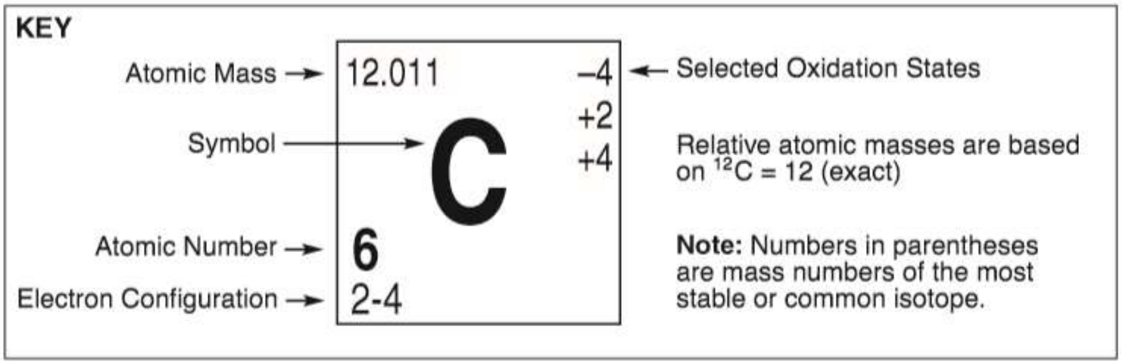
The Periodic Table
Unit 3

The Periodic Table

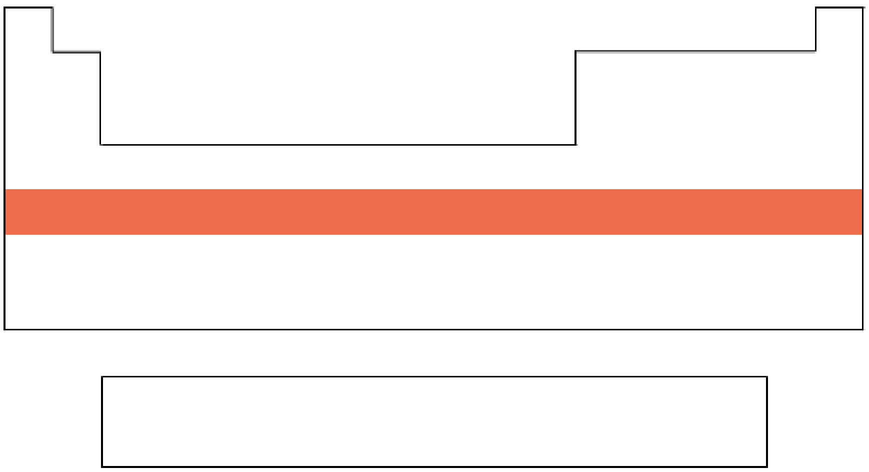
Reading the Periodic Table – use the key



Periods

Horizontal rows 1-7

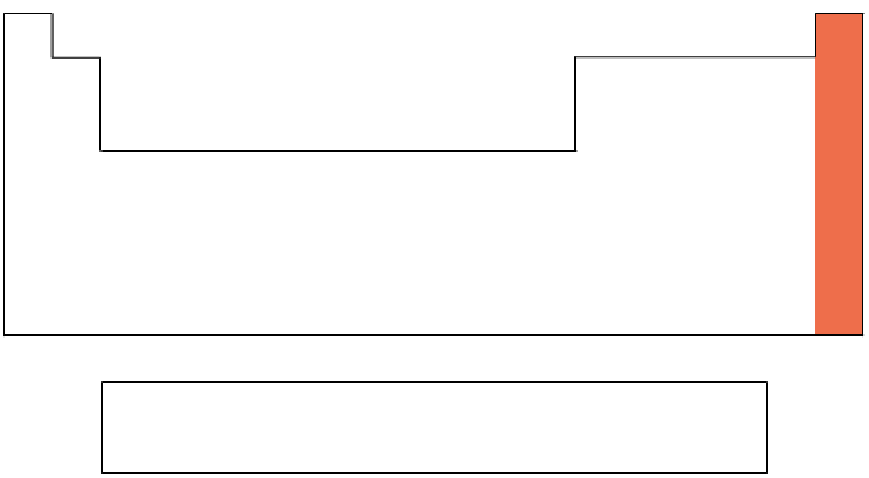
- the period tells the principal energy level (number of electron shells)



Groups

Vertical columns 1-18

- the group tells the number of valence electrons (subtract 10 for Groups 13 – 18, He = 2)



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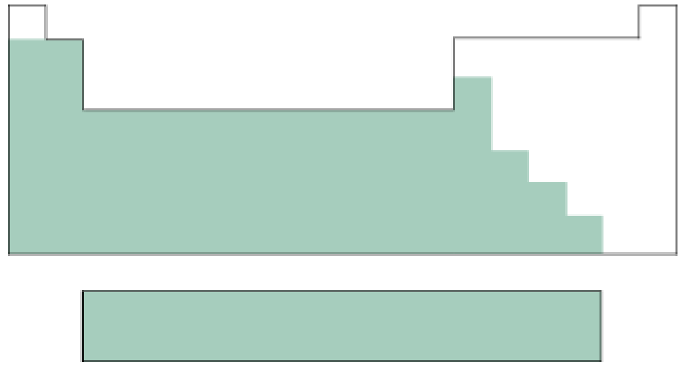
Metals

77% of the elements on the Periodic Table are metals

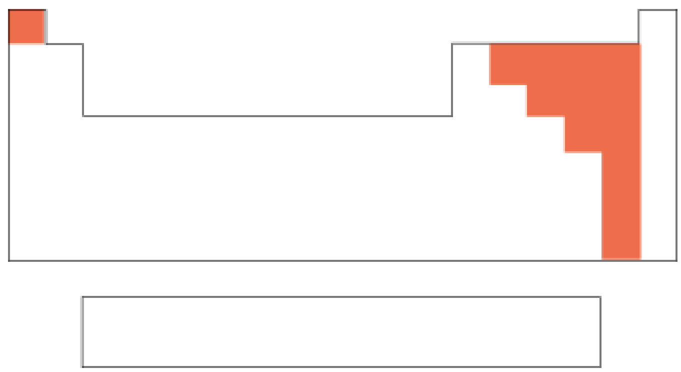
- elements to the left of the zig-zag line are metals (except H)
- metals have low ionization energy (lose electrons easily forming cations)
- metals have low electronegativity (have little attraction for electrons)
- metals have luster (are shiny)
- metals are malleable (can be hammered flat)
- metals are ductile (can be pulled into wires)
- metals are good conductors of heat and electricity
- Fr is the most reactive metal



Nonmetals

About 11% of the elements on the Periodic Table are nonmetals

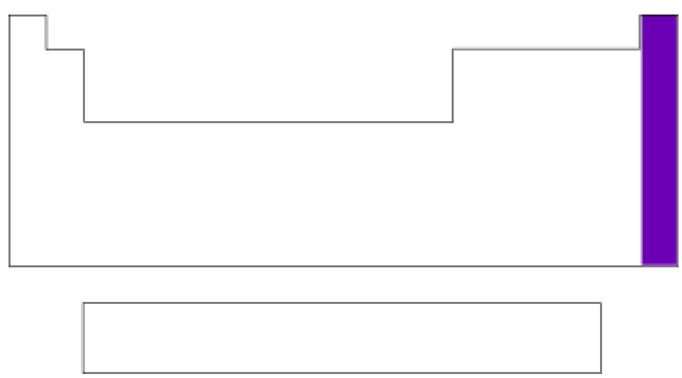
- elements to the right of the zig-zag line are nonmetals (including H)
- nonmetals have high ionization energy (electrons are ***not*** easily lost)
- nonmetals have high electronegativity (tend to gain electrons and form anions)
- nonmetals lack luster and are brittle
- nonmetals are poor conductors of heat and electricity
- most nonmetals are gases at room temperature (Br_2 is a liquid, heavy elements are solids)
- F is the most reactive nonmetal



Noble gases

Group 18 is called the noble gases (about 6% of the elements)

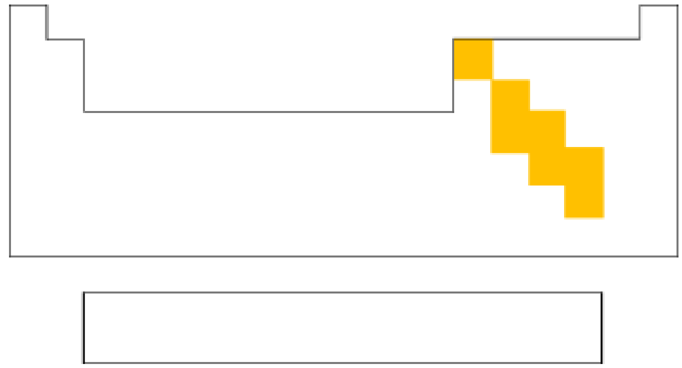
- noble gases have complete outer principal energy levels and are stable (chemically, they have very low reactivity)
- Kr, Xe, and Rn can react with O and F
- noble gases are monatomic
- because dispersion forces (van der Waals' forces) increase with increasing atomic radius, the boiling points of the noble gases increases moving down the group



Metalloids

Elements that touch two of the zig-zag lines (about 6% of the elements)

- metalloids have some properties of both metals and nonmetals
- metalloids include B, Si, Ge, As, Sb, and Te (Po is in doubt)
- Al is ***not*** a metalloid
- metalloids are most often used is semiconductors, for LED lights, and in electronics



Arrangement of Elements on the Periodic Table

Elements on the modern Periodic Table are arranged in order of increasing atomic number

Periodic Law: properties of the elements are a periodic function of their atomic number

1. Another way to think of the periodic law is that if elements are arranged in order of increasing atomic number (*not* atomic mass), their properties will repeat in a periodic pattern (hence the name, Periodic Table)
2. Elements were first organized based on their chemical and physical properties
   This means that the position of an element on the Periodic Table shows information about that element's chemical and physical properties
   Physical properties include: density, conductivity, malleability, solubility, hardness, color, and phase (melting point and boiling point)
   Chemical properties: how an element behaves in a chemical reaction (gain or lose e^-)
3. The result of all this organization is that scientists were able to conclude that the properties of the elements are a result of the valence electrons
   Valence shell: the outermost energy level of an atom
   Valence electrons: electrons in the valence shell of an atom
   Determine the number of valence electrons from the electron configuration
   Atoms in the same Group with have the same number of valence electrons

Allotropes: different forms of the same element in the same phase

- sometimes due to different chemical formulas
  Oxygen exists as O=O (O_2, or oxygen) and O=O-O (O_3 or ozone)
- sometimes due to different crystalline structure
  Carbon exists as diamond or graphite
  Phosphorus exists as white (P_4), red (polymeric), violet, and black forms

Properties that exhibit periodic trends (see Table S):

Atomic radius: half the distance between adjacent nuclei

Ionic radius: radius of an atom after it becomes an ion (gains or loses an e^-)

First ionization energy: energy required to remove the first electron from the outermost (or valence) electron shell

Electronegativity (EN): ability of an atom to attract electrons in a chemical bond (higher EN = higher attraction)

Metallic and nonmetallic properties

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Periodic Trends

Trends are general tendencies, not perfect predictions

Metal, nonmetal, and metalloid trends

- Metals
  - Chemically – tend to lose electrons to form cations
    - Cations tend to be smaller than the atoms that form them (fewer e^-, so less repulsion)
  - Low ionization energies
  - Low electronegativities
  - Good conductors of heat and electricity
  - High luster, malleability, and ductility
  - High melting and boiling points (Hg is the only metal that is liquid at room temperature)
  - Become more reactive down a Group
  - Become less reactive left to right across a Period
- Nonmetals
  - Chemically – tend to gain electrons to form anions
    - Anions tend to be larger than the atoms that form them (more e^-, so more repulsion)
  - High ionization energies
  - High electronegativities
  - Poor conductors of heat and electricity
  - Low luster and tend to be brittle
  - Low melting and boiling points
    - of 16 nonmetals, 11 are gases at STP, Br_2 is the only liquid, C, P_4, Se, and I_2 are solids
  - Become less reactive down a Group
  - Become more reactive left to right across a Period (except noble gases which rarely react)
- Metalloids
  - Metalloids touch two lines on the 'staircase' (except Al which is a metal)
  - Only six elements are metalloids
    - B, Si, Ge, As, Sb, and Te
  - Exhibit properties of both metals and nonmetals (but, chemically, tend to gain e^-)
  - Commonly used for computer chips, glass, alloys, and LED lights

Group trends

Generally, elements in the same Group:

- react similarly
- have the same number of valence electrons (see electron configurations)
- number of electron shells increases (top to bottom) decreasing pull from the nucleus
  - atomic radii increase from top to bottom
  - electronegativities decrease from top to bottom
  - 1st ionization energies decrease from top to bottom

Group 1: the alkali metals (H is not an alkali metal, it can form a 1- ion)

- these elements are never found native (never found in the elemental state naturally)
- all have only *one* valence electron
- lose one electron to form a 1+ cation (positive ion)
- extremely reactive or active metals
- reactivity increases from top to bottom of the Group (Fr is the most reactive metal)
- low 1st ionization energy
- low electronegativity (EN)

Group 2: the alkaline earth metals

- these elements are also never found native
- all have *two* valence electrons
- lose two electrons to form 2+ cations (positive ions)
- very reactive or active metals
- reactivity increases from top to bottom of the Group
- low 1^{st} ionization energy
- low electronegativity (EN)

Group 14: the crystallogens (or the carbon group)

- this Group contains carbon (the basis of almost all life on Earth)
- have *four* valence electrons
- tend to share electrons rather than losing or gaining electrons
- contains nonmetals, metalloids, and metals
  - carbon (at the top) is a nonmetal
  - silicon and germanium are metalloids
  - tin and lead (at the bottom) are metals
    - ► This trend is true for the entire Periodic Table, elements become more metallic toward the bottom of the Group

Group 17: the halogens

- have *seven* valence electrons
- are most likely to gain *one* electron to form a 1^- anion (negative ion)
  - F is the most reactive element (EN = 4.0) and *only* forms 1^- anions
  - Cl, Br, and I can form + oxidation states, usually when reacting with oxygen
- all the halogens are diatomic molecules by nature
- dispersion forces increase from the top to the bottom of the Group
  - F_2 and Cl_2 are gases at STP
  - Br_2 is a liquid at STP
  - I_2 is a solid at STP
    - ► This trend is also true for the entire Periodic Table, dispersion forces increase from the top to the bottom of a Group

Group 18: the noble gases

- these elements are also never found native
- *eight* valence electrons or full octets (except He with 2 valence electrons)
- are very noble or nonreactive
  - He, Ne, and Ar have never been forced to react with other elements
  - Kr and Xe have been reacted with O and F
- form monatomic gases
- dispersion forces increase from the top to the bottom of the Group
  - boiling points (IMFs or bond strengths) increase from top to bottom of the Group

Groups 3 through 12: transition metals: (d block)

- generally have *two* valence electrons, but the number of valence electrons and the number of electrons that can react can vary because some inner electrons are involved
- ions (and solutions of these ions) of these elements can be distinctively colored

Inner Transition Metals: (f block, no Group numbers)

- lanthanides (atomic numbers 57 – 70)
- actinides (atomic numbers 89 – 102)

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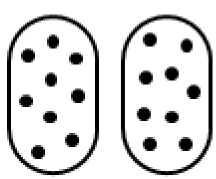
Period trends (with increasing atomic number, or from left to right)

Generally, there is a gradual change in properties as atomic number increases in a Period

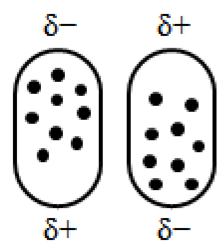
- Period number tell the number of electron shells (always the same in a given Period)
- number of valence electrons increases with increasing atomic number in the main groups (s and p blocks)
- number of protons in the nucleus increases with increasing atomic number and increasing the number of protons increases the pull toward the nucleus
  - oxidation states gradually change from positive to negative
    - metals at the left of a period
    - metalloids at the ‘staircase’
    - nonmetals at the right of a period
  - ionic radii
    - decreases for cations or positive ions (due to decrease in $e^- \leftrightarrow e^-$ repulsion)
    - increases for anions or negative ions (due to increase in $e^- \leftrightarrow e^-$ repulsion)
  - atomic radius decreases from left to right
  - 1st ionization energy increases from left to right
  - electronegativity increases from left to right

Dispersion forces: weak attractions between particles due to the constant motion of electrons

Dispersion forces are weak because they only last for a very brief time



random dispersion results in neutral charge dispersion



electron motion causes asymmetric charge distribution and momentary charges attract

A Brief History of the Periodic Table

1817 – Johann Wolfgang Döbereiner: The Law of Triads

A group of three elements with similar properties, the center element has properties that are the average of the first and third element

Atomic weights: Li = 6.9 Na = 23.0 K = 39.1

1865 – John A Newlands: The Law of Octaves

Made a chart of elements arranged in order of increasing atomic weights and noted that properties repeated every eighth element, like the octave on a musical scale

	No.		No.		No.		No.		No.		No.		No.		No.
H	1	F	8	Cl	15	Co & Ni	22	Br	29	Pd	36	I	42	Pt & Ir	50
Li	2	Na	9	K	16	Cu	23	Rb	30	Ag	37	Cs	44	Os	51
G	3	Mg	10	Ca	17	Zn	24	Sr	31	Cd	38	Ba & V	45	Hg	52
Bo	4	Al	11	Cr	19	Y	25	Ce & La	33	U	40	Ta	46	Tl	53
C	5	Si	12	Ti	18	In	26	Zr	32	Sn	39	W	47	Pb	54
N	6	P	13	Mn	20	As	27	Di & Mo	34	Sb	41	Nb	48	Bi	55
O	7	S	14	Fe	21	Se	28	Ro & Ru	35	Te	43	Au	49	Th	56

Problems:

Put two elements in the same place on several occasions

Some element properties were not a very good match (transition metals)

1864-1870 – J. Lothar Meyer: discovered periodicity of elements arranged by atomic weight

Meyer's work ignored at first because:

- Mendeleev published first
- Mendeleev was more careful
- Mendeleev made accurate predictions of missing elements

1869 – Dmitiri Mendeleev: Periodic Law of the Elements

When arranged in order of increasing atomic weight, the properties of elements repeat periodically

- correctly predicted eight new elements
- incorrectly suggested that improvements in atomic weights would rectify places on his table where the weights of elements appeared out of increasing order (see Te and I)

1911 – Henry Gwyn Jeffreys Moseley: Modern Periodic Law of the Elements

- used x-rays to count the number of protons in the nucleus
- resolved the atomic weight anomalies on the table

When arranged in order of increasing *atomic number*, the properties of elements repeat periodically