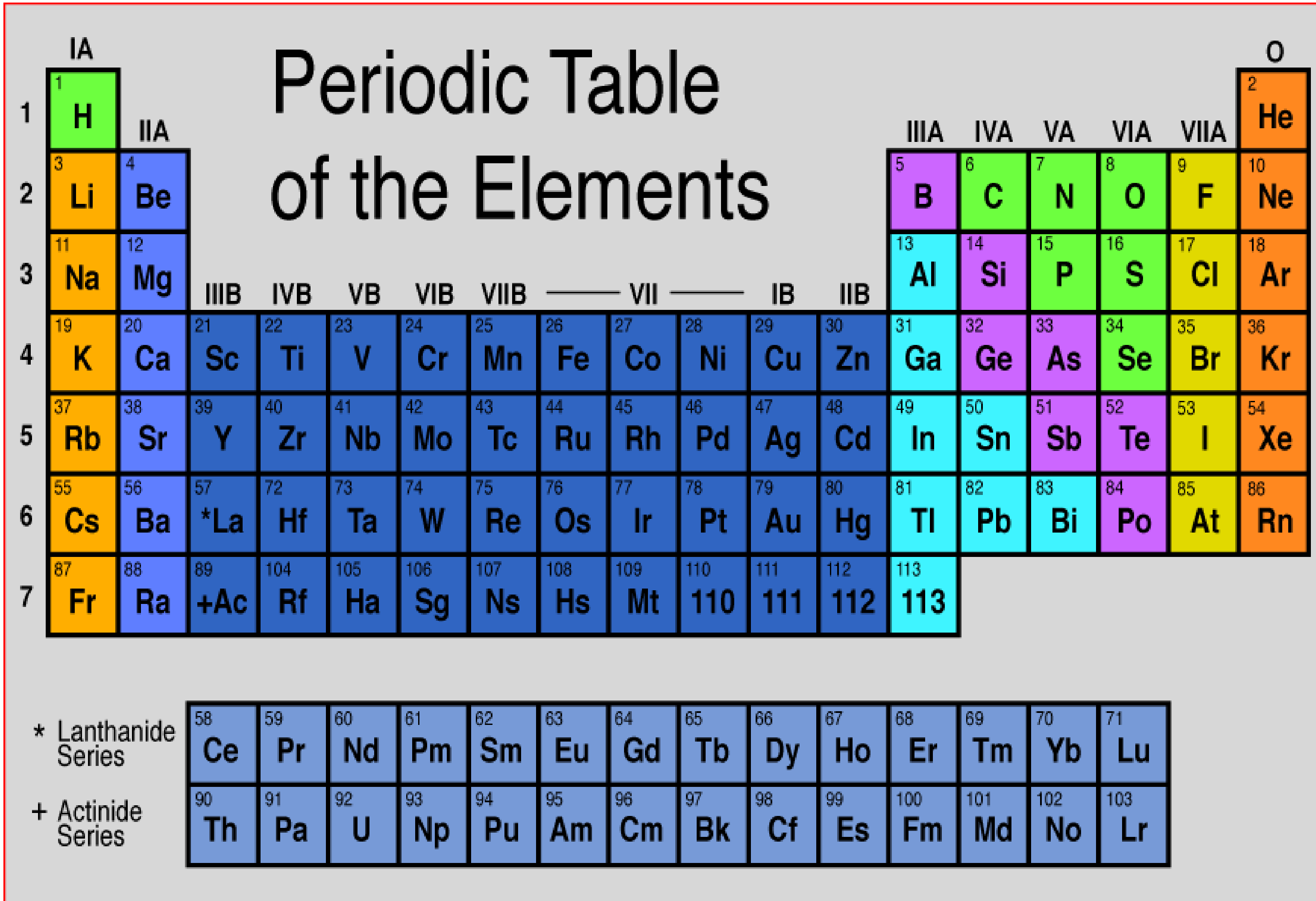
Chapter 6
The Periodic Table



Organizing the Periodic Table

In a grocery store, the products are grouped according to similar characteristics.

With a logical classification system, finding and comparing products is easy.

Similarly, elements are arranged in the periodic table in an organized manner.

Chemists used the properties of elements to sort them into groups.

Mendeleev’s Periodic Table

A Russian chemist and teacher, Dmitri Mendeleev, published a table of the elements in 1869.

Mendeleev developed his table while working on a textbook for his students. He need a way to show the relationship between more than 60 elements.

He wrote the properties of each element on a separate note card. This approach allowed him to move the cards around until he found an organization that worked.

The organization he chose was the periodic table.

The Periodic Law

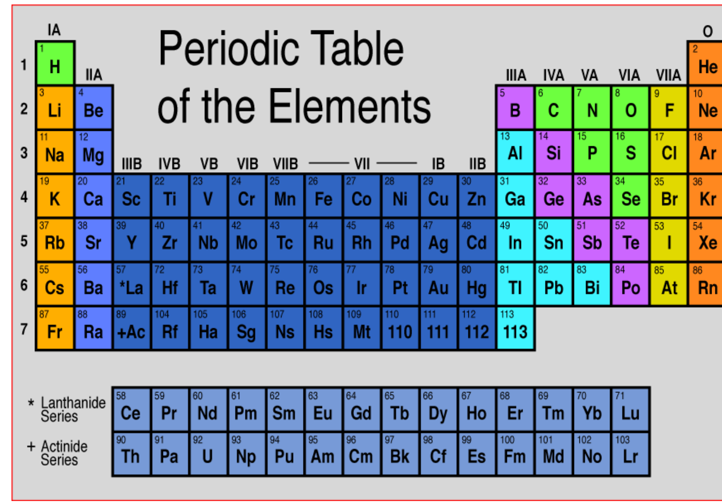
Mendeleev developed his table before scientists knew about the structure of atoms. He did not know that the atoms of each element contain a unique number of protons.

A British physicist, Henry Moseley, determined an atomic number for each known element.

In the modern periodic table, elements are arranged in order of increasing atomic number.

The Periodic Law

The elements within a column or group in the periodic table have similar properties.



The properties of the elements within a period change as you move across a period from left to right.

The pattern of properties within a period repeats as you move from one period to the next.

The Periodic Law

Periodic Law – when elements are arranged in order of increasing atomic number, there is a periodic repetition of their physical and chemical properties.

Group 1 – (alkali metals) are all highly reactive and are rarely found in elemental form in nature

Group 2 – (alkaline earth metals) are silvery colored, soft metals

Group 17- (halogens) the only group which contains elements in all three familiar states of matter at standard temperature and pressure.

Periodic Table of Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII	VII	VII	IB	IB	IIIA	IVA	VA	VIA	VIIA	0
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 106	107 107	108 108	109 109	110 110								

* Lanthanide Series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
+ Actinide Series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Legend - click to find out more...

H - gas | Li - solid | Br - liquid | Tc - synthetic

Non-Metals | Transition Metals | Rare Earth Metals | Halogens

Alkali Metals | Alkali Earth Metals | Other Metals | Inert Elements

Metal, Nonmetals, and Metalloids

The International Union of Pure and Applied Chemistry (IUPAC) set the standard for labeling groups in the periodic table.

They numbered the groups from left to right 1 – 18,

The elements can be grouped into three broad classes based on their general properties.

- Metals
- Nonmetals
- Metalloids

Across the period, the properties of elements become less metallic and more nonmetallic.

Metals

About 80 % of the elements are metals.

Properties of Metals

- Good conductors of heat and electric current.
- Have a high luster or sheen caused by the ability to reflect light
- Solids at room temperature (except Hg)
- Many metals are ductile (can be drawn into wires)
- Most metals are malleable (they can be hammered into thin sheets without breaking)

Nonmetals

Nonmetals are in the upper-right corner of the periodic table.

There is a greater variation in physical properties among nonmetal than among metals.

Properties of Nonmetals

- Most are gases at room temperature. S and P are solids, Br is a liquid.
- Nonmetals tend to have properties that are opposite to those of metals.
- In general, nonmetals are poor conductors of heat and electric current. Solid nonmetals tend to be brittle.

Metalloids

There is a heavy stair-step lines that separates the metals from the nonmetals.

Most of the elements that border this line are metalloids.

<u>Properties of Metalloids</u>

- Generally has properties that are similar to metals and nonmetals.
- Under some conditions they behave like a metal. Under other conditions they behave like a nonmetal.

Questions

How did chemists begin the process of organizing elements?

Used the properties of elements to sort them into groups.

What property did Mendeleev use to organize his periodic table?

In order of increasing atomic mass

How are elements arranged in the modern periodic table?

In order of increasing atomic number

Name the three broad classes of elements.

Metals, nonmetals, and metalloids

Questions

Name two elements that have properties similar to those of the element sodium

Li (lithium), K (potassium), Cs (cesium), Rb (rubidium), Fr (francium)

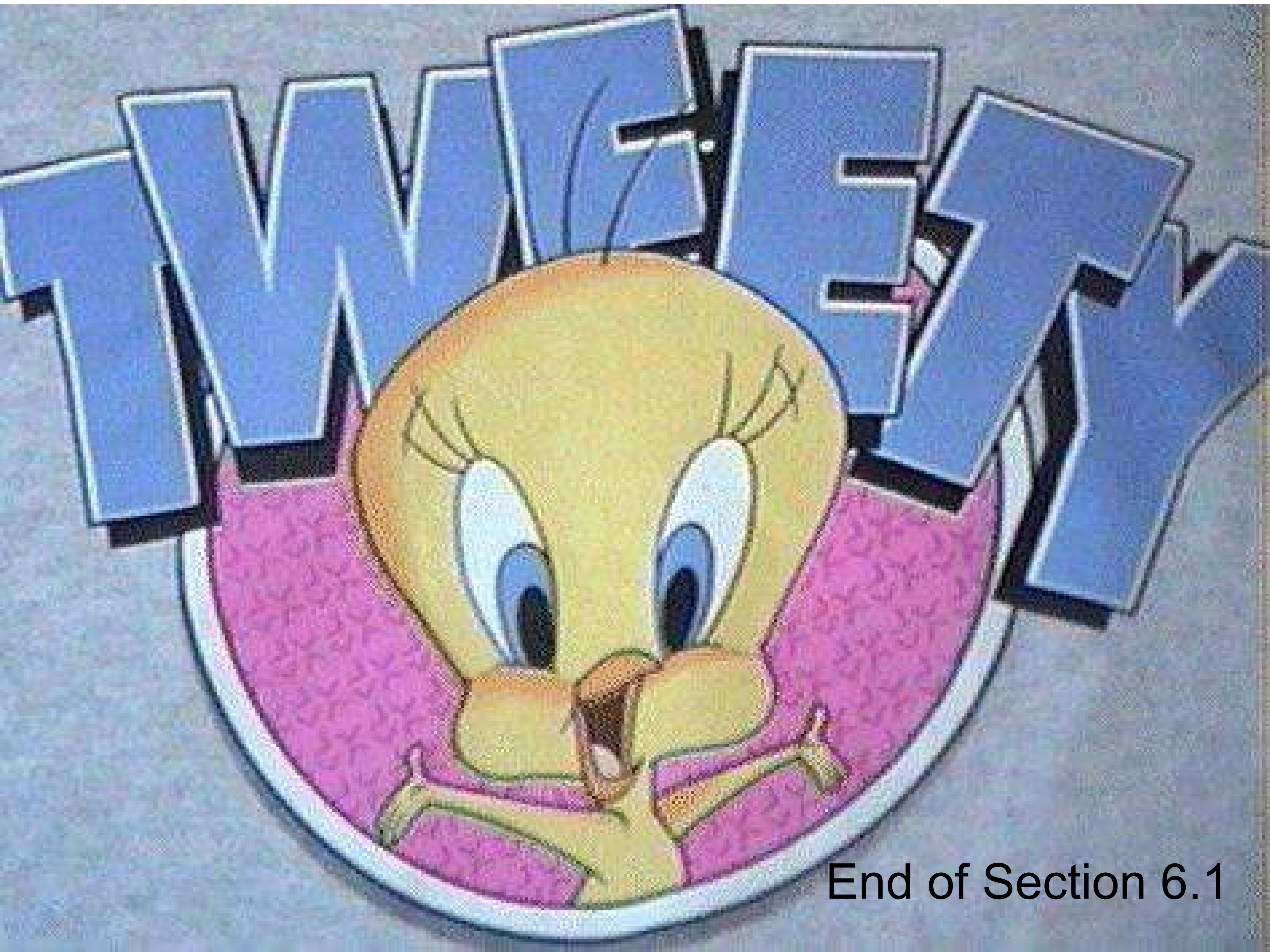
Identify each element as a metal, metalloid or nonmetal.

Gold (Au)
metal
Silicon (Si)
metalloid
Sulfur (S)
Nonmetal
Barium (Ba)
metal

Periodic Table of the Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII			IB	IIB	IIIA	IVA	VA	VIA	VIIA	O
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 Sg	107 Ns	108 Hs	109 Mt	110 110	111 111	112 112	113 113					

* Lanthanide Series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
+ Actinide Series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr



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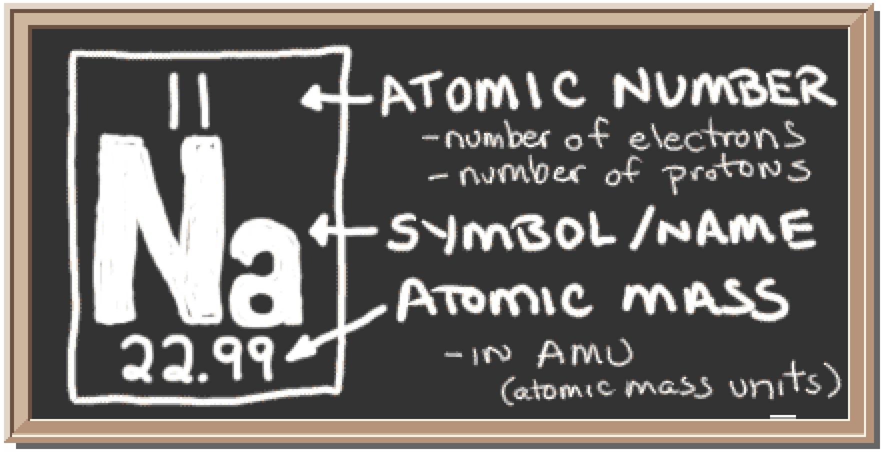
Squares in the Periodic Table

The periodic table displays the symbols and names of the elements along with information about the structure of their atoms.

The symbol for the element is located in the center of the square.

The atomic number is above the symbol.

The element name and average atomic mass are below the symbol.



Squares in the Periodic Table

The background colors in the squares are used to distinguish groups of elements.

Group I elements are called **alkali metals**. Group 2 elements are called **alkaline earth metals**.

The nonmetals of Group 17 are called **halogens.**

Group 18 elements are called **Noble Gases**

Groups 3–12 are called **transition metals**

The two periods usually located at the bottom of the periodic table separate from the main table are called **inner transition elements.** Period 8 is called the **Lanthanide Series** and Period 9 is called the **Actinide Series**

Electron Configuration in Groups

Electrons play a key role in determining the properties of elements.

So there is a connection between an element's electron configuration and its location in the periodic table.

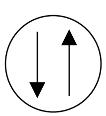
Elements can be sorted into noble gases, representative elements, transition metals, or inner transition metals based on their electron configurations.

The Noble Gases are in Group 18 and are sometimes called inert gases because they rarely take part in a reaction.

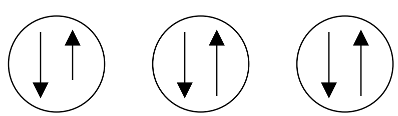
Electron Configuration in Groups

Helium (He)	$1s^2$
Neon (Ne)	$1s^2 2s^2 2p^6$
Argon (Ar)	$1s^2 2s^2 2p^6 3s^2 3p^6$
Krypton (Kr)	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

The highest occupied energy level for each element, (the s & p sublevels) are completely filled with electrons.



s sublevel

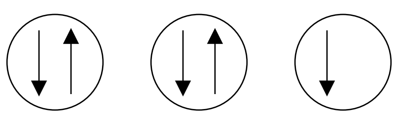


p sublevel

Electron Configuration in Groups

Fluorine (F)	$1s^2 2s^2 2p^5$
Clorine (Cl)	$1s^2 2s^2 2p^6 3s^2 3p^5$
Bromine (Br)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$
Iodine (I)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$

The highest occupied energy level for each element, (the p sublevels) are <u>filled with electrons 5 electrons</u>.



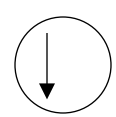
p sublevel

The Representative Elements

Elements in groups 1, 2 and 13 through 17 are often referred to as **representative elements** because they display a wide range of physical and chemical properties.

In atoms of representative elements, the s and p sublevels of the highest occupied energy level are not filled.

Lithium(L)	$1s^2 2s^1$
Sodium (Na)	$1s^2 2s^2 2p^6 3s^1$
Potassium (K)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$



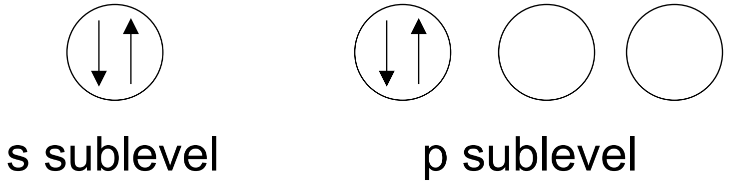
s sublevel

The Representative Elements

Carbon (C)	$1s^2 2s^2 2p^2$
Silicon (Si)	$1s^2 2s^2 2p^6 3s^2 3p^2$
Germanium (Ge)	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$

In atoms of carbon, silicon, and germanium, in Group 14, there are four electrons in the highest occupied energy level

For any representative elements, its group number equals the number of electrons in the highest occupied energy level.



Transition Metals

Elements in groups 3-12 are referred to as **transition elements.**

There are two types of transitions elements: **transition metals** and **inner transition metals**

In atoms of a transition metal, the highest occupied s sublevel and a nearby d sublevel contain electrons.

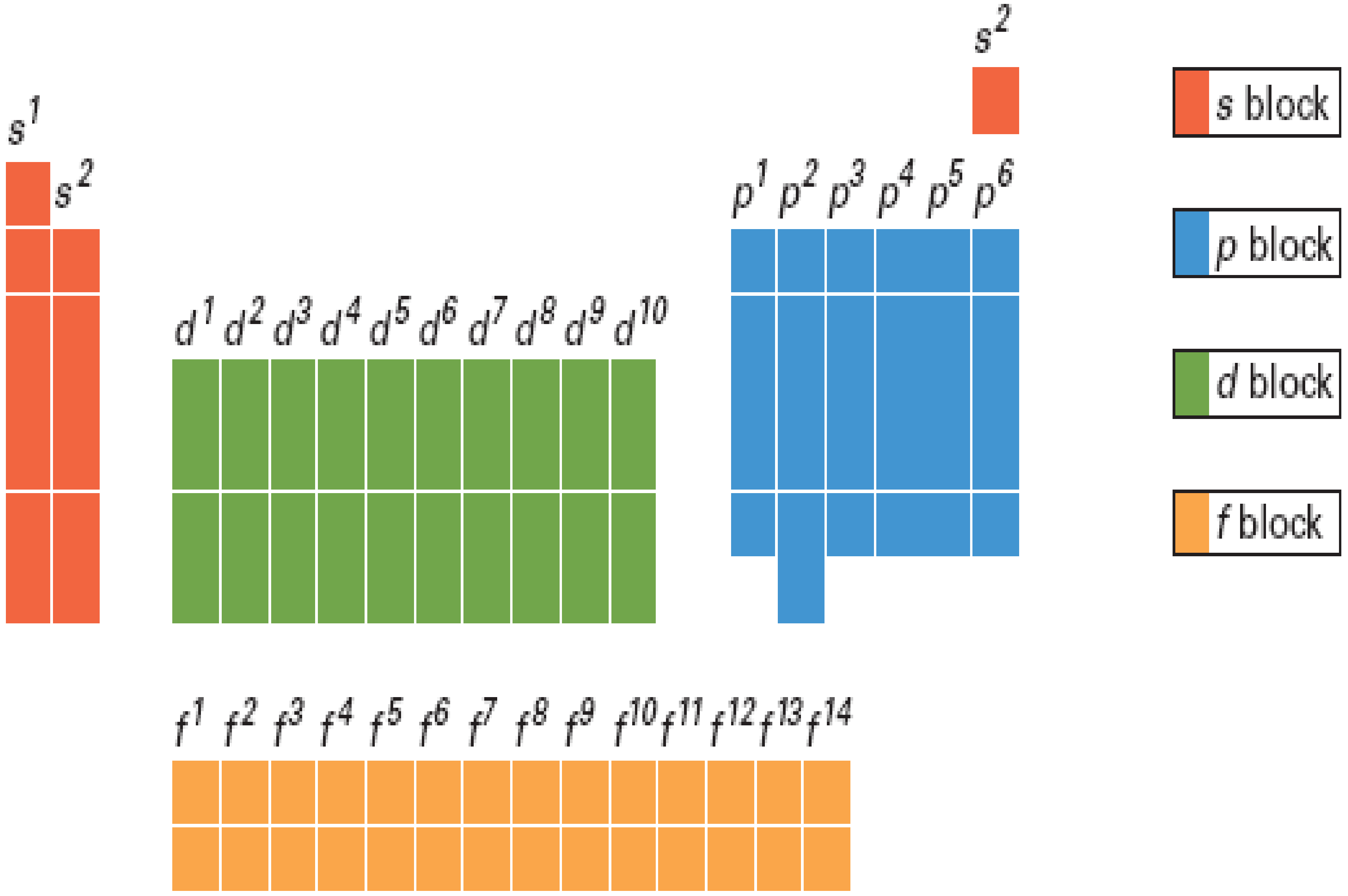
These elements are characterized by the presence of electrons in d orbitals.

Inner Transition Metals

The inner transition metals appear below the main body of the periodic table. _

In atoms of an inner transition metal, the highest occupied s sublevel and a nearby f sublevel generally contain electrons.

The inner transition metals are characterized by f orbitals that contain electrons.



s block
p block
d block
f block

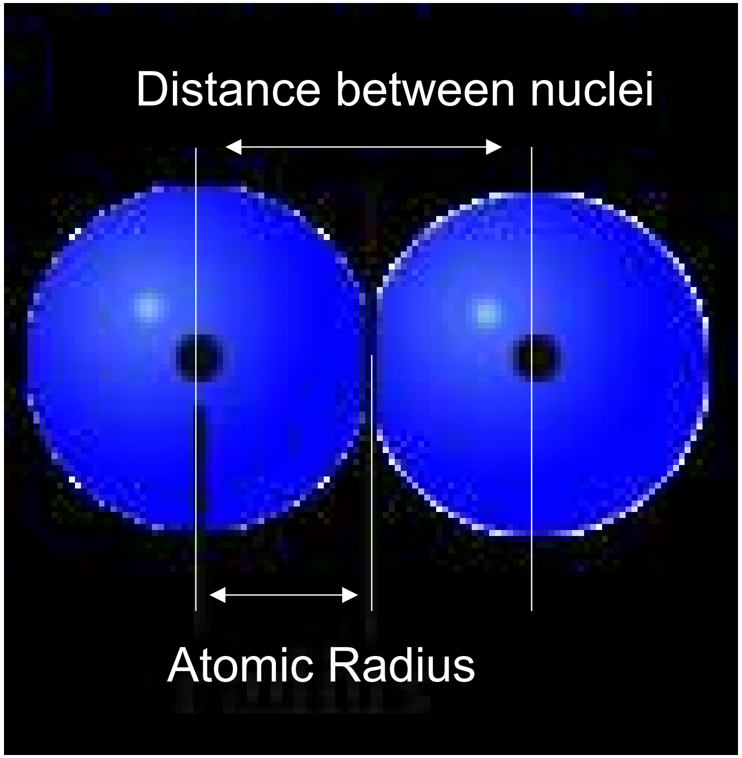
End of Section 6.2

Periodic Trends – Atomic Size

When atoms of the same element are attached to one another they are called **molecules**.

Because the atoms in each molecule are identical, the distance between the nuclei of these atoms can be used to estimate the size of the atoms.

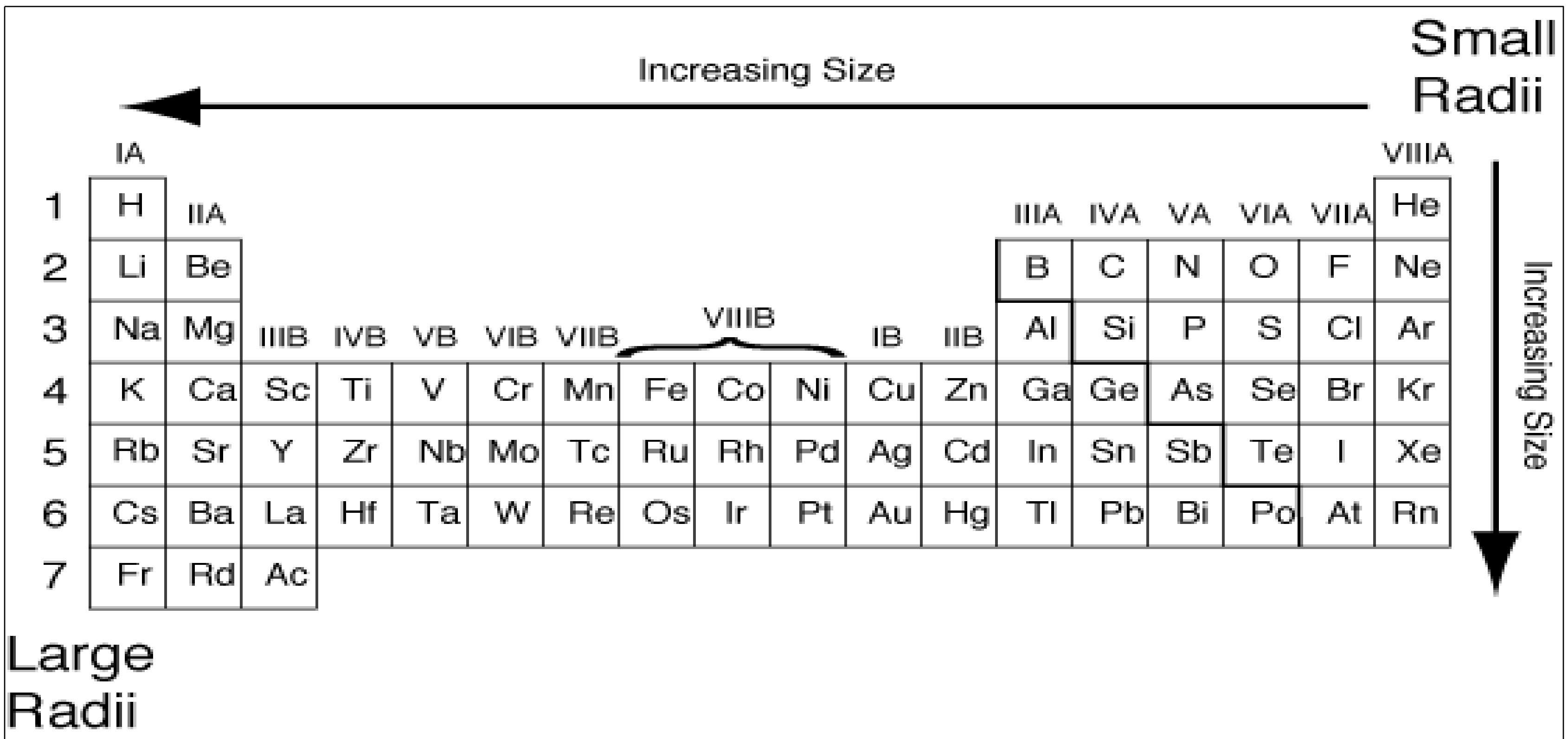
The **atomic radius** is one half of the distance between the nuclei of two atoms of the same element when the atoms are joined.



Atomic Size

The distance between atoms in a molecule are extremely small, so it is often measured in picometers. (10^{12} pm = 1m)

In general, atomic size increases from top to bottom within a group and decreases from left to right across a period.



Atomic Size

As the atomic number increases within a group, the charge on the nucleus increases and the number of occupied energy levels increases.

The increase in positive charge draws electrons closer to the nucleus.

The increase in the number of occupied orbitals shields electrons in the highest occupied energy level from the attraction of protons in the nucleus.

The shielding effect is greater than the effect of the increase in nuclear charge, so the atomic size increases.

Atomic Size

In general, atomic size decreases across a period from left to right.

Each element has one more proton and more more electron than the preceding element.

The increasing nuclear charge pulls the electrons in the highest occupied energy level closer to the nucleus and

Ions

Some compounds are composed of particles called ions. An **ion** is an atoms or group of atoms that has a positive or negative charge.

An atom is electrically neutral because it has equal numbers of protons and electrons.

Positive and negative ions from when electrons are transferred between atoms.

Atoms of metallic elements tend to form ions by losing one or more electrons from their highest occupied energy levels.

A sodium atom tend to lose one electron.

Cations

In the sodium ion, the number of electrons (10) is no longer equal to the number of protons (11).

Because there is more positively charged protons than negatively charged electrons, the sodium ion has a net positive charge.

An ion with a positive charge is called a **cation**.

The charge for a cation is written as a number followed by a plus sign. (Example: 1^{+})

If the charge is 1^{+}, the number 1 is usually omitted from the complete symbol for the ions. (Na^{+})

Anions

Atoms of nonmetallic elements, such as chlorine, tend to form ions by gaining one or more electrons.

A chlorine atom tend to gain one electron.

In a chlorine ion, the number of electrons (18) is no longer equal to the number of protons (17).

Because there are more negatively charged electrons than positively charged protons, the chloride ion has a net negative charge.

An ion with a negative charge is called an **anion**.

Examples: Cl^{-}, S^{2-}

Trends in Ionization Energy

Recall that electrons can move to higher energy levels when atoms absorb energy.

Sometimes there is enough energy to overcome the attraction of the protons in the nucleus.

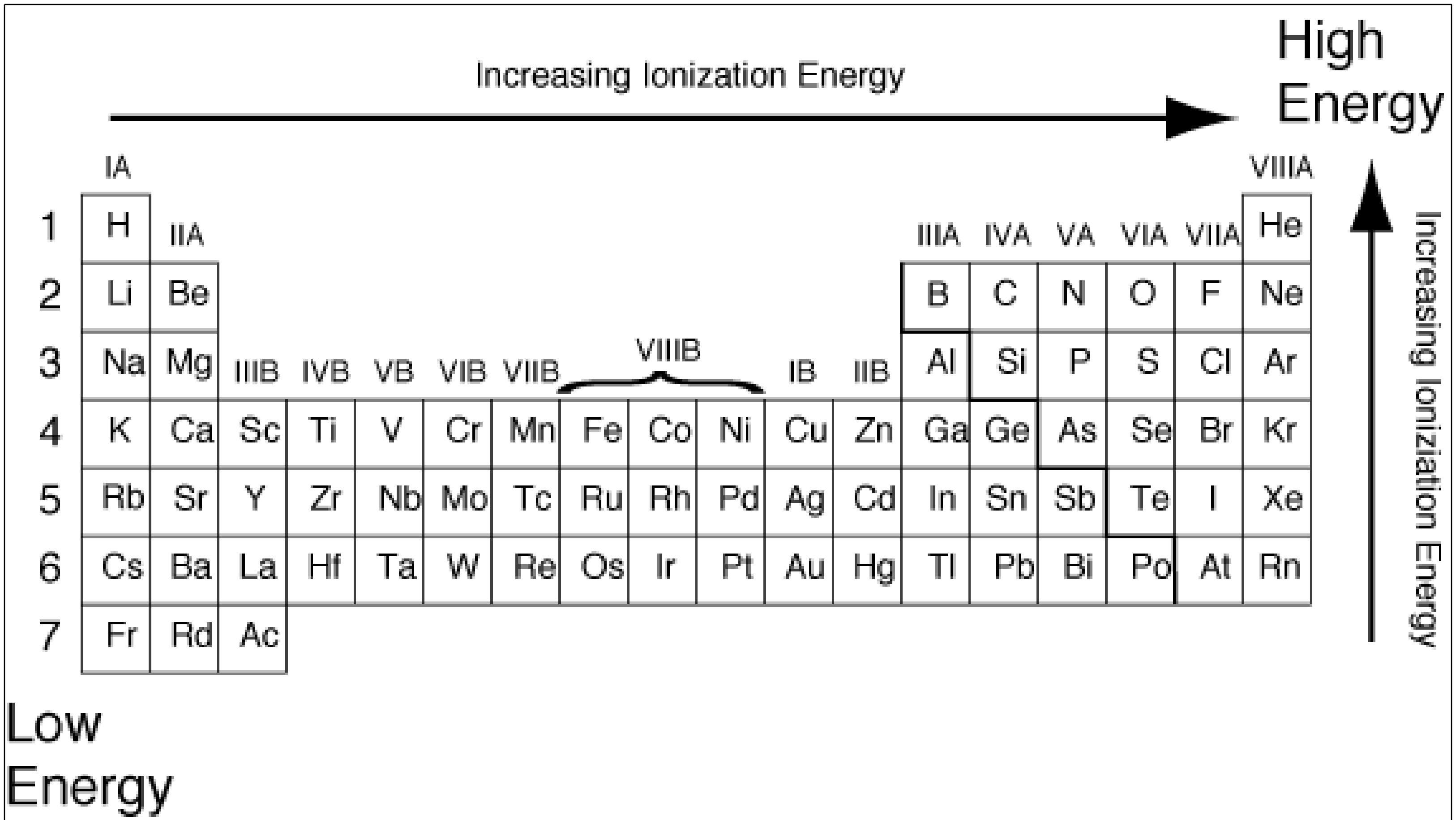
The energy required to remove an electron from an atom is called **ionization energy**.

The energy to remove the first electron from an atom is called the **first ionization energy**.

The cation produced has a 1^+ charge.

Trends in Ionization Energy

First ionization energy tends to decrease from top to bottom within a group and increase from left to right across a period.



Ionization Energy

The energy to remove the first electron from an atom is called the **first ionization energy**. The cation produced has a 1^+ charge.

The **second ionization energy** is the energy required to remove an electron from an ion with a 1^+ charge. The ion produced has a 2^+ charge.

The **third ionization energy** is the energy required to remove an electron from an ion with a 2^+ charge. The ion produced has a 3^+ charge.

Ionization Energy

Ionization energy can help you predict what ions elements will form.

If you look at Li, Na, & K ionization energies, the increase in energy between the first and second ionization energies is large.

It is relatively easy to remove one electron from a Group I metal atom, but it is difficult to remove a second electron, so Group I metals tend to form ions with a 1^+ charge.

Symbol	First IE (kJ/mol)	Second IE (kJ/mol)
Li	520	7297
Na	496	4565
K	419	3069

Group Trends in Ionization Energy

In general, first ionization energy decreases from top to bottom within a group. (recall that the atomic size increases as the atomic number increases within a group)

As the size of the atom increases, nuclear charge has a smaller effect on the electrons in the highest occupied energy level.

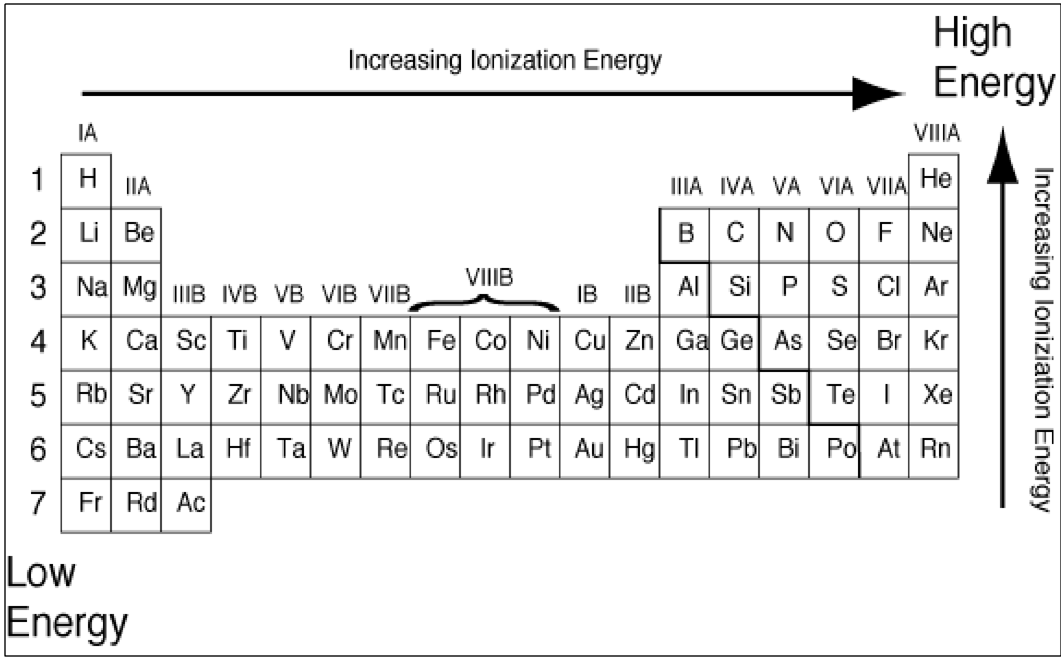
So less energy is required to remove an electron from this energy level and the first ionization energy is lower.

Group Trends in Ionization Energy

In general, the first ionization energy of representative elements tends to increase from left to right across a period.

This trend can be explained by the nuclear charge, which increases, and the shielding effect, which remains constant.

So there is an increase in the attraction of the nucleus for an electron, thus it takes more energy to remove an electron from an atom.



Trends in Ionic Size

During reactions between metals and nonmetals, metal atoms tend to lose electrons and nonmetal atoms tend to gain electrons.

The transfer has a predictable affect on the size of the ions that form.

Cations are always smaller than the atoms from which they form. Anions are always larger than the atoms from which they form.

When a Na atom loses an electron, the attraction between the remaining electrons and the nucleus is increased.

The electrons are drawn closer to the nucleus.

Trends in Ionic Size

Metals that are representative elements tend to lose all their outermost electrons during ionization, so the ion has one fewer occupied energy level.

The trend is the opposite for nonmetals like the halogens in Group 17.

For each of these elements, the ion is much larger than the atom.

As the number of electrons increases, the attraction of the nucleus for any one electron decreases

Trends in Ionic Size

The effective nuclear charge experienced by an electron in the highest occupied orbital of an atom or ion is equal to the total nuclear charge (the number of protons) minus the shielding effect due to electrons in lower energy levels.

The effective nuclear charge determines the atomic and ionic radii.

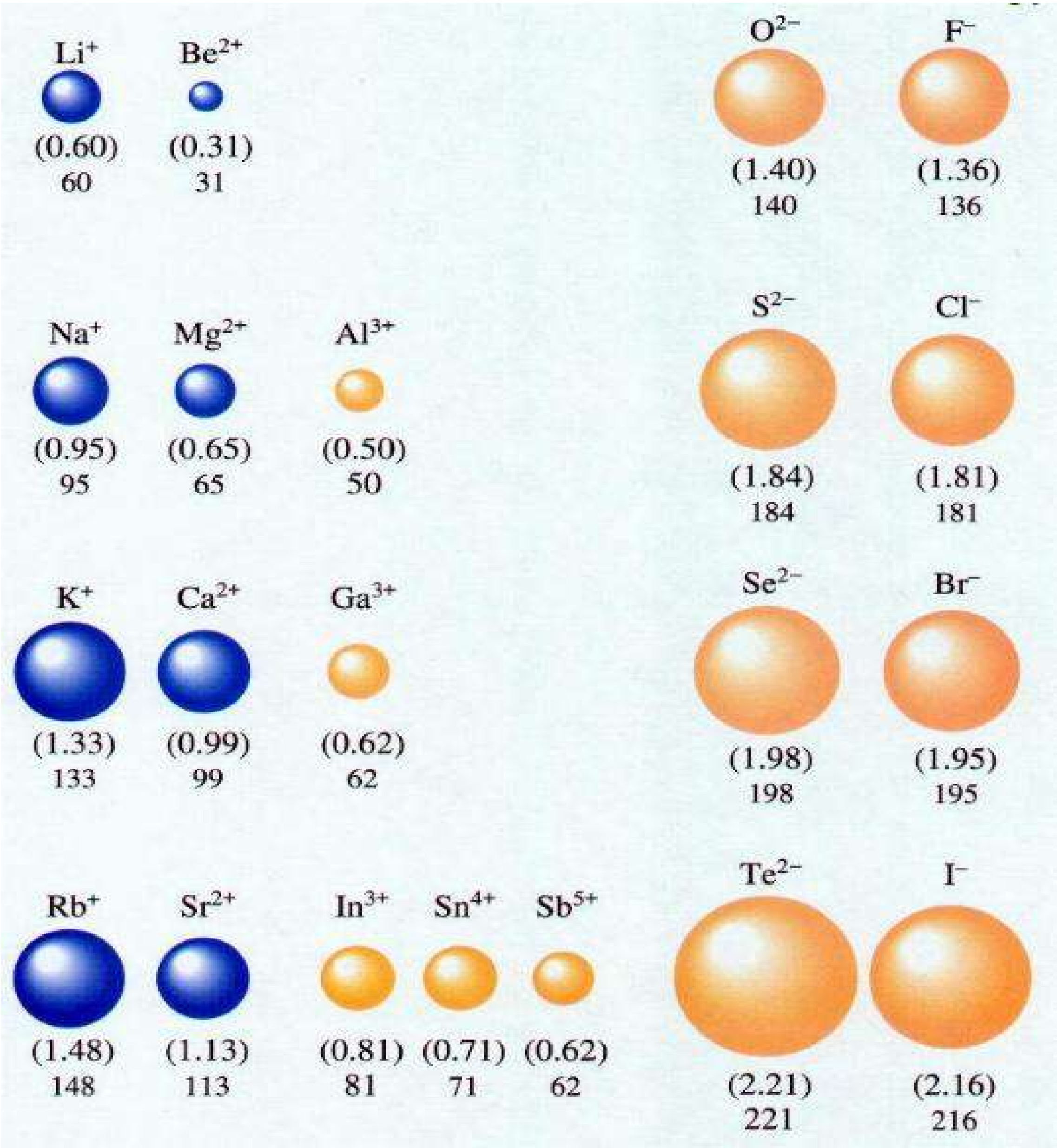
Left to right in any period, the principal quantum number, n, of the highest occupied energy level remains constant, but the effective nuclear charge increases.

Therefore, atomic and ionic radii decrease as you move to the right in a period.

Trends in Ionic Size

Within any group, as you proceed from top to bottom, the effective nuclear charge remains nearly constant, but the principal quantum number increases.

Consequently, atomic and ionic radii increase from top to bottom within a group.



Trends in Electronegativity

There is a property that can be used to predict the type of bond that will form during a reaction.

This property is **electronegativity**, which is the ability of an atom of an element to attract electrons when the atom is in a compound.

In general, electronegativity values decrease from top to bottom within a group.

For representative elements, the values tend to increase from left to right across a period.

Trends in Electronegativity

Metals at the far left of the periodic table have low values. Nonmetals at the far right (excluding noble gases) have high values.

The electronegativity value among the transition metals are not as regular.

The lease electronegative element is cesium. It has the least tendency to attract electrons. When it reacts, it tends to lose electrons and form positive ions.

The most electronegative element is fluorine, and when it is bonded to any other element it either attracts the shared electrons or forms a negative ion.

Trends in Electronegativity

Metals at the far left of the periodic table have low values. Nonmetals at the far right (excluding noble gases) have high values.

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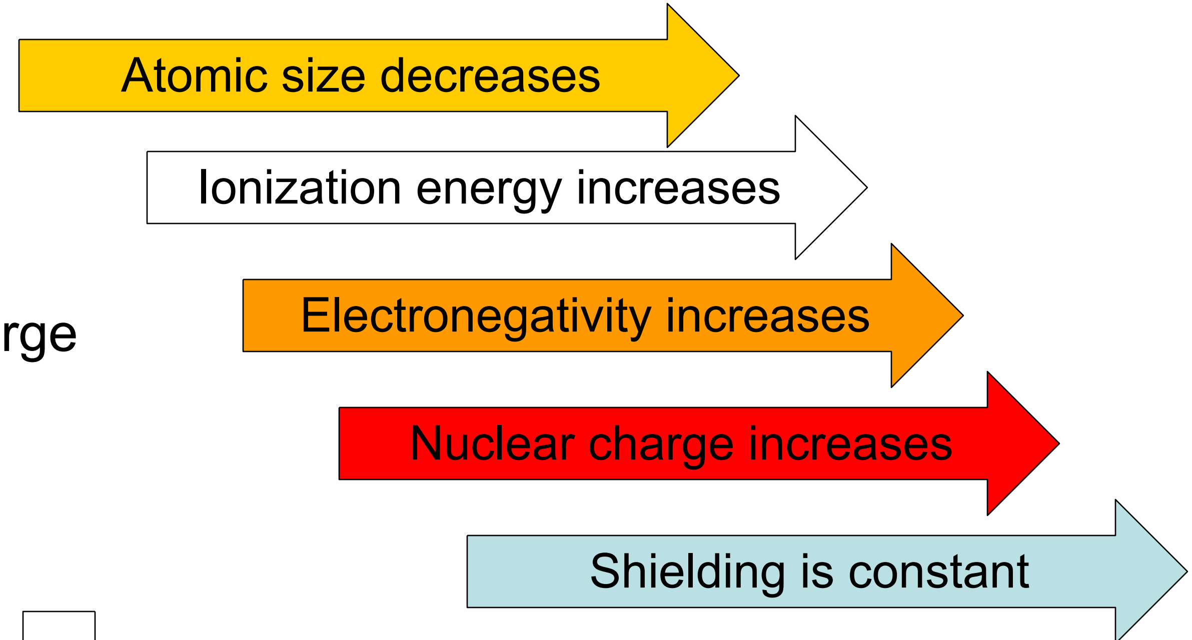
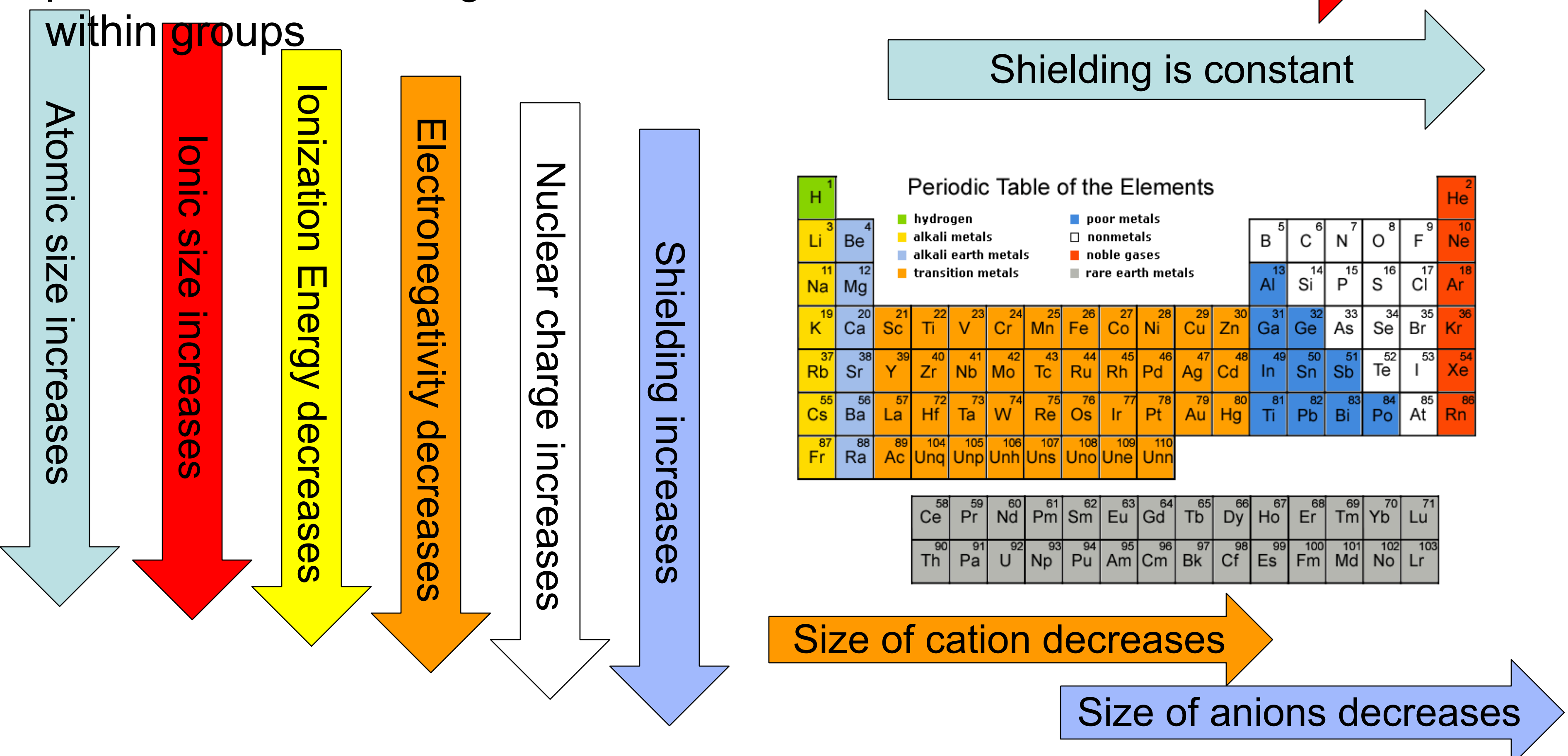
Trends for Groups 1A Through 8A

- Can be explained by variations in atomic structure
- Increase in nuclear charge within groups & across periods, also shielding within groups





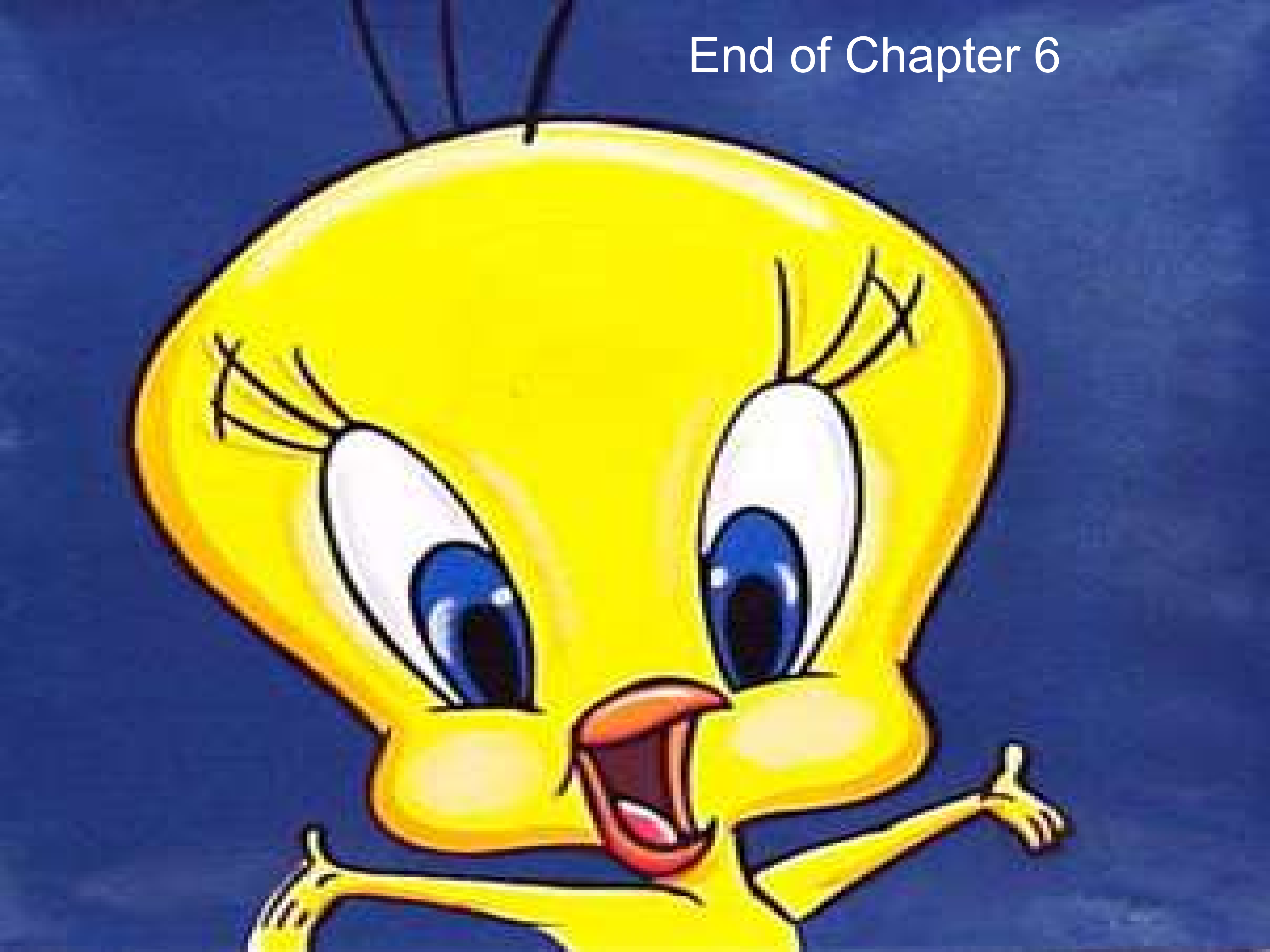
Periodic Table Trends

Metals at the far left of the periodic table have low values. Nonmetals at the far right (excluding noble gases) have high values.

The electronegativity value among the transition metals are not as regular.

The element with the lowest electronegativity value is cesium. It has the least tendency to attract electrons. When it reacts, it tends to lose electrons and form positive ions.

The most electronegative element is fluorine, and when it is bonded to any other element it either attracts the shared electrons or forms a negative ion.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