



#### Periodic TableHistory

Mankind began discovering the properties and uses of elements thousands of years ago.

Over 350 years ago, scientists began to discover, list, classify and organize the elements of our construction.





#### Periodic TableHistory

- In 1869 Dmitri Mendeleev organized the first periodic table by atomic mass and properties. His design allowed room for yet-tobe discovered elements and accounts for the "modern" periodic table.
- In about 1916, Henry Moseley, following the discovery of the nucleus, organized elements by ATOMIC NUMBER. In his research into x-ray wavelengths of elements, he found that nuclear charge correlated with the wavelength produced by each element. This gave significance to the atomic number which had previously been thought of as arbitrary.







The modern periodic table ordered elements by **ATOMIC** NUMBER.





# Periodic Table Organizati

### **GROUPS**

- VERTICAL columns
- Designated by numbers 1-18
- Group # ONE's place represents the number of **valence** electrons an atom possesses.
- This applies to groups 1, 2, 13-18 only.
- Called the Representative Elements
- Ex: all elements in group 16 have 6 valence e





### Periodic Table Organization

#### **GROUPS**

- Valence electrons are electrons in the OUTERMOST energy level of an atom.
- They are responsible for **bonding**.
- Down each level in a group, an energy level is added to the atom.
- · Each outer energy level is repelled by levels of core electrons beneath.
- Elements in a group have similar physical and chemical properties.



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#### Periodic Table Organization

#### **PERIODS**

- HORIZONTAL rows
- Designated by numbers 1-7
- Represents the number of energy levels
- Energy Level the space around the atoms's nucleus where electrons are confined based on the amount of energy they have.





#### **PERIODS**

- From the LEFT to the RIGHT in a period, each element gains 1 proton and 1 electron.
- Properties change from left to right



 Zig-Zag Line - separates metals (on the left) from nonmetals (on the right).



## Periodic Table Organization

#### **METALS**

Metals - tend to lose electrons (e') to form cations.

Cation - atom that has lost one or more electrons. Has a (+) charge.







Ex. Sodium (Na) loses 1 electron to become the sodium cation (Na $^{+1}$ )



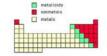
### Periodic Table Organization

#### **NONMETALS**

**Nonmetals** - tend to **gain** electrons (e-) to form **anions**.

Anion - atom that has **gained** one or more electrons. **Has a (-) charge.** 





Ex. Sulfur tends to gain 2 electrons to become the sulfide anion (S $^{\!-2}\!)$ 



### Periodic Table Organization

#### METALLOIDS

**Metalloids** - have properties of metals and nonmetals, aka semimetals



- Found along and to either side of the zig-zag line (except for AI)
- Mostly brittle solids
- Metalloids are **semiconductors** which means they conduct electricity well, but not heat.
- Uses: computer chips (silicon) and eyewash (boron).



### Periodic Table Classification

### Representative Elements (Groups 1, 2, 13-18)

- Alkali Metals (Group 1) n s1
- 1 valence electron
- will lose 1 e to form the +1 cation
- very **reactive**; never found "**free**" (not bound) in nature
- stored under oil to prevent oxidation





### Periodic Table Classification

#### **Representative Elements**

- . Alkali metals (continued)
- Chemical property: most reactive, reacts with H<sub>2</sub>O
- Physical property: soft, silvery metal
- Uses: soap, lights, batteries, nuclear reactors, table salt, medicines, electrolytes in living things



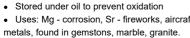




## **Periodic Table Classification**

#### **Representative Elements**

- Alkaline Earth Metals (Group 2) s²
- 2 valence electrons
- will give up 2 e to form +2 ion
- Reactive; never found "free" in nature
- Harder, more dense than alkali metals
- Chemical prop.- less reactive than group 1
- Stored under oil to prevent oxidation
- Uses: Mg corrosion, Sr fireworks, aircraft



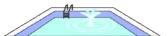




### **Periodic Table Classification**

### **Representative Elements**

- Halogens (Group 17) s<sup>2</sup>p<sup>5</sup>
- 7 valence electron
- gains 1 e- to form -1 ion
- Most reactive of non metals
- Uses: CI pool purification, I iodized salt, F, dental health







#### **Representative Elements**

- Noble Gases (group 18) s²p6
- 8 valence e- (except He)
- Full outer energy level
- Not reactive
- Uses: He balloons, Ne signs, Ar windows









#### Periodic Table Classification

#### **Representative Elements**

- Groups 13 16 (s<sup>2</sup>p<sup>1</sup> s<sup>2</sup>p<sup>4</sup>)
- Named for the first element in each group.
- Ex: Group 13 is called the "Boron" group.
- Properties vary depending on group.
- Metals in these groups have variable properties.
- Nonmetals in these groups have variable properties.



- The commonly known "metallic" elements
- Multiple **oxidation** states, which means they can lose a **variable** number of electrons. (except Ag, Zn, Cd!)
- Many are **brightly** colored and are found free in nature
- Uses: Fe steel, Au/Ag jewelry





Lanthanides (Z = 57-70)

(a.k.a. - rare earth metals)

- (La to Yb) are **silvery**, high melting point, very valuable metals
- Uses: glass industry, modern electronic devices





#### **Inner Transition Elements**

#### **Actinides (Z = 89-102)**

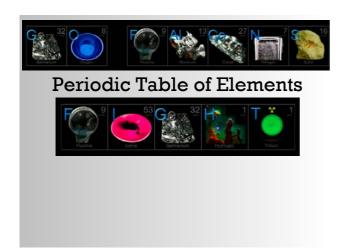
- (Ac to No) are radioactive
- Uses: nuclear fuel, nuclear medicine, research, many industries



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Please take detailed notes.

Be ready to ask questions in class.





#### **Oxidation States**

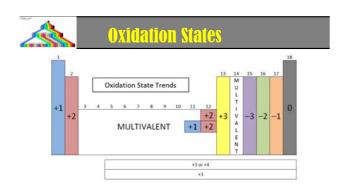
- The Octet Rule: atoms gain or lose electrons to secure 8 electrons in their valence (outer most) shell (energy level)
- Duet Rule: applies to H, He
- Man-made rule! Not law of nature.
- Ion an atom (or group of combined atoms) that has a charge because of the gain or loss of electrons. Ex: Ca<sup>+2</sup>, F<sup>-1</sup>



#### **Avidation States**

- Oxidation State the charge on an ion or an element AFTER it gains or loses electrons to fulfill the Octet Rule. (can be + or -)
- Oxidation state is determined by the # of valence electrons in an atom.
- The **Representative** elements (group 1, 2, 13-18) have *predictable* oxidation states.

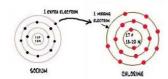
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#### **Ionic Ronding**

• lonic bonds are formed between a CATION (metal) and an ANION (nonmetal) when electrons are **lost and gained** due to forces between the elements

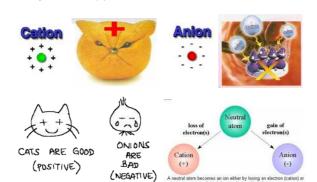


http://www.tubechop.com/watch/1521483



# **Ionic Bonding**

• CATIONS lose electron(s) and ANIONS gain electron(s)





#### Ionic Bonding

- Use the oxidation state map to determine charges.
- To form an ionic compound, we must follow a few easy steps...

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#### Ionic Bonding

- 1. Determine the charge (oxidation state) of the metal and nonmetal
- Write the cation with its charge and the anion with its charge
  Balance the demands of each element (SWAP & DROP)

Exa	mple	Prac	tice	Pra	ctice
Na	CI	Ca	S	Ва	CI



### Covalent (Molecular) Bonding

Covalent bonds form when **nonmetals share** electrons. No ions form. Attraction between atoms arises from the **electronegativity difference** between the elements bonded.



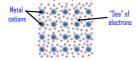
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 b) Distribution of partial charges in a water In water, each hydrogen shares its electron with oxygen. Each atom then fills it has a "filled" valence shell.



#### Metallic Ronding

Metals are made entirely of periodic table metals.



- Metal atoms are packed closely together.
- Their valence electrons are delocalized and travel freely throughout the metal sample. They do not belong to any atom!
   This forms a "sea of electrons"



# **Metallic Bonding**

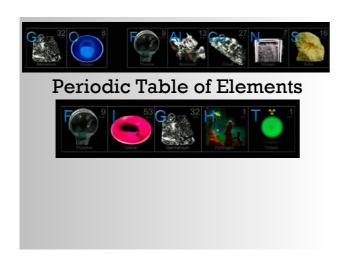
- The metal *cations* are **attracted** to the negatively charged, electron rich region between atoms.
- The electrons insulate the cations from each other.
- Microscopically, the atoms forms a metal crystal.
   Crystal any substance whose particles are arranged in a 3-D orderly pattern.
- Metal alloys, like brass or steel, are made from varying amounts of pure metals mixed together while molten, then allowed to cool. The mixture forms the crystal as it cools.
   unique alloys give them special properties. Examples include stainless steel, structural steel, sterling silver, bronze, and most other metals you can imagine.

End of Video

Please take detailed notes.

Be ready to ask questions in class.

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### PERIODIC TRENI

There are numerous "trends" we can study on the periodic table, but most are governed by two electrical phenomena.

- Z<sub>effective</sub> nuclear charge
   Shielding



# **Z**<sub>effective</sub> Nuclear Charge

- Z<sub>effective</sub> nuclear charge is the overall force the protons in a nucleus exert on the surrounding electrons. In an energy level, electrons are attracted to the nucleus equally, regardless of how many electrons are there.
  - It is influenced by:
- Z (atomic number)
- Electron repulsion (electrons repel one another)

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#### SHIFLDING

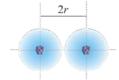
- Shielding is the addition of entire energy levels between the protons in the nucleus and the valence shell of electrons.
  - Regardless of nuclear charge, entire energy levels worth of electrons greatly increase the distance between nucleus and valence shell, reducing attraction.
  - > Electrons repel, further increasing distance



#### Atomic Radius - TRENI

Atomic radius is the *radius* of an atom.  $\bf AR$  is measured in picometers (1x10<sup>-12</sup> m) or Ångstroms. 1 Å= 100 pm.

Measured indirectly from center to center as we only see electron clouds, we cannot "see" the radius.





#### Atomic Radius - TREND

Period Trend (left to right across a period):

- $\bullet~$  As atomic # increases, AR decreases (increasing  $Z_{\mbox{\scriptsize eff}})$
- For each element across a period, one p\* is added to the nucleus and one e\* is added to the valence energy level.

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#### Atomic Radius - TREND

#### Periodic Trend:

 Nuclear Charge increases (each proton makes the nucleus more ++), electrons are pulled INWARD (opposites attract).

Practice: K / Kr has a larger atomic radius

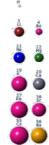




#### Atomic Radius - TREN

#### **Group** Trend (top to bottom):

- As atomic # INCREASES down a group, AR INCREASES
- For each element *down a group*, one ENERGY LEVEL is added to the atom.
- Nuclear charge increases, but **SHIELDING** also increases.





#### Atomic Radius - TREND

#### Group Trend (down):

- SHIELDING as energy levels are added, electrons are less affected by pull of positively charged nucleus because distance is increased and electrons repel each other.
- Practice: Ba / Mg has the larger atomic radius.







XXXX	First Ionization Energy - TREND

---The **energy** required to remove the **outermost** electron from an atom **(kJ/mol)** 

Period Trend: (across a row)

- Generally, as atomic # INCREASES, IE increases <u>because</u> increasing nuclear charge holds electrons with greater force.
- Nuclear charge increases, **electrons** are held more **closely** to the nucleus

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	First Ionization Energy - TREND

Period Trend: (across a row)

• Practice: Ca / Br has as higher IE





Group Trend: (down a column)

- As atomic # INCREASES, IE decreases
- Shielding is the reason!
- Farther from nucleus, more **easily** removed.

Practice: Na / Cs has a higher IE



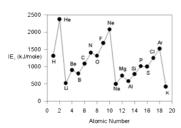
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Challenge: Which element (from entire PT) has the highest IE?



The graphical trend, two slight dips.

First Ionization Energy versus Atomic Number





#### Second I onization Energy - TRENI

The energy required to remove each electron from an atom increases successively. Removing a "core" electron requires significantly more energy than valence electrons.





### Electronegativity - TREND

- --- A measure of an atom's ability to **pull on another** atom's electrons. Abbreviated "EN". No units!
- Electronegativity difference in a bond of > 1.7 indicates ionic bonding
- Electronegativity difference of < 1.7 indicates (molecular) <u>covalent</u> bonding





## **Electronegativity - TREND**

Period Trend: (across a row - left to right)

- As atomic # INCREASES, EN increase.
- Nuclear charge is increasing.





Period Trend: (across a row)

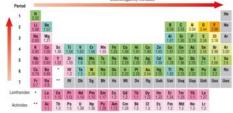
• Practice: Be / O has a higher EN



- Group Trend: (Down a column)

   As atomic # INCREASES, EN decreases
- SHIELDING
- Elements at **top** of a group have e- closely held by nucleus because they are closest to nucleus.
- Elements at the bottom of the group have more distance between nucleus and outer e-
- Practice: N / Sb has a higher EN





Challenge: Why are noble gases not listed?



#### **Reactivity - TRENI**

- ---- Refers to how **likely** an atom is to react with other substances, depends on...
- how easily electrons can be removed (ionization energy)
- how easily atoms grab electrons (electronegativity)

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#### **Reactivity - TREND**

#### **Metals (low ionization energy):**

<u>Period Trend</u> (across a row) - As atomic # INCREASES, reactivity **decreases**.

<u>Group Trend</u> (down a column) - As atomic # INCREASES reactivity **increases**.

<u>WHY?</u> If the IE is low, an atom is more likely to lose an electron, which results in a chemical *reaction* with another atom.



#### Reactivity - TRENII

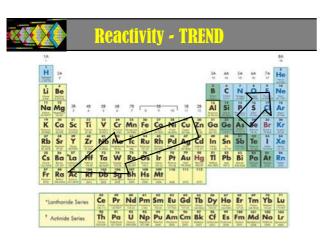
#### Nonmetals (high electronegativity):

<u>Period Trend</u> (across a row) - As atomic # INCREASES, reactivity **increase**.

<u>Group Trend</u> (down a column) - As atomic # INCREASES reactivity **decreases**.

 $\underline{\it WHY?}$  IF EN is high, the atom takes electrons from other atoms more easily, which results in a chemical reaction.

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Across a period (L to R), metallic character decreases.

Down a group, metallic character increases.



End of Video

Please take detailed notes.

Be ready to ask questions in class.



Density is odd. Across a period, it peaks in the middle of the d and f blocks. Down a group, density increases!





Generally, across a period, MP peaks in the middle of the d and f blocks, with metalloids showing high MP also (think applications!)

Down a group, MP has no appreciable trend!

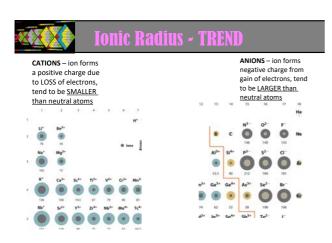




Generally, across a period, BP peaks in the middle of the d and f blocks, with metalloids showing high BP also (think applications!)

Down a group, BP has no appreciable trend!



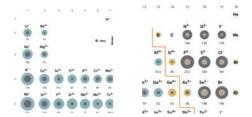




PERIOD TREND (left to right):

As atomic # INCREASES, IR of CATIONS and ANIONS DECREASES.

- In cations, loss of the valence electron shell electrons makes the atom smaller.
- In anions, the increasing nuclear charge pulls all electrons inward.





**GROUP TREND** (top to bottom)

As atomic # INCREASES down a group,  $\ensuremath{\mathbf{IR}}$  of CATIONS and ANIONS tends to  $\ensuremath{\mathbf{INCREASE}}$ 

• In both cases, shielding contributes to the trend.

