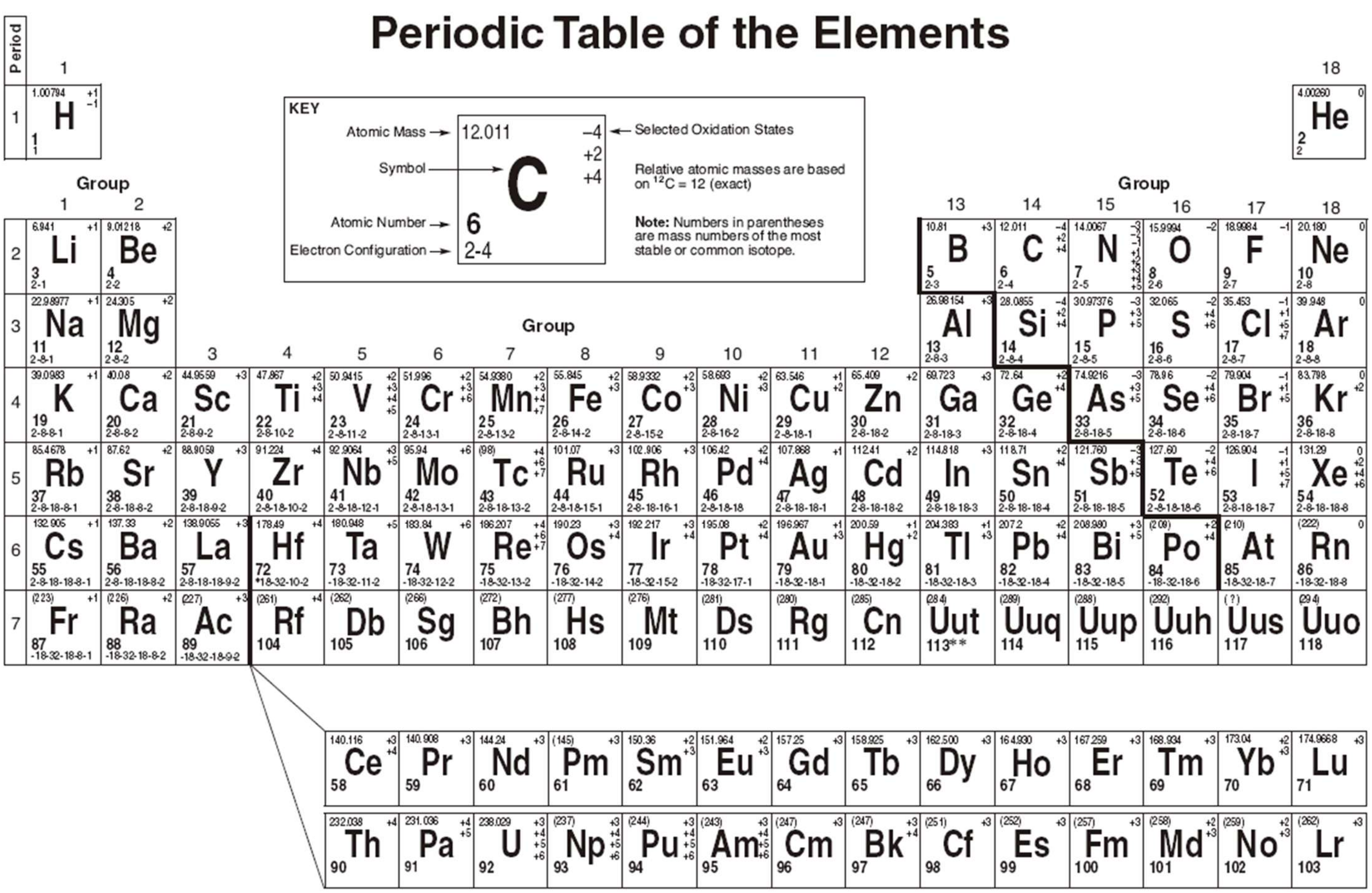
TRENDS OF THE PERIODIC TABLE

NAME



*denotes the presence of (2-8-) for elements 72 and above

**The systematic names and symbols for elements of atomic numbers 113 and above will be used until the approval of trivial names by IUPAC.

Source: *CRC Handbook of Chemistry and Physics*, 91st ed., 2010–2011, CRC Press

☺

Parts of the Periodic Table

There are two kinds of elements, metals and nonmetals. They have very different properties from each other. The metals are all on the left side of the "staircase" that starts at B and ends at At. The only exception to that is the strangest atom of them all, HYDROGEN, which is a nonmetal but sits at the top of group 1.

The table has 18 columns which we call groups. They are numbered across the top of the table. The rows go across the table, and they are numbered from 1 to 7. If you put your finger into Group 8 and Period 6, you touch Os (osmium) the densest of all atoms.

Some groups and periods have special names.

Group 1 (Li to Fr) are called the ALKALI metals.
Group 2 (Be to Ra) are the ALKALINE EARTH metals.

Groups 3 to 13 PLUS the metals under the staircase are called the TRANISTIONAL metals.

Group 17 (F to At) are the HALOGENS (salt-makers)
Group 18 (He to Rn) are the NOBLE GASES, which are almost complete inert (unreactive)

There are 7 elements called metalloids (B, Si, Ge, As, Sb, Te, and At). They all touch the staircase on the periodic table, (so do Al and Po, we call them the "AlPo dog-food" exception. These two are metals that tough the stairs, they are not metalloids. A metalloid is a metal (Sb and Ge) that have some nonmetallic properties—which is weird, or a nonmetal (B, Si, As, Te and At) that have some metallic properties. These 7 elements are sometimes called semi-metals.

In Group 3 something very weird happens. Scandium and Yttrium start the group, but in the box "La", there is a black bar going up and down on the right side. Below "La" is "Ac" with that same black bar. Element box 57 really contains ALL the elements from La through Lu (#71), and element box 89 really contains ALL the elements from Ac through Lr. The lower two rows of elements are called in inner transitional metals, and they ALL fit into group 3. All 32 of these elements have very similar properties and all belong in one group, but drawing that group calls for some fancy artwork (and that dumb black bar to remind you).

Box 57 (La to Lu) is called the lanthanide series of metals. Box 89 is known as the actinide series (Ac to Lr). Different tables show this better, make sure you see the big periodic table in the rear of our classroom today.

If you look at the CARBON "key" on the table, each atom has an atomic mass, an atomic number, and an electron configuration. There are also some "selected oxidation states" for most elements.

The MASS = the sum total of the protons and neutrons in the nucleus (approx. 1 AMU each)
The Atomic Number = the number of protons in the atom.
The electron configuration shows the simple locations of the electrons (how many electrons are in each shell)

Since all atoms have the SAME number of protons and electrons, all atoms are electrically neutral, they have no charge. (protons = electrons)

The selected oxidation numbers are used to determine the ratios (formulas) of molecular compounds (nonmetals bonding to other nonmetals). All of the oxidation numbers in all compounds must sum to zero.

Trends of the Periodic Table Basics

Trends are patterns that atoms on the periodic table of elements follow. Trends hold true “most” of the time, but there are exceptions, or “blips”, where the trend seems to do the wrong thing. It is important when deciding a particular trend that you examine 3-4 atoms in a group—or in a period, to see what the numbers are really doing. Choosing just two atoms might show you the exception to a trend rather than the trend.

The 7 trends we study in class are

1. atomic radius (relative size, measured in picometers, pm)
2. average weighted atomic mass (measured in amu)
3. net nuclear charge (how positive is the nucleus, related to # of protons)
4. ion size (cations or anions)
5. electro negativity (relates to bonding)
6. 1st Ionization Energy (energy required to change a mole of atoms → a mole of +1 cations)
7. metal property or non-metal property (how metallic or non-metallic is this atom?)

GROUP TRENDS: THE TREND THAT THE ATOMS FOLLOW GOING <u>DOWN</u> ANY PARTICULAR GROUP

PERIOD TRENDS: THE TREND THAT THE ATOMS FOLLOW GOING <u>ACROSS</u> ANY PARTICULAR PERIOD

Trend #1: Atomic Radius Size

Reference table S provides us with the atomic radius of most atoms. It is the measure of distance from the center of the nucleus to outer most electron shell.

This is measured in picometers (1×10^{-12} meters - trillionths of a meter).

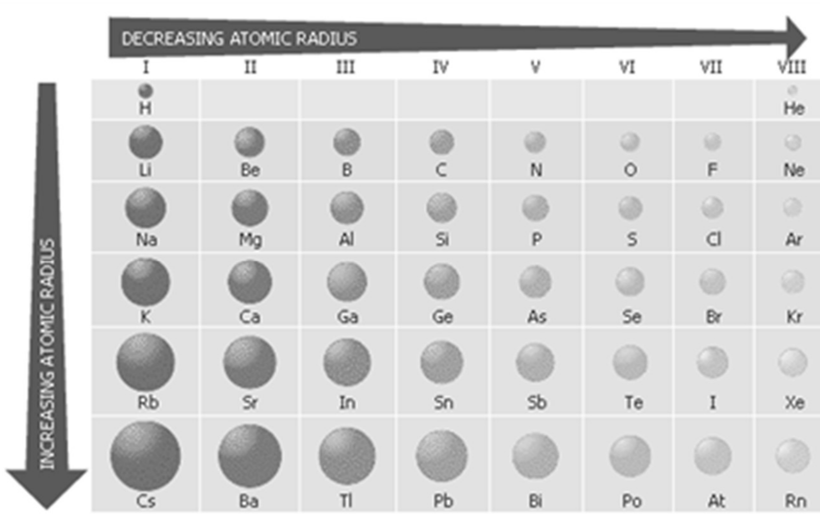
<u>The group trend for atomic size is INCREASING</u>.

That is because the atoms of period 1 (H + He only) each have only one shell.
Period 2 atoms have two shells—which make them bigger.
Period 3 atoms have three shells, bigger still, etc. Period number = number of shells.
More shells = bigger atoms.
The group trend for atomic radius is increasing, without exception.

<u>The period trend for atomic size is DECREASING</u>.

As you go across any of the 7 periods, each atom has the SAME number of shells, but each successive atom adds another positive proton into the nucleus. The more positive protons “pulling” the electrons, in the same number of shells, pull the electrons closer to the center, making the atoms shrink moving across a period.

Usually the smallest atom in the period would be the noble gas. Group 18 atoms, the noble gases, have the MOST PROTONS of all atom in any period.

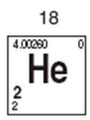


Trend #2: Atomic Mass

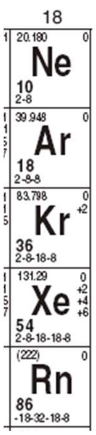
Atomic mass is measured in amu, atomic mass units. The average weighted atomic mass for each atom is listed on the Periodic Table of Elements. Generally speaking the smallest atoms are those with the lowest atomic numbers, and they get heavier as this number increases.

Atomic mass is approximately equal to the number of protons and neutrons in the nucleus. In high school we accept that the mass of an electrons is so small we can disregard.

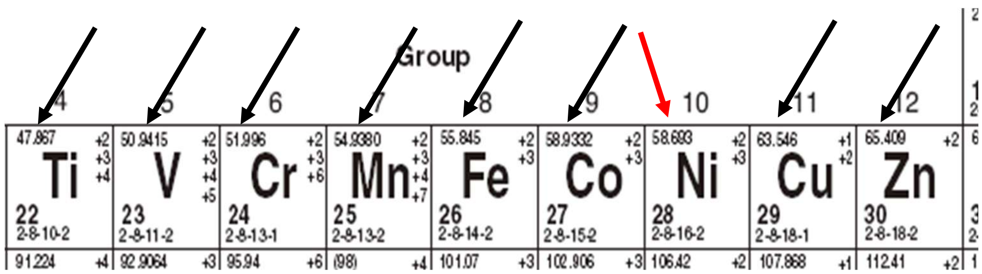
All atoms have isotopes, chemically identical atoms with different masses because they have different numbers of neutrons. Neutrons are neutral, they don't really affect the chemical properties, so all isotopes for a given atom react the same way.



<ins>The group trend for atomic mass is INCREASING</ins>. Check group 18 here. Going down each group on the periodic table each box adds many protons and many neutrons.



<ins>The period trend for atomic mass is also INCREASING</ins>, as we move to the right we are always adding at least one proton per box, and usually one or more neutrons as well. There are some exceptions (see Co → Ni). Exceptions exist due to isotopes of some atoms.



Trend #3: Net Nuclear Charge

The subatomic particles: electrons, protons, & neutrons all have particular charges. Electrons are negative (-1) and are all located outside the nucleus. Neutrons are neutral (Ø) and even though they are in the nucleus, add NO CHARGE to the nucleus. The protons of the nucleus are positively charged (+1) and these protons alone are the measure of net nuclear charge.

This trend is a measure of <ins>how positive</ins> the nucleus of the atom is. It is measured by how many protons, each with a +1 charge, are in a nucleus of an atom. Since each atom has a certain number of protons (the ATOMIC NUMBER), it's easy enough to count the net nuclear charges.

Examples:
He has 2 protons and 2 neutrons in the nucleus, He has a +2 net nuclear charge
Ne has 10 protons and 10 neutrons in the nucleus, Ne has a +10 net nuclear charge
Ar has 18 protons and 22 neutrons in the nucleus, Ar has a +18 net nuclear charge.

<ins>The group trend for net nuclear charge is INCREASING</ins>.

<ins>The period trend for net nuclear charge is INCREASING</ins>.

There are NO exceptions to these trends. Net Nuclear Charge REQUIRES a + sign.
Helium has 2 protons, but the net nuclear charge for helium is +2. <ins>2 is not the same as +2.</ins>

Trend #4: Cation and Anion Sizes

Ions come in 2 varieties, cations are atoms that have lost electrons and become positively charged, they are always metals. Anions are atoms that have gained electrons and become net negatively charged, these are always non-metals.

Ions form by gaining or losing enough electrons to get that "special" stable, noble gas electron configuration. When an ion forms, it obtains a noble gas electron configuration, which is called being ISOELECTRIC to a noble gas. These ions are not noble gases, they obtain the same electron configuration as a noble gas.

When an atom becomes a cation it usually loses ALL of its valence, or outermost electrons. Group one atoms all lose only one electron. Group 2 atoms all lose 2 electrons as they become +2 cations. Metals always lose all of their valence electrons, to become isoelectric to a noble gas.

<u>Cations are always smaller than the atoms they form from</u>.
The sodium cation is smaller than the sodium atom. The calcium cation is smaller than the calcium atom. The aluminum cation is smaller than the aluminum atom.

Cations are always quite a bit smaller than their atoms—a whole shell smaller.

When a non-metal atom becomes an anion it gains enough electrons to obtain a full outer shell, so it can be ISOELECTRIC to a noble gas. Non-metals can gain one, two, or even three electrons to fill the outer valence shell. When the atoms gain electrons in the valence shells this valence shell stretches a bit to accommodate this influx of negative charge. The electrons all repel each other, and the extra electron will force all the electrons in that shell a bit further away from each other.

<u>Anions are always slightly larger than the atoms that they form from</u>.
Nitrogen is smaller than the nitride anion. Oxygen is smaller than the oxide anion. Fluorine is smaller than the fluoride anion.

Anions gain electrons, which stretches out their VALENCE shell when they form,

Atomic size increases going down a group because each atom lower on the table has more shells. Same for cations, although a cations are smaller than the atoms they formed from, each successive cation has more shells than the cation above it.

Cations get smaller going across a period as each cation has the same number of electrons as all the other cations in the period, but they gain electrons across the period, so more protons are pulling on the same number of electrons. Cations Na^{+1} , Mg^{+2}, and Al^{+3} all have 10 electrons in a 2-8 configuration but they will have 11, 12, and 13 protons respectively. Sodium is larger than magnesium because Mg has that extra proton pulling the outer shell in tighter. The aluminum +3 ion is smallest because it has yet another proton pulling on the same number of electrons in the same number of shells, 13 positive protons pulls "harder" than 12.

<u>THE GROUP TREND FOR CATION SIZE IS INCREASING</u>. <u>THE PERIOD TREND FOR CATION SIZE IS DECREASING</u>.

Anions are all slightly larger than the atoms they formed from because they add extra electrons into the outer orbital, where it can fit, but this forces all the negatively charged electrons a bit further away from themselves. Looking at group 16, oxygen sulfur and selenium, all get larger as atoms moving down the table. Each anion is larger than it's atom, and the anions get progressively larger as well. Each anion below has one more orbital than the ion above it.

For the period trend of anion size let us look at period 3: phosphorous, sulfur and chlorine.

The atoms get progressively smaller going across the period (more protons pulling on the same number of electron shells). The anions for these three form as P^{-3}, S^{-2}, and Cl^{-1} respectively. Each has TEN electrons in the same 2-8 configuration, but since there are more and more protons in these ions moving across the period, the anions get smaller moving across the period.

THE GROUP TREND FOR ANION SIZE IS INCREASING.

THE PERIOD TREND FOR ANION SIZE IS DECREASING.

Group Trends for Ions
examples: Group 2 and Group 16

Group 2 ION electron config.	Group 16 ION electron config.
Be^{+2} 2	O^{-2} 2-8
Mg^{+2} 2-8	S^{-2} 2-8-8
Ca^{+2} 2-8-8	Se^{-2} 2-8-18-8
Sr^{+2} 2-8-18-8	Te^{-2} 2-8-18-18-8
The ions get larger going "down" any group, they each have one more full electron shell	

Period Trends for Ions
examples: Period 3 cations and anions
Ion Electron Configurations listed

Na^{+1}	Mg^{+2}	Al^{+3}
2-8	2-8	2-8
11 protons	12 protons	13 protons
All have 2-8 electron counts in 2 shells. The extra protons pull harder on these shells. Cations get smaller across a period.		
P^{-3}	S^{-2}	Cl^{-1}
2-8-8	2-8-8	2-8-8
15 protons	16 protons	17 protons
All have the same 2-8-8 electrons, in 3 shells. The extra protons pull harder on these shells. Anions get smaller across a period.		

Trend #5: Electronegativity

Electronegativity is the tendency of an atom to attract electrons to itself when it makes a bond.

When atoms make covalent bonds, they "share" electrons. Each atom puts up one electron, and when a pair of electrons is shared, a single covalent bond forms.

How well they share is measured using the difference in their electronegativity values. The values were set by Linus Pauling (Nobel prize for this work). The higher the Electronegativity value, the harder the atoms pulls electrons to itself. The lower the value, the less attraction the atom has for these bonding electrons.

If two atoms in the bond have the SAME electronegativity values, there is NO difference in EN Value, which means the atoms share the electrons evenly, or, the bond is NONPOLAR.

If the atoms in a bond have different EN Values, the atom with the higher EN Value "pulls" harder, getting the bonding electrons more of the time, making that side of the bond NEGATIVE. This leaves the atom with the lower EN Value get the electrons in the bond less often, which leaves that side of the bond mostly POSITIVE.

The greater the difference the less fairly the atoms share the electrons, which makes the bonds MORE POLAR

Examples of NONPOLAR BONDS include: H_2, Cl_2, and Br_2. Each shares a single pair of electrons, forming a nonpolar covalent bond. Since both of the H atoms have the same 2.2 EN Value, the difference between these values is zero. That makes this bond nonpolar.

With HCl, hydrogen monochloride forms, the H atom (2.2 EN) and the Cl atom (3.2 EN) have a (big) 1.0 difference in electronegativity. This bond is a POLAR BOND.

When H_2O forms, there are 2 H atoms that bond to ONE oxygen atom. Each bond is individual; they both get examined separately.

The H atom (2.2 EN) and the O atom (3.4 EN) have a big 1.2 EN difference. This bond is very polar.

Water makes two very polar covalent bonds inside the one molecule.

Substance	Atom	The other atom	EN Difference	Bond Polarity
H_2	H 2.2	H 2.2	0	nonpolar
Cl_2	Cl 3.2	Cl 3.2	0	nonpolar
Br_2	Br 3.0	Br 3.0	0	nonpolar
HCl	H 2.2	Cl 3.2	1.0	polar
H_2O	H 2.2	O 3.4	1.2	very polar

Comparing H—Cl and the H—O bonds, we compare the differences in EN Values. The H—O bond is MORE POLAR than the H—Cl bond. We can use the EN Value differences to "rank" bond polarity.

In Regents Chem bonds are nonpolar with a zero EN Value difference, or they are POLAR because there is a difference in electronegativity. In college, sometimes a very small electronegativity difference is called nonpolar, but not in high school chem.

Linus Pauling created this concept of electronegativity and determined that fluorine has the greatest tendency to gain electrons in a bond compared to ALL OTHER ATOMS. Since he created this scale, he could do what he wanted, and he did.

Dr. Pauling decided that F would has an EN value of 4.0, the highest value on the table. All other atoms would be compared or ranked to Fluorine. Since all atoms are compared to a single "standard" atom, this makes electronegativity a relative scale. All atoms are measured, in relation to, or compared to a standard that Pauling decided upon.

Electronegativity is an example of a relative scale. Electronegativity has no units.

He choose the numbers 0 → 4.0 for no particular reason. The numbers don't "mean" anything, they are just numbers that "rank" the atoms. Since the numbers don't mean anything, this is an ARBITRARY scale too.

Table S shows all the EN values for the elements. Some atoms, the smaller noble gases have no EN values. These atoms do not make bonds ever, they have NO TENDENCY to gain (or lose) electrons.

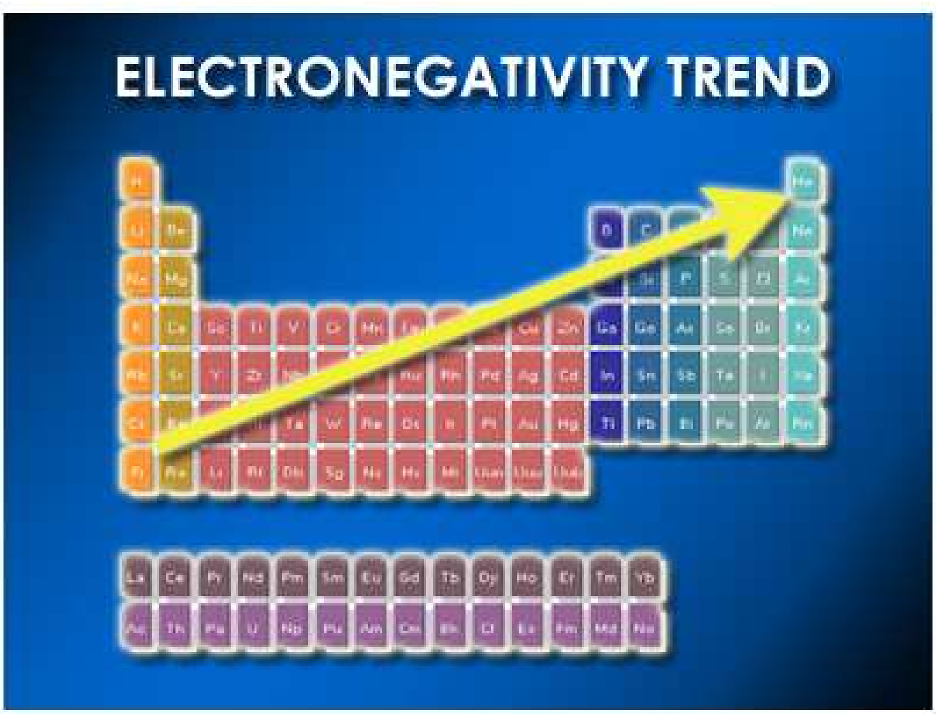
THE GROUP TREND FOR ELECTRONEGATIVITY IS DECREASING. THE PERIOD TREND FOR ELECTRONEGATIVITY IS INCREASING.

The general table trend is the closer you are to F, the higher the EN value (except for noble gases).

The noble gas XENON has an electronegativity, and under some unusual conditions (very high pressure and catalysts) it can be forced into making some bonds. This is an exception, noble gases tend to be relatively INERT, or non-reactive.

The period trend for electronegativity is increasing, because in any period you have one shell, and more and more protons pulling tighter and tighter. There is a greater inward attraction to gain electrons as you move across the table. Until you get to group 18, which already has full shells.

The group trend for electronegativity is decreasing. This seems odd at first, more and more protons are in the atoms going down any group, but this increased positive charge in the nucleus pulling inwards is offset by the distance the outermost valence shell is to the nucleus. More protons helps pull inward, distance hurts more than the increased number of protons helps.



The closer to FLUORINE, the higher the electrongativity value (noble gases excepted).

The bottom left side of the table has the lowest electronegativity values.

Trend #6: 1st Ionization Energy

When atoms of group 1 become cations and “lose” an electron, even though they “want” to do this to gain the noble gas electron configuration, it requires some energy. The electrons do not FALL OFF of atoms. 1st Ionization energy is the energy required to pull a mole of electrons off a mole of atoms.

The amount of energy required to take a mole of atoms and make them a mole of +1 cations is called the 1st IONIZATION ENERGY. The unit is kJ/mole or kilo-joules per mole.

The metals have lower first ionization energy requirements for several reasons. Metals will tend to lose electrons easier than nonmetals. Metals form cations. Nonmetals gain electrons to form anions.

The largest atoms in any period are in group 1, and these atoms have the lowest 1st ionization energy. Moving across any period, you have the same number of orbitals, but increasingly more protons, pulling the atoms smaller. For the same reason that the atoms get smaller (greater inward attraction) the more difficult it becomes to pull these electrons off.

THE PERIOD TREND FOR FIRST IONIZATION ENERGY IS INCREASING.

Going down any group, since the atoms are getting larger (more shells). This extra distance of electrons to the protons hurts more than the extra protons helps, so it becomes easier to get the electrons off the bigger atoms, even with extra protons.

THE GROUP TREND FOR 1ST IONIZATION ENERGY IS DECREASING.

A mole of atoms can be converted into a mole of +1 ions by applying the 1st ionization energy. Anything (almost) is possible. Making fluorine a +1 cation is possible, but hard. It requires a lot of energy compared to lithium. Noble gases can be forced into +1 cations too. That is quite hard to do—but possible.

Atoms like Mg and Ca make +2 ions. Al makes a +3. To convert a mole of Mg into a mole of +2 ions requires the application of the 1st Ionization energy PLUS the application of the 2nd Ionization energy (the energy required to remove a second electron from a mole of +1 ions to make them form into +2 ions.)

There is also a 3rd Ionization energy. (We concern ourselves only with the 1st Ionization energy, not the 2nd or 3rd, or others. It’s added here because you are smart enough to understand it).

Noble gases tend to be “perfect” atoms, and it’s very hard to remove their electrons, as the Table S shows us their very high 1st ionization energies. The atom with the highest 1st ionization energy is helium, only two electrons but super small. It’s so hard to remove those electrons!

Trend #7: Metallic Property and Non-Metallic Property

Metals are on the left side of the Periodic Table of Elements. They have many properties that make them “metallic”, such as luster, electrical conductivity, heat conductivity, they’re malleable, ductile, they have low specific heat capacity constants, higher density, higher melting point, and only form cations, etc.

If each property could somehow be ranked, if we “measure” all metals against each other, the MOST METALLIC METAL would be Francium, Fr.

In fact, the closer to Fr a metal is on the table (LITERALLY, in inches) the more metallic it is.
Examples - polonium, lead, silver and zirconium are all metals. Zr is the "closest" to Fr on the table, therefore, Zr is the most metallic of these four metals.

Non-metals are on the right side of the table (plus Hydrogen). They have pretty much the opposite properties of metals. Nonmetal solids are brittle, not able to change shape. They don't conduct heat or electricity, they form only anions and are dull colored. The nonmetal gases are clearly not metallic.

If these nonmetals were ranked, which is the most nonmetallic of them all?
Helium is the MOST NON-METALLIC nonmetal. The closer atoms are to He, the more non-metallic it is.

Example, C, Cl and Ne are all non-metals, neon is the closest of these to helium, so neon is the most non-metallic of these three nonmetals.

Sometimes kids think of crazy questions, like which is more metallic, Cesium or radium, or which is more non-metallic, F or Ne. These questions cannot be answered by our simple "proximity" rule. Even the regents exam writers are not that cruel.

Parts of the Periodic Table of Elements

Even the name of the table is important. When the atoms are arranged in increasing order of atomic number there is a periodic repetition of their chemical properties, in the groups. That is the PERIODIC LAW and why it's the PERIODIC table of the elements.

Group 1 = the alkali metals
Group 2 = the alkaline Earth metals
Group 3 through 12 <u>plus</u> some the atoms under the staircase = Transitional metals
Group 17 = the halogens
Group 18 = the noble gases
The bottom 2 rows of the table are detached. These are the Inner Transitional metals.
They all fit into GROUP 3, in periods 6 + 7

Hydrogen is an exception; it's a non-metal that *sometimes* acts like a group 1 metal when bonding

7 Metalloids touch the staircase line (includes B, Si, Ge, As, Sb, Te and At) but not Al & Po.
Aluminum & Polonium touch the stairs, but they're metals, exceptions to this rule. Besides, AlPo is dogfood.

All atomic masses are based against carbon-12, an atom with 6 protons and 6 neutrons. It's said to have the exact mass of 12 amu. One AMU is one twelfth the mass of one C-12 atom.

At STP, all metals are solids, except for Hg, which is liquid.

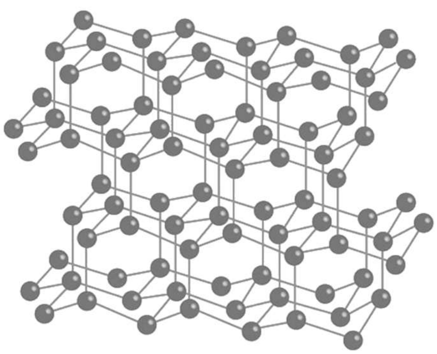
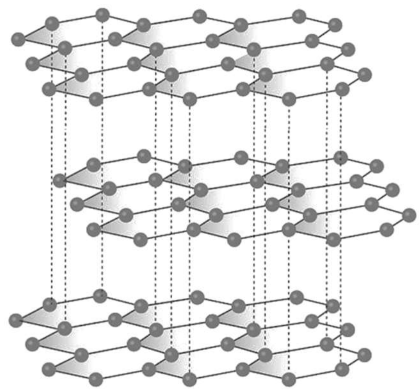
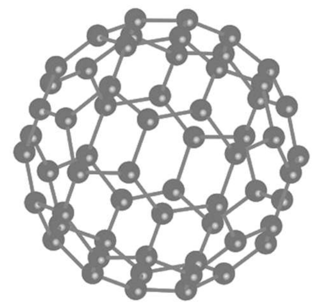
At STP most non-metals are gases, but Br is a liquid, B, C, Si, P, S, As, Se, Te and I are solids.

The modern periodic table was first devised by Dimitri Mendeleev, a Russian scientist. It was a tremendous achievement of figuring out a pattern for 62 elements into a table that made sense. It was the shape of the table that made this possible and was likely the hardest part to imagine for him.

ALLOTROPES

Allotropes are pure forms of an element, but they bond together in a different way, so it has different properties.

3 allotropes of carbon, left to right are diamond, graphite, and the Buckminster fullerene
The 3 forms of the pure carbon have different structures and different properties.

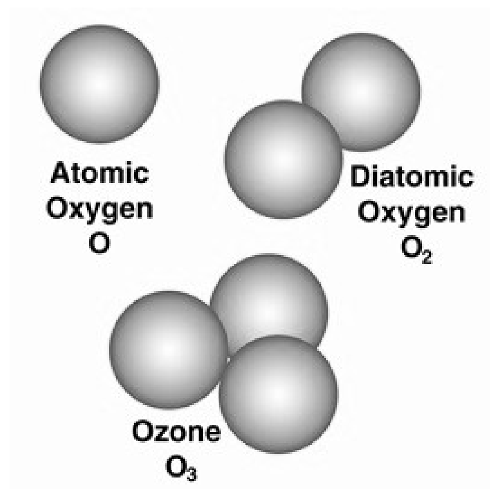




Oxygen can exist as unstable single atoms.

Or as O_2 diatomic oxygen that we breathe to stay alive and is key to combustion reactions.

The third oxygen type is ozone O_3, which protects us from harmful rays from the Sun. Ozone is not good to breathe.

These are all different, although they are all pure oxygen.



Trends of the Periodic Table Notes

1. ___________________________________ was the Russian man who invented the modern Periodic Table.

2. The COLUMNS on the table are called __________________________. There are ________ on the table.

3. The ROWS that go across the table are called ________________________. There are _______ of these.

4. Group 1 (Li to Fr) are called the ____ metals

5. Group 2 (Be to Ra) are called the ____ metals

6. Groups 3 to 12 (plus under the stairs) are the ____ metals

7. On the right side of the staircase (and H) are the ____

8. Group 17 (F, Cl, Br, and I) are called the ____

9. Group 18 (He, Ne, Ar, Kr, Xe, and Rn) are the ____________________________________gases

10. Seven (out of 9) atoms that touch the staircase are called the ____________________________________.

11. The 2 disconnected rows at the bottom are the ____ metals.

They ALL fit into GROUP _____ and PERIODS ___ + ____

12. Atoms in the same group share many ______________ properties, because they have the same number of

valence (outermost) ___________________________ , which means they all bond in similar ways.

13. Skip this one.

14. What does PERIODIC mean?

15. The elements of the Periodic Table are arranged in order of increasing

16. The Periodic Law states….

17. Similar properties "show up" periodically, IN THE ___________________________.

18. At the top of the table, the properties repeat ____________________________________ *atoms*,

but in the middle of the table it's __________________________________ *atoms*.

19. The periods of the table go __________________ to ____________________.

20. The periods contain many elements that have ____.

21. Period numbers correspond to the number of ________________________ that all atoms in a period have

22. Period	Example element	Element names	Electron configuration	Number of electron shells
	H		1	shell
	Be		2-2	shells
	S		2-8-6	shells
	Mn		2-8-13-2	shells
	Xe		2-8-18-18-8	shells
	Ba		2-8-18-18-8-2	shells
	Ra		2-8-18-32-18-8-2	shells

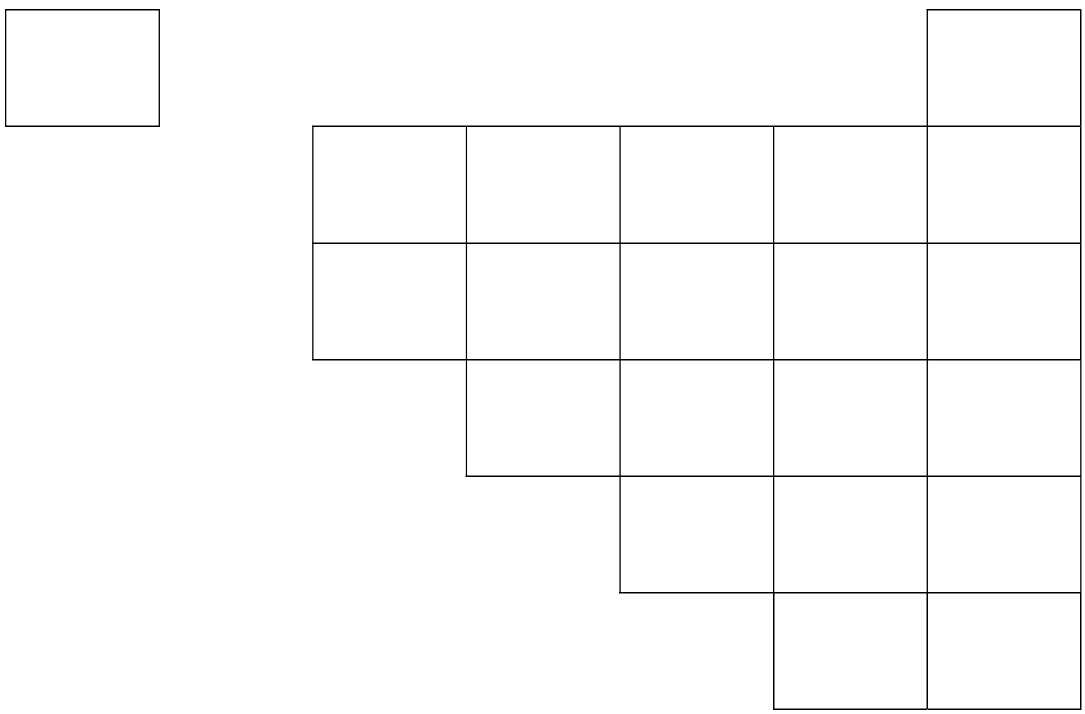
23. Subatomic Particles	Location	Charge	Mass	Symbol
	nucleus			
	nucleus			
	outside of the nucleus			

This is a cool way to show an element and it’s important numbers $^{59}_{27}\mathrm{Co}$	24. How many protons, neutrons and electrons in cobalt?

25.

List ALL the nonmetals

How many nonmetals are on the whole Periodic Table?



26. List symbols of all the metalloids. Each number corresponds to an atomic number.						
5	14	32	33	51	52	85

27. How many elements are METALS? 118 total elements - 22 nonmetals = _____________ metals

8. We will examine ___ trends on the periodic table.

Trends going down are group are called →

Trends going across any period are called →

29. These trends are…			
	The trends on the periodic table	have these units	find details in
1	atomic mass	amu	Table S
2	atomic radius	pm or picometers	Table S
3	net nuclear charge	no units	Periodic table
4	1st Ionization energy	kJ/mole	Table S
5	metallic vs. nonmetallic properties	no units	Periodic table
6	cation sizes and anion sizes	no units	Periodic table
7	electronegativity values	no units	Table S

30. What is the group trend for atomic mass?		
Group 2	Mass in amu	
Be	amu	
Mg	amu	
Ca	amu	
Sr	amu	

31. What is the period trend for atomic mass?				
Period 3	Na	Mg	Al	Si
Mass in amu	amu	amu	amu	amu

The period trend for atomic mass…

32. What is the group trend for atomic radius?

Group 1	Radius in pm
Li	pm
Na	pm
K	pm
Rb	pm

33. What is the period trend for atomic radius?

Atom	Li	Be	B	C
Radius in pm	pm	pm	pm	pm

34. What is the group trend for net nuclear charge?	Atom	Atomic Number (number of protons)	Net Nuclear Charge
	Be		
	Mg		
	Ca		
	Sr		
	Ba		

35. What is the Period Trend for Net Nuclear Charge?				
atom	K	Ca	Sc	Ti
atomic number				
net nuclear charge				

36. 1st Ionization Energy is the amount of energy required to turn 1 mole of atoms → 1 mole of +1 cations.

Look at table S, what is the 1st Ionization Energy for LITHIUM? ____________________

______________________ → (____________________ + ________________________)

Lithium electrons do not fall off, ____

What about Mg? Or Al?

37. What is the group trend for 1st Ionization Energy?

Group 1	1st ionization energy kJ/mole	Electron configuration
Li		
Na		
K		
Rb		

38. What is the period trend for 1st Ionization Energy? (period 5)				
Atom	Rb	Sr	Y	Zr
1st Ionization Energy	kJ/mole	kJ/mole	kJ/mole	kJ/mole

45. List some metallic properties	46. List some nonmetallic properties

47. If you could rank all the metals, for all their properties,

the most metallic element is ______________________

48. If you could rank all the nonmetals for all their properties,

the most nonmetallic element is _______________________ (this is dopey)

49 Circle the most metallic of these: strontium copper lead

50 Circle the most nonmetallic of these: sulfur bromine neon

Cation Sizes and Anion Sizes

We have no charts to look over to determine the actual sizes of any ions, but we can still figure out the trends of cation sizes and of anion sizes by thinking.

51. How big is a PICOMETER?

52. group 1 CATIONS	Electron configurations of cations	53. group 17 ANIONS	Electron configurations of anions
Li^{+1}		F^{-1}	
Na^{+1}		Cl^{-1}	
K^{+1}		Br^{-1}	
Rb^{+1}		I^{-1}	

54a. State the group trend for cation size.

54b. State the group trend for anion size.

54c. State the group trend for all ion sizes.

55. Why does this trend exist?

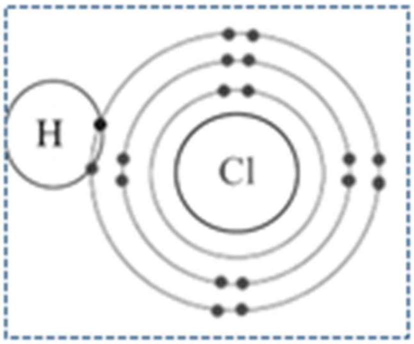
56. What is the period trend for cation size?			
CATIONS Period 3			
Electron configurations			

56. What is the period trend for anion size?			
ANIONS Period 2	N^{-3}	O^{-2}	F^{-1}
56. Electron configurations			

ELECTRONEGATIVITY

61. Dr. Linus Pauling defined Electronegativity as

62. Let's imagine the bond between H + Cl in HCl.
H has one electron; Cl has seven valence electrons.
They share electrons with each other so that they both get a "FULL" valence electron shell.



Cl *electronegativity Value*:

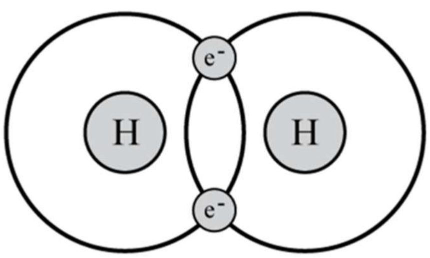
H *electronegativity Value*:

____

Difference in Electronegativity

THIS IS A →

63. Let's imagine the bond between 2 Hydrogen atoms in H_2.
H has 1 valence electron; the other H atom also has 1 valence electron.
By sharing their electrons both atoms get a "FULL" valence electron shell.



H *electronegativity Value*:

H *electronegativity Value*:

____

Difference in Electronegativity

THIS IS A →

64. The greater the difference in 2 atoms electronegativity makes a ______________________________ bond.

65. No difference in electronegativity makes a ___________________________________ bond.

66. By comparing bonds electronegativity differences, you can rank bonds from

67. Draw this:		The arrow itself is a dipole arrow. It shows the polarity of the bond. The arrowhead points to where the electron “goes” (more negative). The arrow tail is a “positive” sign, it shows H is more positive

68. What is the group trend for electronegativity value?

Group 1 Atoms	electronegativity values		Group 17 Atoms	electronegativity values
Li			F	
Na			Cl	
K			Br	
Rb			I	

69. What is the period trend for electronegativity value?

	Li	Be	B	C	N	O	F	Ne
Electro negativity Values								

70.

71. Exceptions to the trends.

Here are the atomic masses for Period 4. Which way does the trend go (up or down?)

atoms	Mn	Fe	Co	Ni	Cu
amu	54.9380 u	55.845 u	58.9332 u	58.693 u	64.456 u
Trend arrows					

73. Does nickel destroy the trend?

74. These are the atomic radius measures in period 2. Fill in arrows to show the trend.

Atoms	Li	Be	B	C	N	O	F	Ne
Radius in pm	130.	99	84	75	71	64	60	62
Arrows	start	↓						

State the period trend for atomic radius. Mention neon

75. Noble gases have no tendency to make bonds ever, they don't have electronegativity values either, right?			
Atomic number	Symbols	Atom Names	Electronegativity value
2		Helium	
10		Neon	
18		Argon	
36		Krypton	
54		Xenon	
86		Radon	

76. Whoa! What's up with Xe?

77. Are there exceptions to Net Nuclear Charge?

78. Predict the approximate sizes of these cations & anions					
Atom	Atom electron config.	Atomic Radius (look up in table S)	Ion	Ion Electron configuration	Ionic Radius
Li lithium	2-1	pm	Li^{+1}	2	pm
Mg magnesium	2-2-2	pm	Mg^{+2}	2-8	pm
Sc scandium	2-8-9-2	pm	Sc^{+3}	2-8-8	pm
O oxygen	2-6	pm	O^{-2}	2-8	pm
P phosphorous	2-8-5	pm	P^{-3}	2-8-8	pm

79. Cations are always…

Anions are always…

80. A relative scale is one that…

81. Electronegativity is a relative scale, all atoms being relative to ____________________________
Dr. Pauling determined that fluorine has the greatest tendency to gain electrons when making a bond.

82. An arbitrary scale is one that uses numbers that ____.
Dr. Pauling chose 4.0 for his highest value, given only to fluorine. All other values descend to zero.

83. The electronegativity scale is both…

84. Allotropes are pure forms of an element that have different bonding patterns which leads to allotropes having different physical properties and different chemical properties too.

85. Common examples of allotropes include

Carbon	Oxygen
1	1
2	2
3	

Group Trend For	Fill in full sentences and memorize these all.
atomic mass	
atomic radius	
ionic radius	
net nuclear charge	
1st Ionization Energy	
electronegativity	
metallic property	
nonmetallic property	
Period Trend For	Fill in full sentences and memorize these all.
atomic mass	
atomic radius	
ionic radius	
net nuclear charge	
1st Ionization Energy	
electronegativity	
metallic property	
nonmetallic property	