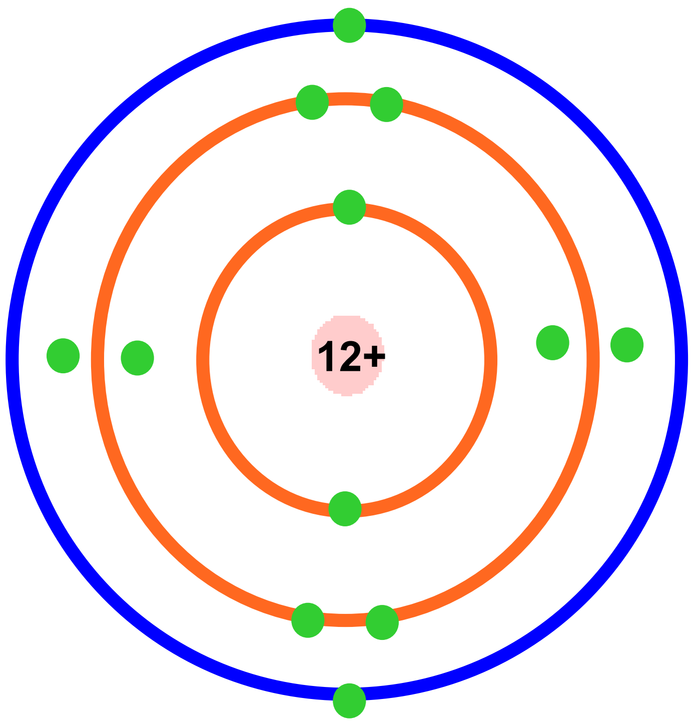
2 electrons
in inner energy levels

8 electrons
in the outer energy level.

These outer electrons are called valence electrons.

30 How many valence electrons does magnesium have?

- A 2
- B 8
- C 10
- D 12



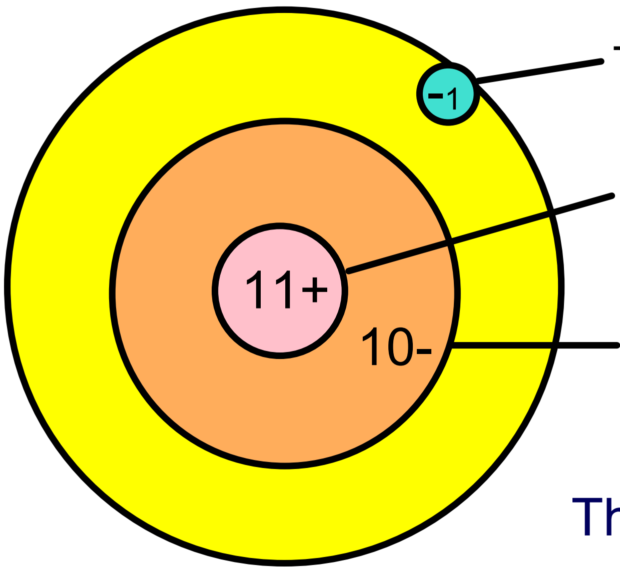
31 Which of the following elements has the largest amount of inner shell electrons: aluminum, silicon or phosphorus?

- A Al
- B Si
- C P
- D They all have the same number of inner shell electrons.

Effective Nuclear Charge

In a multi-electron atom, electrons are both attracted to the positive nucleus and repelled by other electrons.

The nuclear charge that an electron experiences depends on both factors. For example, the valence electron of sodium is attracted to the positive nucleus but is repelled by the negative inner electrons.

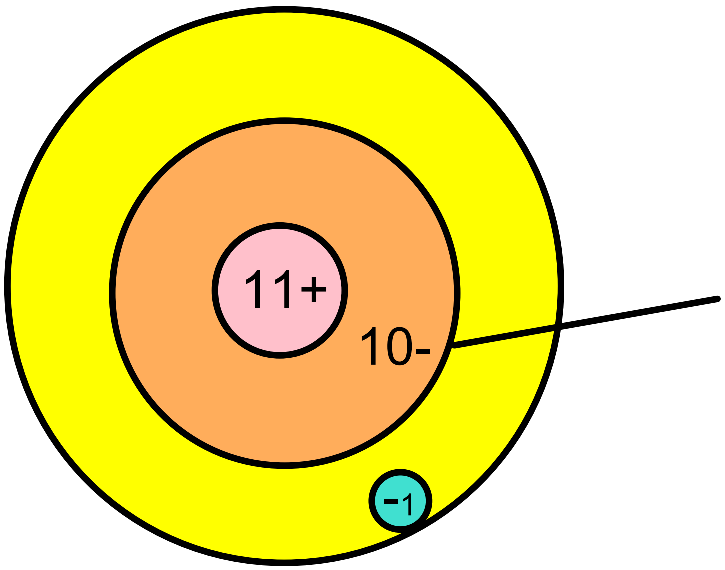


The total charge on the valence electron is:

$+11 + -10 = +1$

Effective Nuclear Charge

The inner shell electrons prevent the valence electron from feeling the full attractive force of the positive protons. In other words, the inner electrons are shielding the valence electrons from the nucleus.



These 10 inner electrons prevent the 1 valence electron from feeling the full attractive force of the 11 protons.

Effective Nuclear Charge

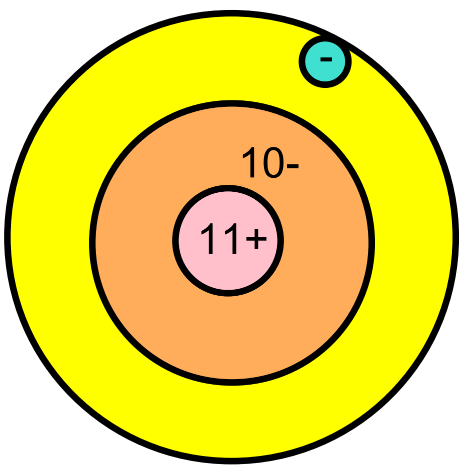
Effective nuclear charge is the amount of charge that the outer electron actually feels.

The formula for effective nuclear charge is:

$$\mathbf{Z_{eff} = Z - S}$$

Z is the atomic number (the number of protons).

S is the shielding constant, the number of inner electrons that shields the valence electrons from the protons.



For sodium:

$Z_{eff} = 11 - 10 = 1$

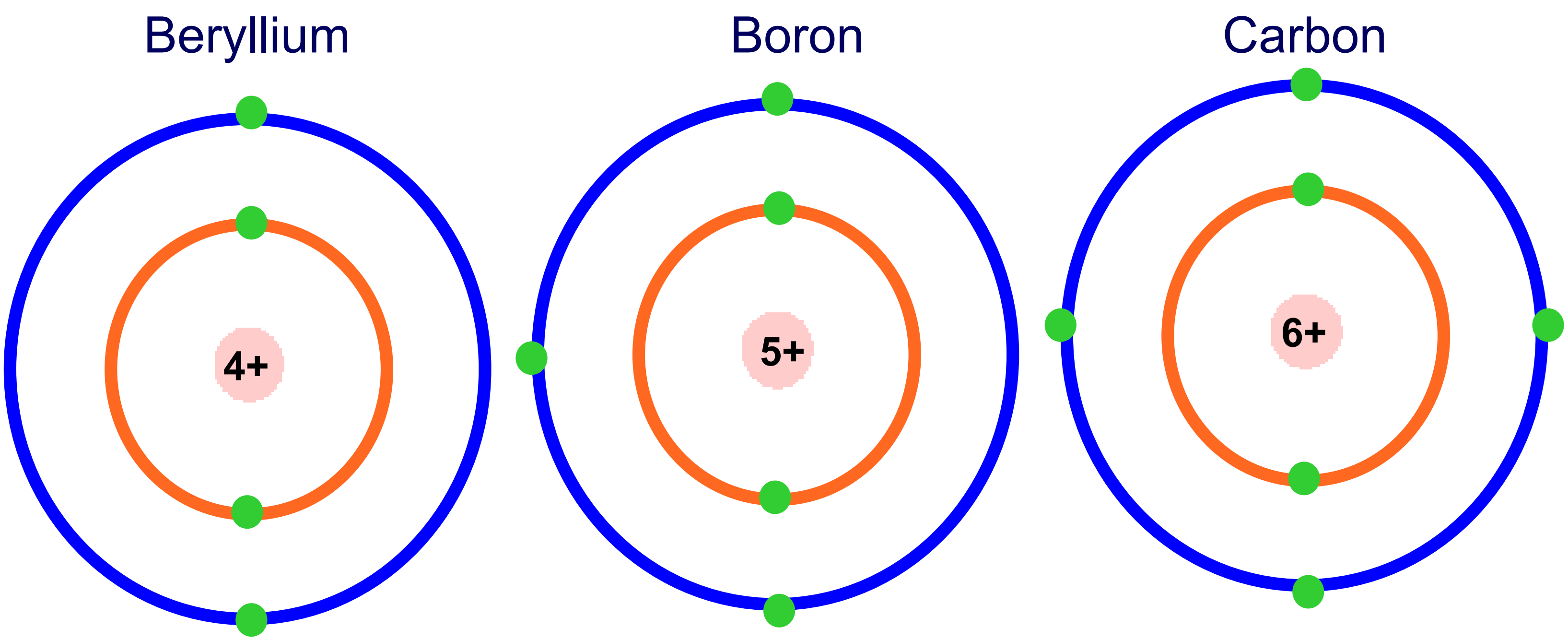
Effective Nuclear Charge

Beryllium, boron and carbon are all in the same period of the periodic table. Compare their shielding constants.

Beryllium	Boron	Carbon
Move for answer.	Move for answer.	Move for answer.

Effective Nuclear Charge

Elements in the same period will have the same shielding constant because their valence electrons are located in the same energy level.



Each has a different atomic number. Boron and carbon have different subshells from beryllium. BUT, they are all in the same energy level, so they have the same number of shielding electrons.

Effective Nuclear Charge

Now look at effective nuclear charge. Compare the values for beryllium, boron and carbon.

Beryllium	Boron	Carbon
Move for answer.	Move for answer.	Move for answer.

What do these values tell you?

32 What is the shielding constant, S, for Boron (B)?

33 What is the effective nuclear charge, Z_{eff} on electrons in the outer most shell for Boron?

34 What is the shielding constant, S, for Aluminum (Al)?

35 What is the effective nuclear charge on electrons in the outer most shell for Aluminum?

36 Which of the following would have the highest effective nuclear charge?

- A Aluminum
- B Phosphorus
- C Chlorine
- D Neon

37 In which subshell does an electron in an arsenic (As) atom experience the greatest shielding?

- A 2p
- B 4p
- C 3s
- D 1s

38 Two elements are studied: one with atomic number X and one with atomic number X+1. Assuming element X is not a noble gas, which element has the larger shielding constant?

- A Element X
- B Element X+1
- C They are both the same.
- D More information is needed.

39 Two elements are studied: one with atomic number X and one with atomic number X+1. It is known that element X is a noble gas. Which element has the larger shielding constant?

- A Element X
- B Element X+1
- C They are both the same.
- D More information is needed.

40 In which subshell does an electron in a calcium atom expe
the greatest effective nuclear charge?

- A 1s
- B 2s
- C 2p
- D 3s

41 Compare the following elements: potassium, cobalt and selenium. Which atom feels the strongest attractive force between the nucleus and the valence electrons?

- A K
- B Co
- C Se
- D They all experience the same magnitude of force.

Coulomb's Law

The magnitude of the force between the protons in the nucleus and electrons in the orbitals can be calculated using Coulomb's Law.

$$F = \frac{kq_1 q_2}{r^2}$$

k = Coulomb's constant

q_1 = the charge on the first object

q_2 = the charge on the second object

r^2 = the distance between the two objects

42 According to Coulomb's Law, the stronger the charge of the objects, the ___ the force between the objects.

- A stronger
- B weaker

$$F = \frac{kq_1 q_2}{r^2}$$

43 According to Coulomb's Law, the greater the distance between two objects, the ___ the force between the objects.

A stronger

B weaker

$$F = \frac{kq_1 q_2}{r^2}$$

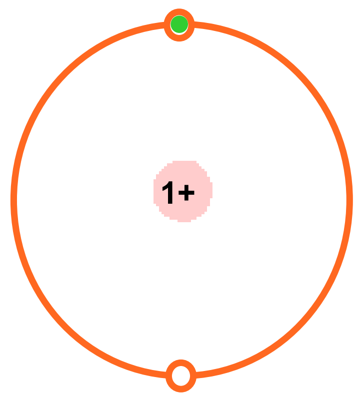
Hydrogen

Applying Coulomb's Law to atoms provides useful information about those atoms.

Consider hydrogen. Z_{eff} for hydrogen is 1.

Z_{eff} = 1 proton - 0 inner electron

Z_{eff} = 1



The charge between the valence electron and the nucleus is 1e.

Plugging this into Coulomb's Law:

$$F = \frac{kq_1 q_2}{r^2} \longrightarrow F = \frac{kZ_{eff}(e)^2}{r^2} \longrightarrow F = \frac{ke^2}{r^2}$$

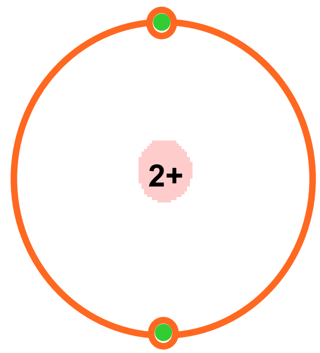
Helium

Now let's apply Coulomb's Law to helium.

Z_{eff} for hydrogen is 2.

Z_{eff} = 2 protons - 0 inner electron

Z_{eff} = 2



The charge between the valence electron and the nucleus is 2e.

Plugging this into Coulomb's Law:

$$F = \frac{kq_1 q_2}{r^2} \longrightarrow F = \frac{kZ_{eff}(e)^2}{r^2} \longrightarrow F = \frac{k(2e)^2}{r^2}$$

Hydrogen vs Helium

Now we can compare hydrogen and helium.

Hydrogen

The force between the valence electron and the nucleus is:

$$F = \frac{ke^2}{r^2}$$

Helium

The force between the valence electrons and the nucleus is:

$$F = \frac{k(2e)^2}{r^2}$$

(Initially, the radius is the same for both since both have valence electrons in the same energy level.)

The force between the nucleus and the electrons in helium is much larger than the force between the nucleus and the electron in hydrogen.

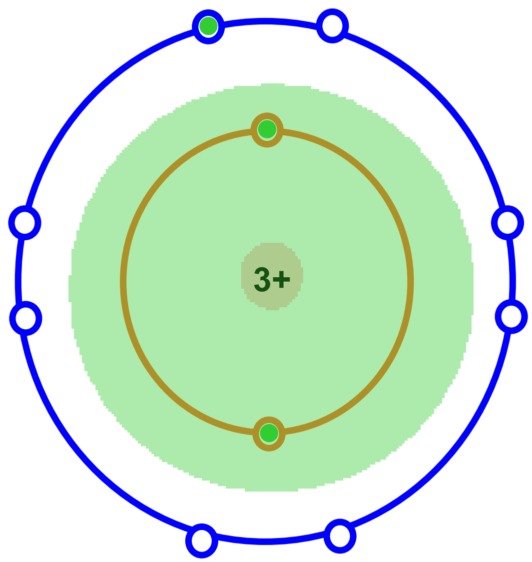
How does this affect the radii of the atoms?

Lithium

$Z_{eff} = Z - S$

$Z_{eff} = 3 - 2$

$Z_{eff} = 1$

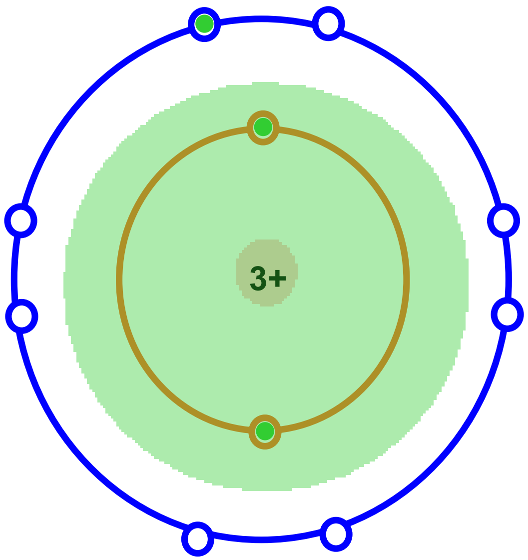


Plugging this into Coulomb's Law:

$$F = \frac{kq_1 q_2}{r^2} \longrightarrow F = \frac{kZ_{eff}(e)^2}{r^2} \longrightarrow F = \frac{ke^2}{r^2}$$

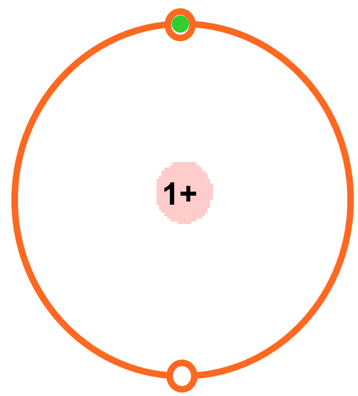
Lithium vs Hydrogen

Lithium



$$F = \frac{ke^2}{r^2}$$

Hydrogen



$$F = \frac{ke^2}{r^2}$$

The Z_{eff} is the same for both atoms. However, lithium has valence electrons in a higher energy level.

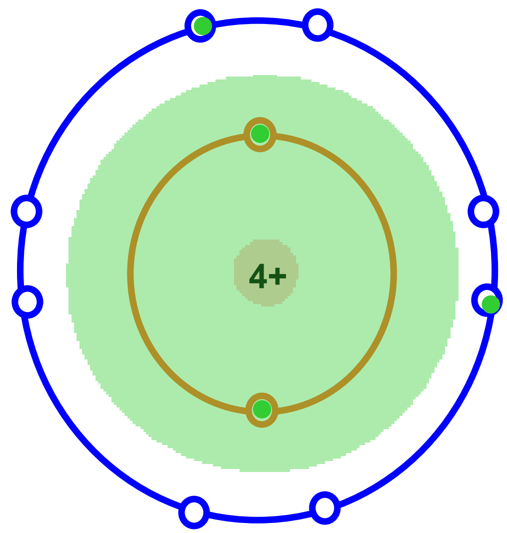
How does this affect the radii of the atoms?

Beryllium

$Z_{eff} = Z - S$

$Z_{eff} = 4 - 2$

$Z_{eff} = 2$



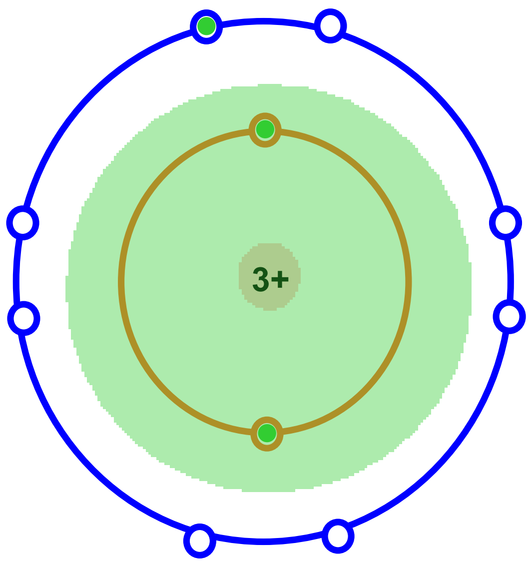
Plug this into Coulomb's Law.

$$F = \frac{k q_1 q_2}{r^2}$$

Slide for answer.

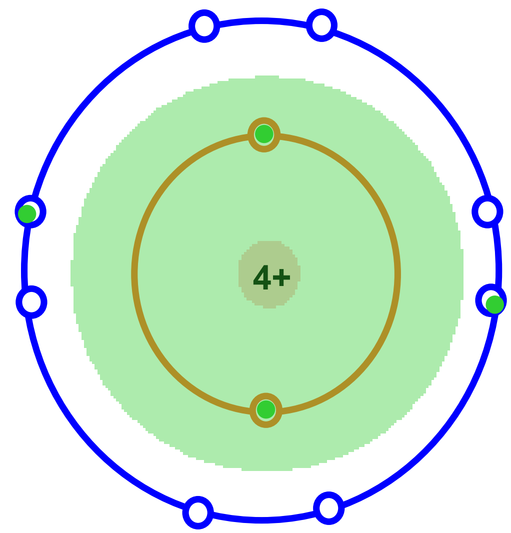
Lithium vs Beryllium

Lithium



$$F = \frac{ke^2}{r^2}$$

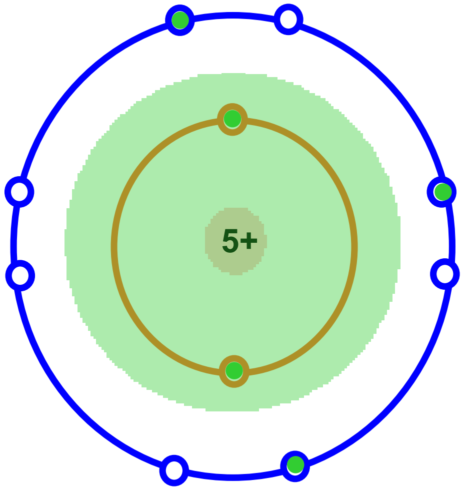
Beryllium



$$F = \frac{k(2e)^2}{r^2}$$

How do the radii of beryllium and lithium compare?

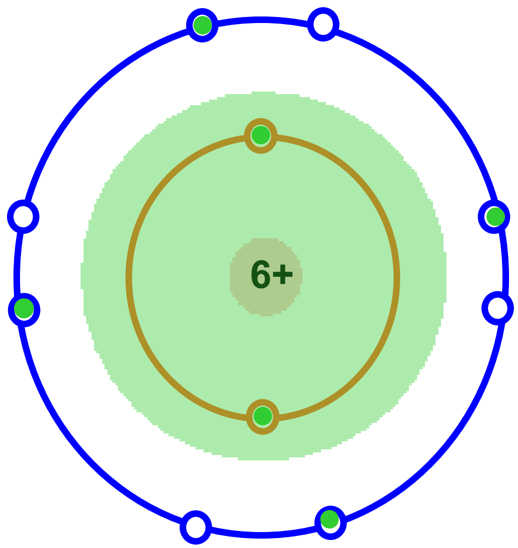
44 What is Z_{eff} for Boron (B)?



45 Compare the radial size of boron to lithium and beryllium.

- A Li>Be>B
- B Li<Be<B
- C Li>B>Be
- D Be<Li<B

46 What is Z_{eff} for Carbon (C)?



47 Compare the radial size of carbon to boron and nitrogen.

- A C>N>B
- B C<N<B
- C B>C>N
- D B<C<N

48 Which of the following equations correctly calculates the Coulombic force between the valence electrons and the nucleus of an oxygen atom?

- A $F = k(2e)^2/r^2$
- B $F = k(4e)^2/r^2$
- C $F = k(6e)^2/r^2$
- D $F = k(8e)^2/r^2$

49 Give the atomic number of the smallest element in the 2nd period.

Periodic Trends: Atomic Radius

Return to Table
of Contents

Atomic Radii Trend

50

Students type their answers here

What is the trend in atomic size across a period?
What is the trend in atomic size down a group?
(Pull the box away to see the answers.)

51 Across a period from left to right Z_{eff} ____.

- A increases
- B decreases
- C remains the same

52 Down a group from top to bottom Z_{eff} ____.

- A increases
- B decreases
- C remains the same

53 Atomic radius generally increases as we move __________.

- A down a group and from right to left across a period
- B up a group and from left to right across a period
- C down a group and from left to right across a period
- D up a group and from right to left across a period

54 Which one of the following atoms has the smallest radius?

- A O
- B F
- C S
- D Cl

55 Which one of the following atoms has the largest radius?

- A Cs
- B Al
- C Be
- D Ne

56 Which one of the following atoms has the smallest radius?

- A Fe
- B N
- C S
- D I

57 Of the following, which gives the correct order for atomic radius for Mg, Na, P, Si and Ar?

- A Mg > Na > P > Si > Ar
- B Ar > Si > P > Na > Mg
- C Si > P > Ar > Na > Mg
- D Na > Mg > Si > P > Ar

58 Which of the following correctly lists the five atoms in order of increasing size (smallest to largest)?

- A $O < F < S < Mg < Ba$
- B $F < O < S < Mg < Ba$
- C $F < O < S < Ba < Mg$
- D $F < S < O < Mg < Ba$

59 Two elements are studied. One with atomic number X and one with atomic number X+1. Assuming element X is not a Noble Gas, which element has the larger atomic radius?

- A Element X
- B Element X+1
- C They are both the same.
- D More information is needed.

60 Two elements are studied. One with atomic number X and one with atomic number X+1. It is known that element X is a Noble Gas. Which element has the larger atomic radius?

- A Element X
- B Element X+1
- C They are both the same.
- D More information is needed.

Summary of Atomic Radius Trends

- Across a period, effective nuclear charge increases while energy level remains the same. The force of attraction between the nucleus and valence electrons gets stronger. Valence electrons are pulled in tighter, so radius gets smaller.

$$F = \frac{kq_1 q_2}{r^2}$$

This value gets larger, so force is larger. (Radius is smaller.)

- Down a period, effective nuclear charge remains the same while the energy level increases. The increased distance from the nucleus to valence electrons makes the force of attraction decrease. Electrons are not held as tightly, so radius gets larger.

$$F = \frac{kq_1 q_2}{r^2}$$

This value gets larger, so force is smaller. (Radius is larger.)

Click here for an animation on the atomic radius trend.

Periodic Trends: Ionization Energy

Return to Table of Contents

.

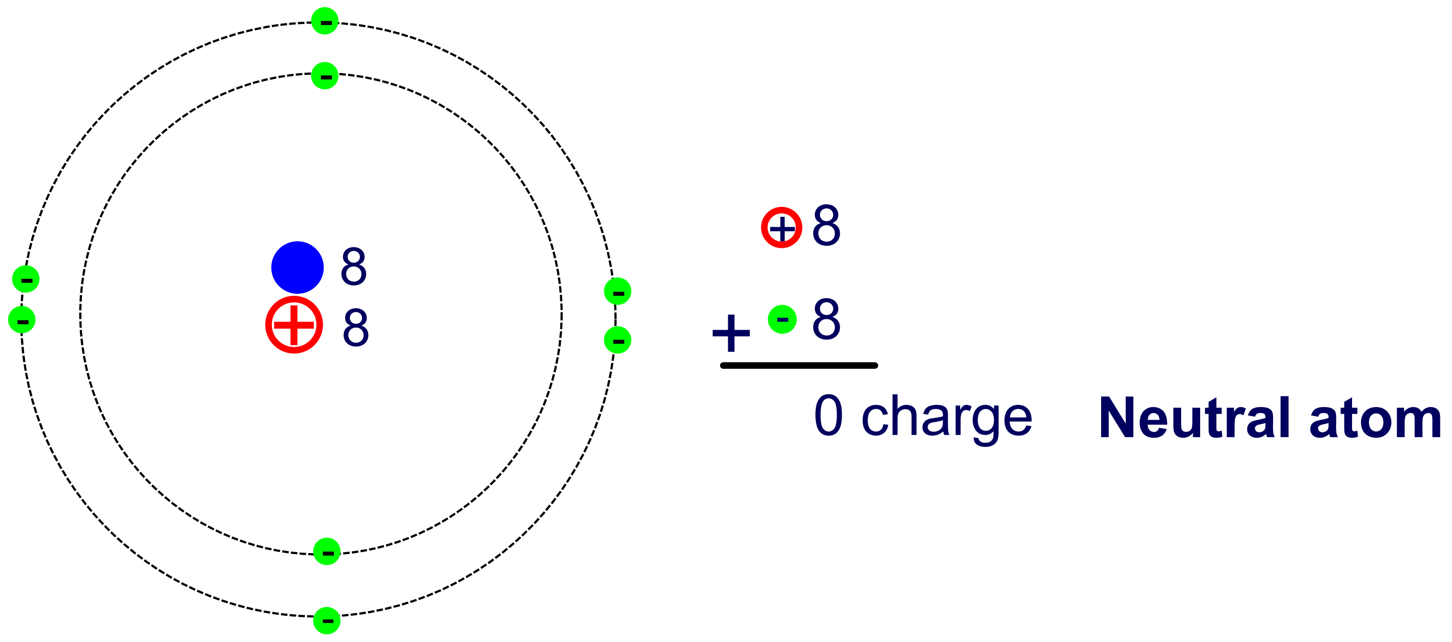
Ionization Energy

Atoms of the same element have equal numbers of

protons and electrons.

Neutral Oxygen ---> 8 (+) protons and 8 (-) electrons

Neutral Magnesium ---> 12 (+) protons and 12 (-) electrons

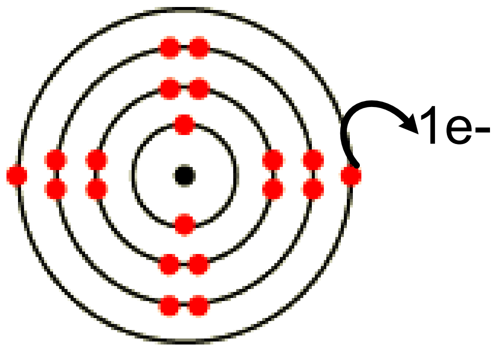


Ionization Energy

The ionization energy is the amount of energy required toremove an electron from an atom.

Removing an electron creates a positively charged atom called a cation.

Calcium cation



$$Ca \longrightarrow Ca^{+} + e^{-}$$

⊕ 20

\+ ● 19

+1 charge

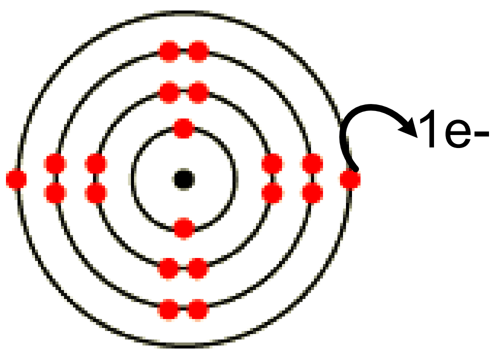
Ionization Energy

The ionization energy is the amount of energy required toremove an electron from an atom. Removing an electron creates a positively charged atom called a cation.

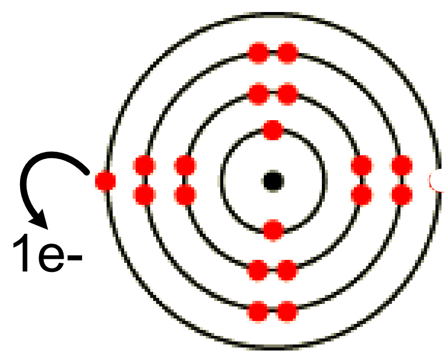
The first ionization energy is the energy required to remove the first electron.

The second ionization energy is the energy required to remove the second electron, etc.

$Ca \longrightarrow Ca^{+} + e^{-}$



$Ca^{+} \longrightarrow Ca^{2+} + e^{-}$

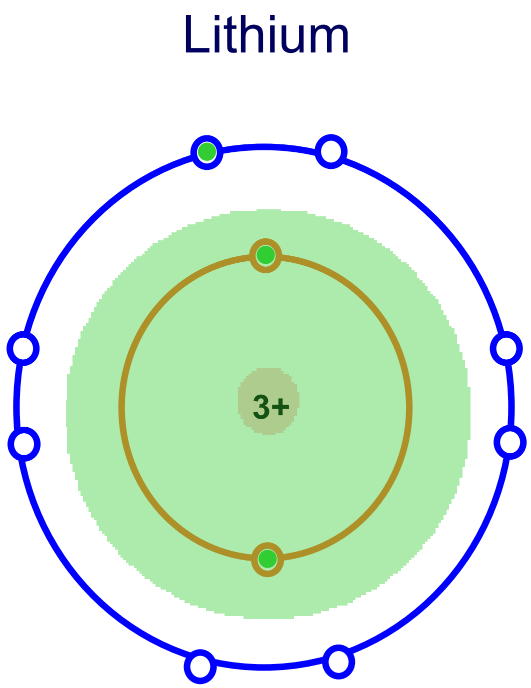


61 If an electron is removed from a sodium (Na) atom, what charge does the Na cation have?

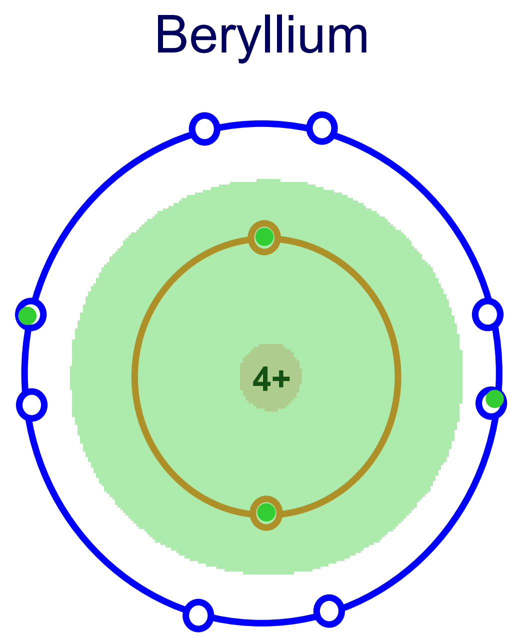
62 If two electrons are removed from a Magnesium (Mg) atom, what charge does the Mg cation have?

Lithium vs Beryllium

Applying Coulomb's Law helps us to understand how ionization energy changes among elements.



$$F = \frac{ke^2}{r^2}$$

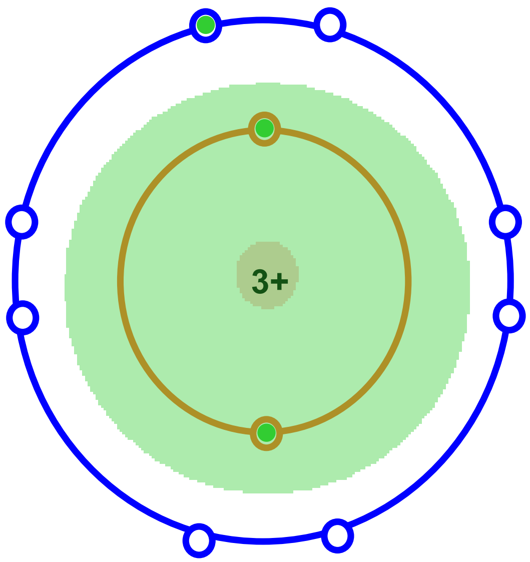


$$F = \frac{k(2e)^2}{r^2}$$

Which atom is held together more closely?

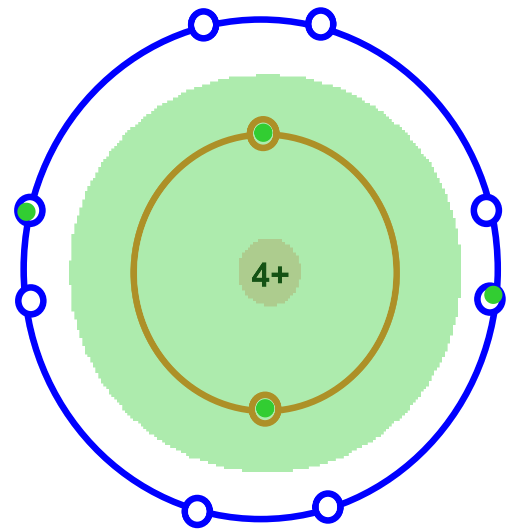
Lithium vs Beryllium

Lithium



$$F = \frac{ke^2}{r^2}$$

Beryllium



$$F = \frac{k(2e)^2}{r^2}$$

Since beryllium holds onto its electrons tighter, it will require more energy to take away an electron. The ionization energy of beryllium is higher than lithium.

Ionization Energy and Coulomb's Law

As the force increases, the atom holds onto electrons tighter. These electrons will require more energy (ionization energy) to take them away than an atom with a lower force.

As force increases, ionization energy increases.

Think back to atomic radius. How does atomic radius relate to Coulomb's Law? How does it relate to ionization energy?

Trends in First Ionization Energies

Compare ionization energies for magnesium, aluminum and silicon.

First, find Coulomb's equation for each. Then, order the elements in increasing ionization energy.

Magnesium	Aluminum	Silicon
Pull for answer	Pull for answer	Pull for answer

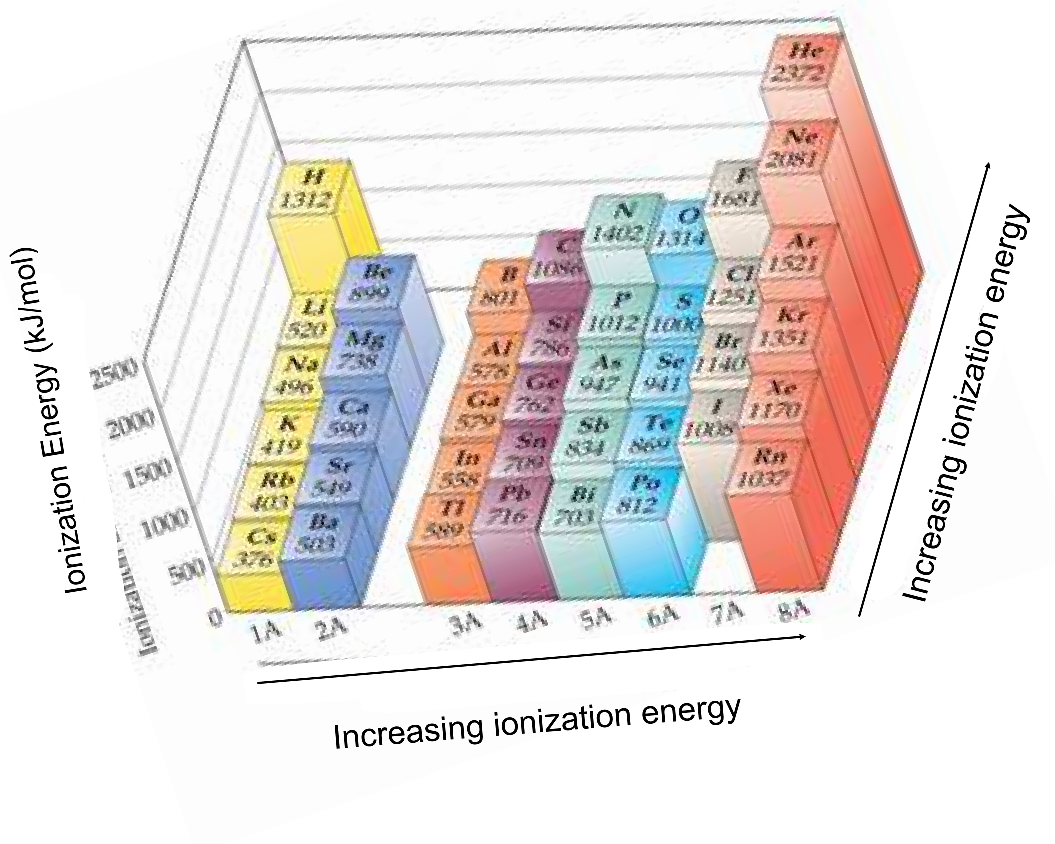
Increasing order of ionization energies: Pull for answer

How does ionization energy change as you go across a period?

Trends in First Ionization Energies

Across a period, Z_{eff} increases and the force on electrons increases. This makes it harder for an electron to be taken away.

Ionization energy increases across a period.



Trends in First Ionization Energies

Compare ionization energies for sodium and potassium.

First, find Coulomb's equation for each. Then, order the elements in increasing ionization energy.

Sodium

Pull for answer

Potassium

Pull for answer

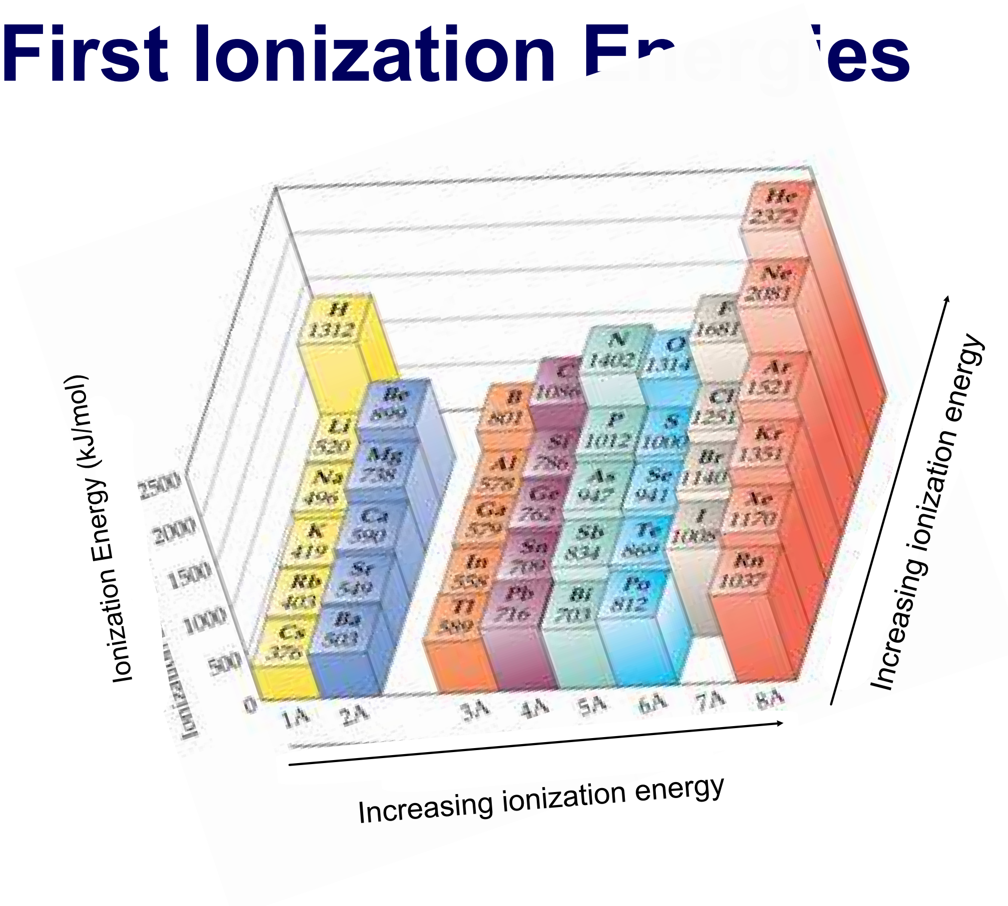
Increasing order of ionization energies: Pull for answer

How does ionization energy change as you go down a group?

Trends in First Ionization Energies

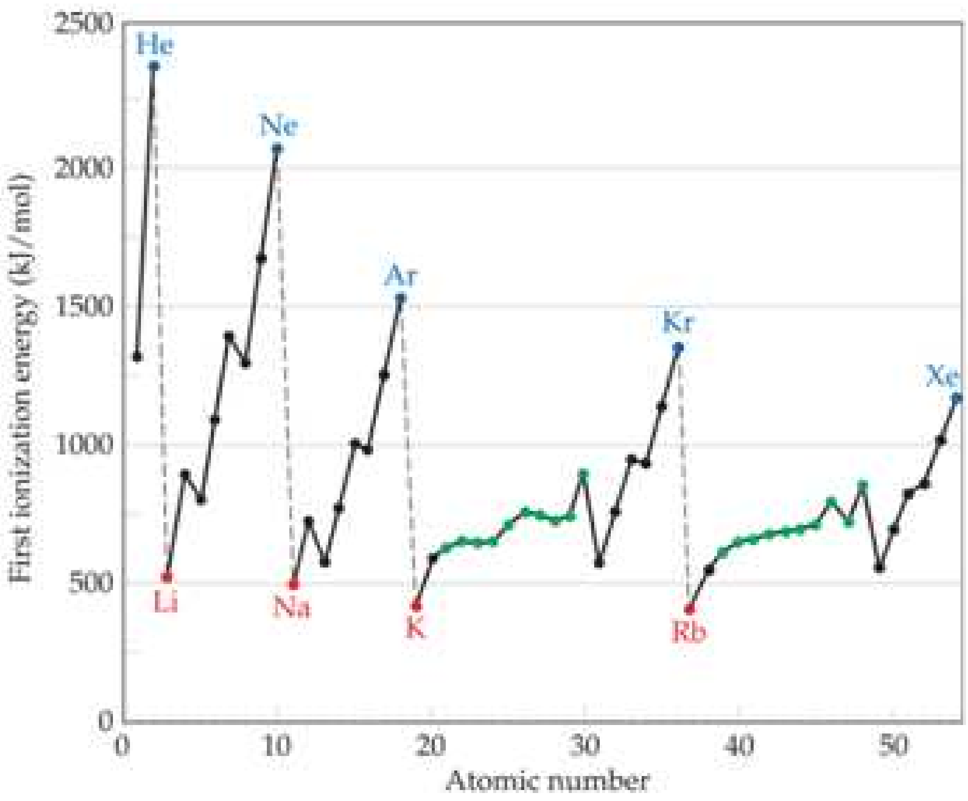
Down a group, Z_{eff} stays the same but the extra energy levels make the radius larger which make the force less. It is easier to take electrons away.

Ionization energy decreases as you go down a period.



Click here for an animation on Ionization Energy

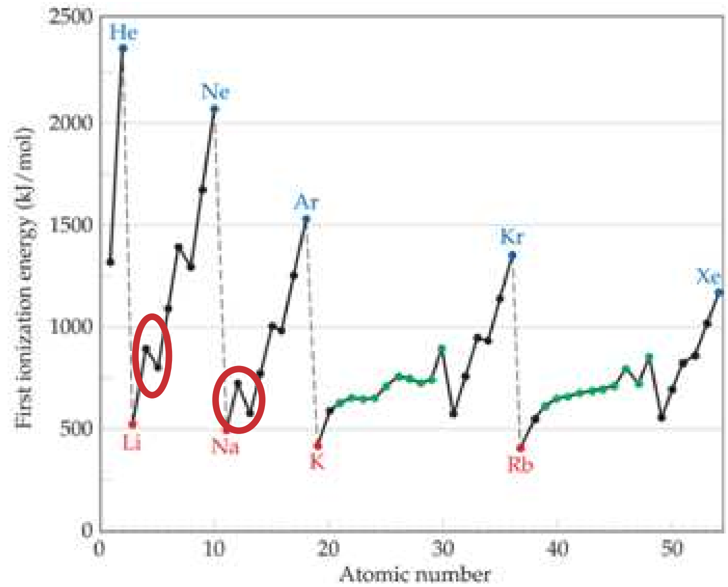
Trends in First Ionization Energies



However, there are two apparent discontinuities in this trend.

Discontinuity #1

The first is between Groups 2 and 13 (3A). As you can see on the chart to the right, the ionization energy actually decreases from Group 2 to Group 13 elements. The electron removed for Group 13 elements is from a *p* orbital and removing this electron actually adds stability.



The electron removed is farther from nucleus, there is a small amount of repulsion by the *s* electrons.

The atom gains stability by having a full *s* orbital, and an empty *p* orbital.

Discontinuity #1

More energy is required to remove an electron from Group 2 elements than Group 13 elements. Draw the orbital diagrams for Group 2 Boron and Group 13 Beryllium to illustrate why.

Boron ____ ____ ____ ____ ____

1s 2s 2p

Beryllium ____ ____ ____ ____ ____

1s 2s 2p

The atom gains stability by having a full *s* orbital, and an empty *p* orbital.

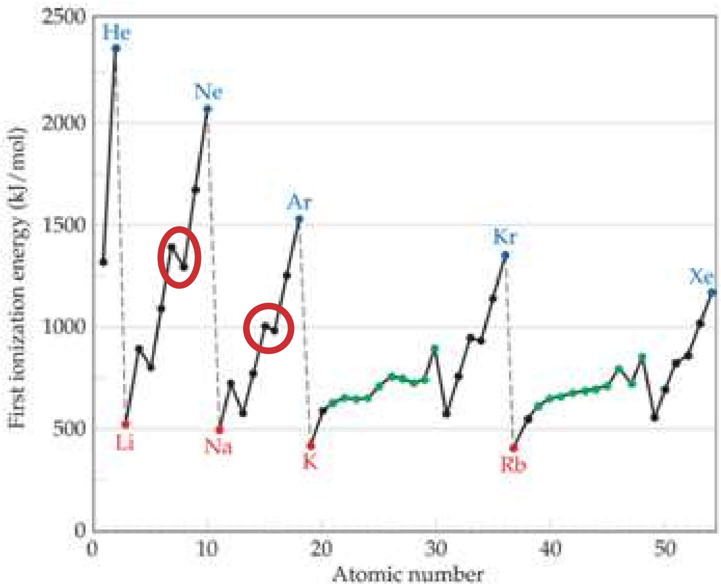
63

Discontinuity #2

Students type their answers here

The second is between Groups 15 and 16.

Using your knowledge of electron configurations and the stability of atoms explain why the first ionization energy for a Group 16 element would be less than that for a Group 15 element in the same period.



64 Of the elements below, ___________ has the largest first ionization energy.

- A Li
- B K
- C Rb
- D H

65 Of the following atoms, which has the largest first ionization energy?

- A Br
- B O
- C C
- D P

66 Of the following elements, which has the largest first ionization energy?

- A Na
- B Al
- C Se
- D Cl

67 Which noble gas has the lowest first ionization energy (enter the atomic number)?

Periodic Trends: Electronegativity

Return to Table of Contents

Electronegativity

Electronegativity is the ability of an atom to attract other electrons.

Using Coulomb's Law, an atom with a high attractive force with its own electrons will also have a high attractive force with other electrons.

Use Coulomb's Law to rank boron, carbon and nitrogen in terms of increasing force.

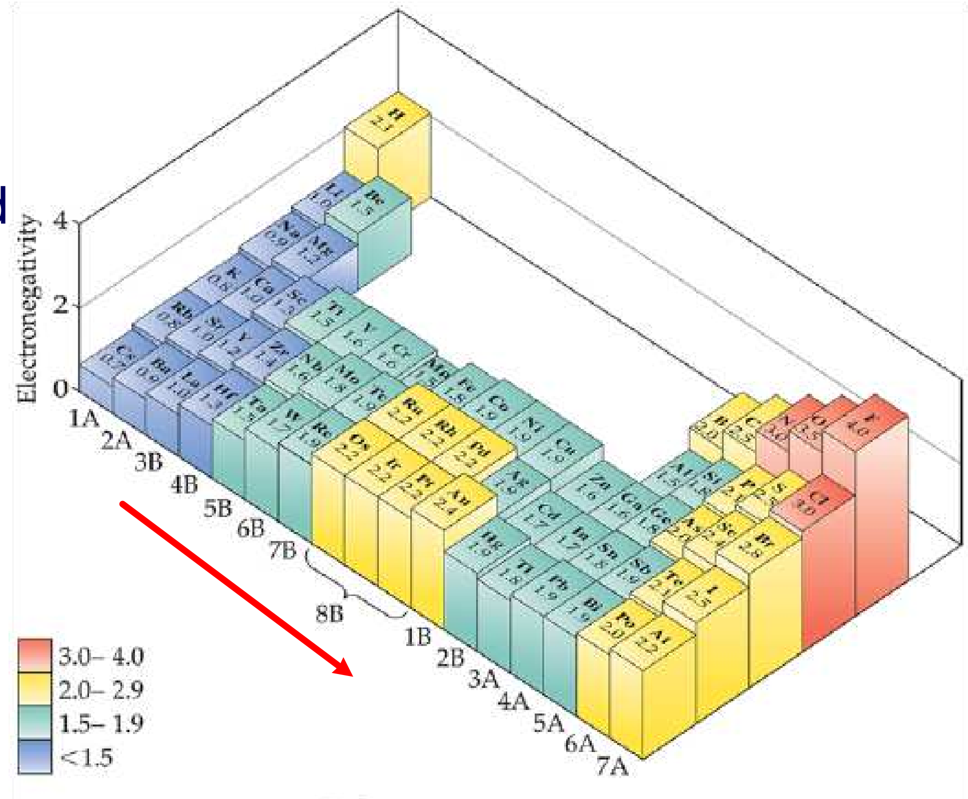
Pull for answer →

How does electronegativity relate to ionization energy and atomic radius?

Electronegativity Trends

As you go across a period, the Z_{eff} increases and the force between nucleus and electrons increases. As this force increases, it is easier for the atom to attract other electrons, so electronegativity increases.

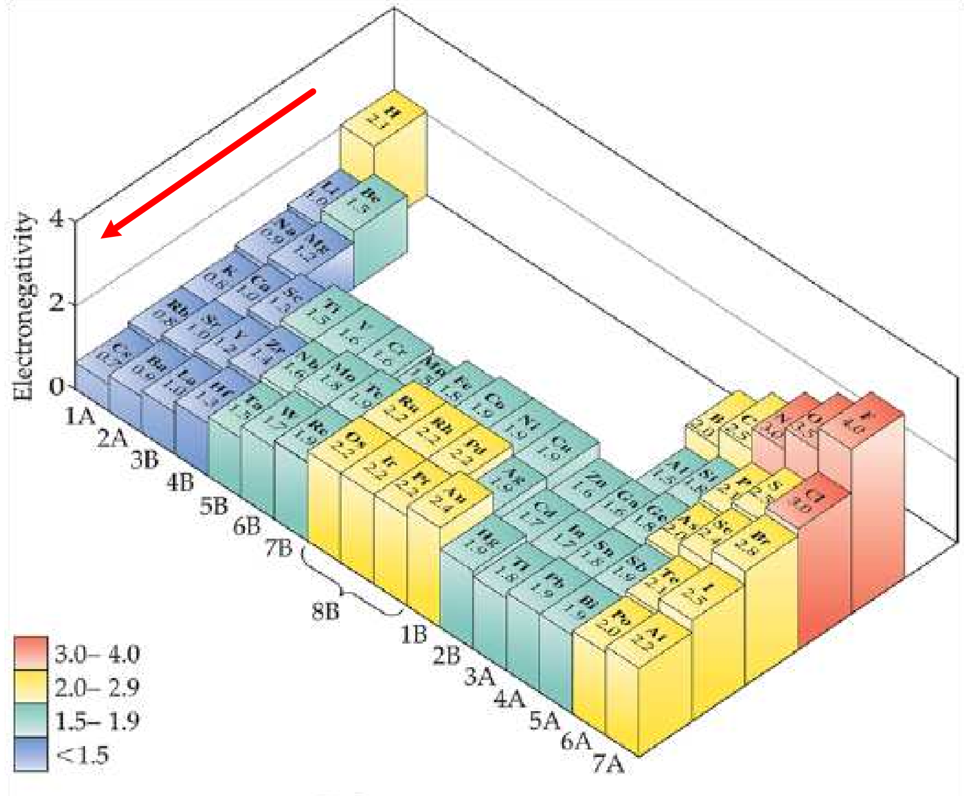
Electronegativity increases as you go across a period.



Electronegativity Trends

As you go down a group, the increased energy levels increase the radius. The force between nucleus and electrons decreases and it is harder for the atom to attract other electrons.

Electronegativity decreases down a group.



Electronegativity Exception #1

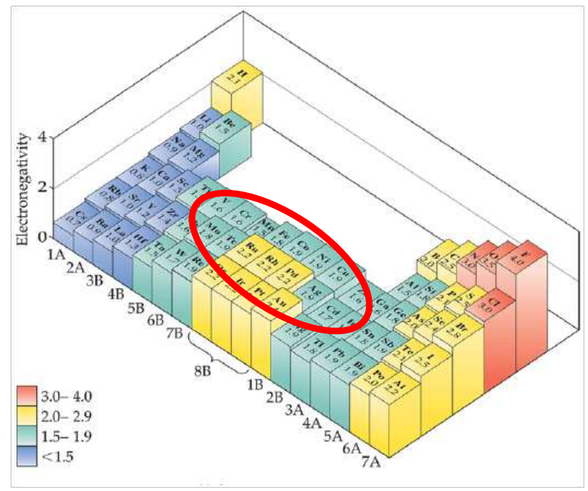
The Noble Gases are some of the smallest atoms, but are usually left out of electronegativity trends since they neither want electrons nor want to get rid of electrons.

Using your knowledge of electron configurations, why do you think noble gases are left out of electronegativity trends?

Electronegativity Exception #2

The Transition Metals have some unexpected trends in electronegativity because of their d and sometimes f orbitals.

The electrons located inthe 3d orbitals (and all d and f orbitals after that) do not contribute as much to the shielding constants of the elements as electrons in the s and p orbitals.



As such, elements with configurations that end in a **d** or **f** orbital will frequently have atomic radii that do not match up with the normal trend.

68 The ability of an atom in a molecule to attract electrons is best quantified by its ___________.

- A electronegativity
- B electron charge-to-mass ratio
- C atomic radius
- D number of protons

69 Electronegativity ___________ from left to right within a period and ___________ from top to bottom within a group.

- A decreases, increases
- B increases, increases
- C increases, decreases
- D decreases, decreases

70 Which of the following correctly ranks the elements from highest to lowest electronegativity?

- A Cl > S > P
- B Br > Cl > F
- C K > Na > Li
- D N > O > F

Summary of Electronegativity & First Ionization Energy Trends

Electronegativity & Ionization Energy increases left to right across a period.

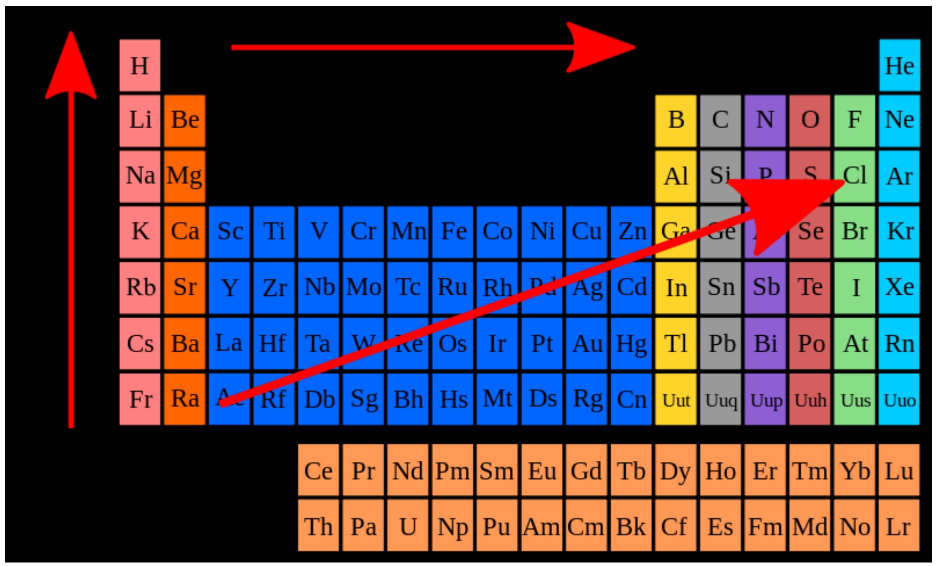
Z_{eff} increases and the force of attraction between the nucleus and valence electrons is strengthened. More energy is required to remove these electrons.

Electronegativity & First Ionization Energy decrease going down a group.

The size of shells increases significantly. The distance between the nucleus and outer electrons increases. The force of attraction decreases.

*Additional Ionization Energies

It requires more energy to remove each successive electron.
ie: *second* ionization energy is greater than *first*,
third ionization energy is greater than *second*, etc.



When all valence electrons have been removed, leaving the atom with a full p subshell, the ionization energy becomes incredibly large.

71 An atom has the following values for its first four ionization energies. Which of the following elements would fit this data?

- A Li
- B Be
- C C
- D F

1st IE = 899.5 kJ/mol
2nd IE = 1,757 kJ/mol
3rd IE = 14,849 kJ/mol
4th IE = 21,007 kJ/mol

Periodic Trends:
Metallic Character

Return to Table
of Contents

Metallic Character

For a metal to conduct electricity or heat, it needs to have electrons that are free to move through it, not tightly bound to a particular atom.

The metallic character of an element is a measure of how loosely it holds onto its outer electrons.

72

Metallic Character

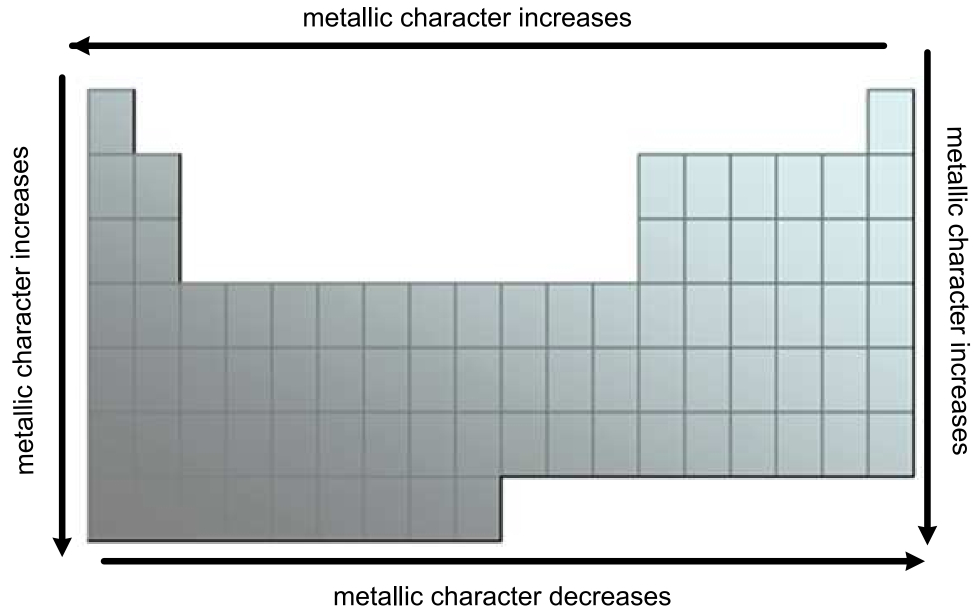
Students type their answers here

So the metallic character of an element is inversely related to its electronegativity.

On the periodic chart, metallic character increases as you go…

from right to left across a row.

from the top to the bottom of a column.



What is the relationship between first ionization energy and metallic character?

73 Because of the relationship between metallic character and electronegativity, you can say that metals tend to ___.

- A take in electrons, becoming positive.
- B give off electrons, becoming negative.
- C take in electrons, becoming negative.
- D give off electrons, becoming positive.

74 Of the elements below, _____ is the most metallic.

- A Sodium
- B Magnesium
- C Calcium
- D Cesium

75 Which of the elements below is the most metallic.

- A Na
- B Mg
- C Al
- D K

76 Which of the atoms below is the most metallic?

- A Br
- B O
- C Cl
- D N

77 Which of the atoms below is the most metallic?

- A Si
- B Cl
- C Rb
- D Ca