

AP CHEM TOPIC#10

STRUCTURE (ATOMIC) NOTES

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ELECTROMAGNETIC RADIATION SECTION #1

- One of the means E travels through space
- All EM (electromagnetic) waves travel at the speed of light
- Waves have three primary characteristics
 - Wavelength
 - Symbol: λ (lower case lambda)
 - Label: meters
 - The distance between two consecutive peaks or troughs in a wave
 - Frequency
 - Symbol: ν (lower case nu)
 - Label: 1/sec or Hz (Hertz)
 - Number of waves (cycles) per second
 - Speed
 - Light: $c = 2.9979 \times 10^8 \text{ m/s}$
 - Label: m/s
- Speed of light equation

- $c = \lambda \nu$
 - λ (lambda) must be in meters when used in equation
- **Sample Exercise #1** – Frequency of Electromagnetic Radiation
The brilliant colors seen in fireworks are due to the emission of light with wavelengths around 650nm when strontium salts such as $\text{Sr}(\text{NO}_3)_2$ and SrCO_3 are heated. Calculate the frequency of red light of wavelength 650nm. (Ans: $4.61 \times 10^{14} \text{Hz}$)

THE NATURE OF MATTER SECTION #2

- Max Planck
 - Energy gained/lost in whole number multiples of $h\nu$
 - h is Planck's constant
 - $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$
 - Energy can be quantized (packets)
 - Packet called a quantum
 - **Sample Exercise #2** – The Energy of a Photon
The blue color in fireworks is often achieved by heating copper (I) chloride (CuCl) to about 1200°C . Then the compound emits blue light having a wavelength of 450nm. What is the increment of energy (the quantum) that is emitted at $4.50 \times 10^2 \text{ nm}$ by CuCl ? (Ans: $4.41 \times 10^{-19} \text{ J}$)
- Albert Einstein
 - Proposed EM radiation is quantized
 - Stream of particles called "photons"
 - $E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$
 - $E = mc^2$
 - Energy has mass
 - By substitution, $m = \frac{h}{\lambda c}$
- Arthur Compton
 - Experiment
 - Collisions between x-rays and electrons
 - Showed photons do exhibit apparent mass
 - No rest mass (classical sense)
 - Only in relative
 - Discovered neutron
- Summary of work by Einstein and Planck
 - Energy – quantized – discrete units called quanta
 - Dual nature of light
 - Wave
 - Stream of photons
- Louis de Broglie
 - Equation used to calculate the wavelength of a particle
 - $\lambda = h/mv$

- **Sample Exercise #3** – Calculations of Wavelength
Compare the wavelength for an electron ($m = 9.11 \times 10^{-31} \text{kg}$) traveling at a speed of $1.0 \times 10^7 \text{m/s}$ with that for a ball ($m = 0.10 \text{kg}$) traveling at 35m/s . (Ans: $\lambda_e = 7.27 \times 10^{-11} \text{m}$, $\lambda_b = 1.9 \times 10^{-34} \text{m}$)
 - Note the λ for the ball is very short while the λ for the electron is on the same order as the spacing between the atoms in a typical crystal.
 - Why is this important?
 - To prove, shoot electrons at a nickel crystal, observed a diffraction pattern (wave property)
 - Matter obeys, but λ way too small to test
 - All matter exhibits both particulate and wave properties

THE ATOMIC STRUCTURE OF HYDROGEN SECTION #3

- Continuous spectrum
 - White light passed through a prism
 - Breaks light into representative colors
- Hydrogen emission spectrum
 - Line spectrum
 - Certain lines of light at specific wavelengths/frequencies/energies
 - Allows only certain energies for the electron in an hydrogen atom
 - Quantized
 - $\Delta E = h\nu = hc/\lambda$

THE BOHR MODEL SECTION #4

- Quantum model for hydrogen
 - E moves around nucleus in discrete orbits
 - $E = -2.178 \times 10^{-18} \text{J} (Z^2/n^2)$, n integer and Z is nuclear charge
 - (-) sign means E of electron bound to nucleus is lower than at an infinite range
 - $\Delta E = E_{(\text{of final state})} - E_{(\text{of initial state})}$
 - **Sample Exercise #4** – Energy Quantization in Hydrogen
Calculate the E required to excite the hydrogen electron from level $n = 1$ to level $n = 2$. Also calculate the wavelength of light that must be absorbed by a hydrogen atom in its ground state to reach this excited state. (Ans: $1.216 \times 10^{-7} \text{m}$)
- Two important points about Bohr model:
 1. Correctly fits the quantized energy levels of the hydrogen atom and postulates only allowed circular orbits for electron.
 2. When electron is at an infinite distance from the nucleus, energy state is zero (zero-energy state). As the electron becomes more bound to the nucleus (closer) the energy becomes more negative in relation to zero-energy state. In other words, the closer the electron is brought to the nucleus the more E is released from system.
 - General equation for the electron moving one level (n_{initial}) to another level (n_{final})

$$\Delta E = E_{\text{final}} - E_{\text{initial}} = -2.178 \times 10^{-18} \text{J} (1/n_f^2 - 1/n_i^2)$$

- **Sample Exercise #5** – Electron Energies

Calculate the E required to remove an electron from a hydrogen atom in its ground state.

(Ans: $2.178 \times 10^{-18} \text{J}$)

- Ultimately, the Bohr model only works for hydrogen, but is significant because it showed the observed quantization of E in atoms could be explained using simple assumptions.

THE QUANTUM MECHANICAL VIEW OF THE ATOM SECTION #5

- Werner Heisenberg (1901-1976)
 - Heisenberg Uncertainty Principle
 - Fundamental limitation to just how precisely we can know both the position and momentum of a particle at a given time.
 - If we measure where it is we will not know for certain where it will go.
 - On the quantum level, any method of observation will change the system.
- Louis de Broglie (1892-1987)
 - Electron exhibits wave properties
- Erwin Schrodinger (1887-1961)
 - Treated electron as a standing wave
 - Orbits with whole number of wavelengths, standing wave will fit
 - Orbits with destructive interference are not allowed
 - Wave function (ψ) – function of the coordinates x , y , and, z of the electron in 3D space
 - A specific wave function is called an orbital
 - Electron density (probability) – 90% within a certain region of space
 - Picture an orbital as a 3-D electron density map
- The Physical Meaning of a Wave Function
 - The square of the wave function indicates the probability of finding an electron near a particular point in space.
 - Differentiates between two positions
 - Determines which position is most likely to find the electron
 - Gives no information concerning when the electron will be at either position or how it moves between positions.
 - Vagueness consistent with Heisenberg Uncertainty principle
 - Represented as a probability distribution
 - Uses color intensity to indicate the probability value near a given point in space.
 - Electron density map
 - Electron density same as electron probability
 - Atomic orbital
 - Orbital size
 - The radius of a sphere that encloses 90% of the total electron probability

QUANTUM NUMBERS SECTION #6

- n – principal quantum number (energy level)
 - 1,2,3,4,5,6, and 7

- Rows on a periodic table
 - Size and energy of the orbital
 - As $n \uparrow$'s, orbital becomes larger, e^- spends more time farther away from nucleus, and higher energy (because e^- is less tightly bound by nucleus)
- l – angular momentum number (shape of orbital)
 - sometimes called a subshell
 - Integers from 0 to $n-1$ for each value of n
 - For $l = 0$ s orbital
 - For $l = 1$ p orbital
 - For $l = 2$ d orbital
 - For $l = 3$ f orbital
- m_l – magnetic quantum number (orientation of orbital in space)
 - Integral values of l to $-l$
 - **Sample Exercise #6** – Electron Subshells
For the quantum level $n = 5$, determine the number of allowed subshells (different values of l), and give the designation of each. (Answer: 5 (s, p, d, f , and g))

ORBITAL SHAPES AND ENERGIES SECTION #7

- s -orbitals
 - spherical shaped and increase in size as n increases
 - single orbital
 - $3s > 2s > 1s$
- p -orbital
 - dumbbell shaped (two lobes separated by a region of zero probability (node))
 - three orbitals
 - start with $n = 2$
 - labeled according to xyz coordinates
 - $2p_x$ centered around x -axis
 - Surface area of orbital increase when n increases
- d -orbital
 - start with $n = 3$, no $2d$ or $1d$
 - five orbitals
- f -orbital
 - start with $n = 4$
 - seven orbitals
- For the hydrogen atom, all orbitals with the same value for n have the same energy
 - Degenerate
 - Ground state – lowest E state
 - Excited state – any E state higher than ground state
 - A summary of the hydrogen atom

- In the quantum (wave) model, the electron is viewed as a standing wave. This representation leads to a series of wave functions (orbitals) that describe the possible energies and spatial distributions available for electrons.
- In agreement with the Heisenberg Uncertainty Principle, the model cannot specify the detailed electron motions. Instead, the square of the wave function represents the probability distribution of the electron in that orbital. This allows us to picture orbitals in terms of probability distributions, or electron density maps.
- The size of an orbital is arbitrarily defined as the surface that contains 90% of the total electron probability.
- The hydrogen atom has many types of orbitals. In ground state, the single electron resides in the $1s$ orbital. The electron can be excited to higher-energy orbitals if energy is put into the atom.

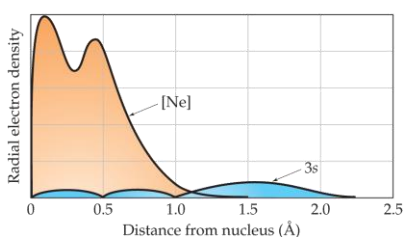
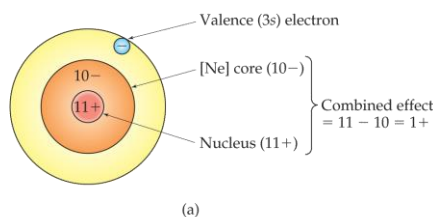
ELECTRON SPIN AND THE PAULI PRINCIPLE SECTION #8

- Fourth quantum number (m_s)
 - Samuel Goudsmit and George Uhlenbeck
 - Needed to account for the emission spectra of atoms
 - Indicated a magnetic moment
 - Two spins
 - Spin quantum number, m_s
 - $+\frac{1}{2}$ and $-\frac{1}{2}$
 - Wolfgang Pauli (1900-1958)
 - In a given atom, no two electrons can have the same 4 quantum numbers (n , l , m_l , and m_s) or two electrons with the same spin cannot occupy the same orbital
 - Two electrons per orbital with opposite spins
 - Pauli exclusion principle

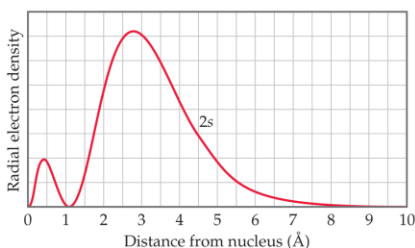
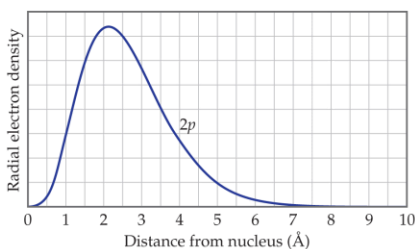
POLYELECTRONIC ATOMS SECTION #9

- Atoms with more than 1 electron
- Consider helium
 - Three energy contributions
 - (1) KE of moving electrons
 - (2) PE due to attraction between electrons and nucleus
 - (3) PE due to repulsions between electrons
 - Can be described in terms of the quantum mechanical model
 - On the other hand, the Schrodinger wave equation cannot be solved exactly
 - Difficulty in dealing with electron-electron repulsion
 - Due to the inability to determine the pathway of electrons
 - Cannot calculate the repulsions exactly
 - Called electron correlation problem
 - All polyelectronic atoms
 - Use approximation

- Treat each electron as if it were moving in a field of charge that is the result of the nuclear attraction and the average repulsions of all the other electrons
- e.g. - sodium atom
 - Outermost electron
 - Attracted to nucleus
 - Repulsions from inner 10 electrons
 - Net effect, electron not held as tightly
 - Shielded (screened) from nucleus by inner electrons
 - This picture polyelectric atoms leads to hydrogen like orbitals for these atoms
 - All orbitals in a principal quantum number are not of the same energy (unlike H)
 - $E_{ns} < E_{np} < E_{nd} < E_{nf}$
 - $4s < 4p < 4d < 4f$
 - Electrons prefer the order $s, p, d, \& f$
 - Why?
 - Probability profiles show s electrons penetrate shield more effectively than $p, d,$ or f electrons



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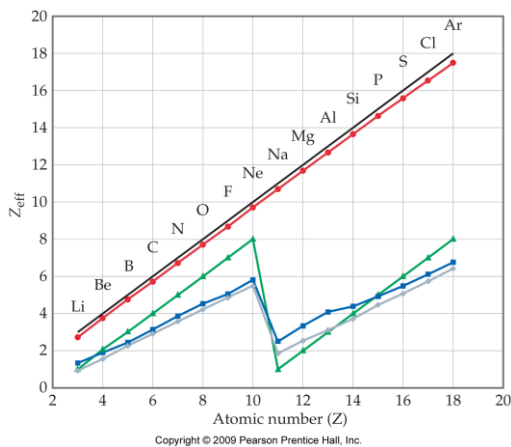


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(a) The effective nuclear charge experienced by the valence electron in sodium depends mostly on the 11+ charge of the nucleus and the 10- charge of the neon core. If the neon core were totally effective in shielding the valence electron from the nucleus, then the valence electron would experience an effective nuclear charge of 1+.

(b) The 3s electron has some probability of being inside the Ne core. Because of this "penetration," the core is not completely effective in screening the 3s electron from the nucleus. Thus, the effective nuclear charge experienced by the 3s electron is somewhat greater than 1+.

The radial probability function for the 2s orbital of the hydrogen atom (red curve) shows a "bump" of probability close to the nucleus, whereas that for the 2p orbital (blue curve) does not. As a result, an electron in the 2s orbital for a many-electron atom "sees" more of the nuclear charge than does an electron in the 2p orbital—the effective nuclear charge experienced by the 2s electron is greater than that for the 2p electron. This difference leads to our observation that in a many-electron atom the orbitals for a given n value increase in energy with increasing l value; that is, ns is lower in energy than np , which is lower in energy than nd .



This figure shows the variation in effective nuclear charge, Z_{eff} , as felt by the innermost electrons, those in the $1s$ subshell (red circles), which closely tracks the increase in nuclear charge, Z (black line). The Z_{eff} felt by the outermost valence electrons (blue squares) not only is significantly smaller than Z , it does not evolve linearly with increasing atomic number. It is also possible to estimate Z_{eff} using certain rules. The values shown as green triangles were obtained by assuming the core electrons are completely effective at screening and the valence electrons do not screen the nuclear charge at all. The values shown as gray diamonds were estimated using Slater's rules.

THE HISTORY OF THE PERIODIC TABLE SECTION #10

- Johann Dobereiner (1780 – 1849)
 - Triads
 - Cl, Br, and I
 - Average 1st and last to find properties of middle
 - Very limited
- John Newlands
 - Octaves
 - Certain properties occur at intervals of 8
 - Similar to musical scale
- Dmitri Mendeleev (1834-1907) & Julius Lothar Meyer (1830-1895)
 - Arranged based on chemical properties of known elements
 - Mendeleev given most of the credit
 - Used atomic masses
 - Due to predictive power of table
 - Predicted scandium, gallium, and germanium (*ekasilicon*)
 - Corrected atomic mass for indium
 - From 76 (InO) to 113 (In₂O₃)
 - Also for beryllium and uranium
- John Mosely
 - Worked with Rutherford
 - Discovered atomic number
 - Unique for each element
 - Number of protons on nucleus
 - Switched atomic number for atomic masses

TABLE 7.1 ■ Comparison of the Properties of Eka-Silicon Predicted by Mendeleev with the Observed Properties of Germanium

Property	Mendeleev's Predictions for Eka-Silicon (made in 1871)	Observed Properties of Germanium (discovered in 1886)
Atomic weight	72	72.59
Density (g/cm ³)	5.5	5.35
Specific heat (J/g-K)	0.305	0.309
Melting point (°C)	High	947
Color	Dark gray	Grayish white
Formula of oxide	XO ₂	GeO ₂
Density of oxide (g/cm ³)	4.7	4.70
Formula of chloride	XCl ₄	GeCl ₄
Boiling point of chloride (°C)	A little under 100	84

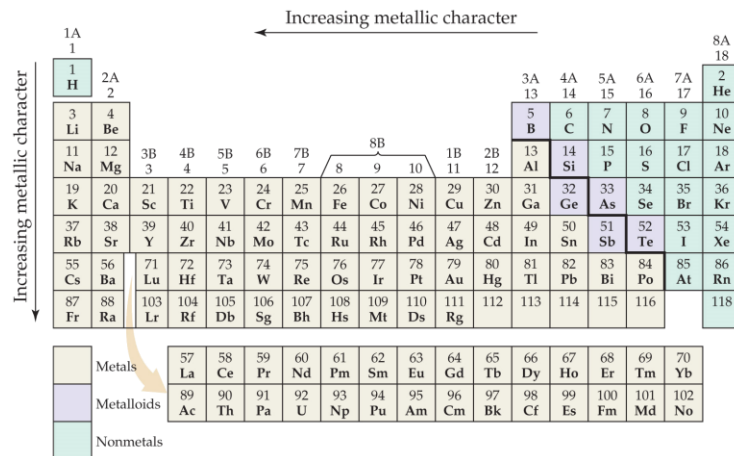
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- Corrected the arrangement of several elements

- Te and I

Ancient Times	1735–1843	1894–1918
Middle Ages–1700	1843–1886	1923–1961
	1965–	

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TABLE 7.3 ■ Characteristic Properties of Metals and Nonmetals

Metals	Nonmetals
Have a shiny luster; various colors, although most are silvery	Do not have a luster; various colors
Solids are malleable and ductile	Solids are usually brittle; some are hard, some are soft
Good conductors of heat and electricity	Poor conductors of heat and electricity
Most metal oxides are ionic solids that are basic	Most nonmetal oxides are molecular substances that form acidic solutions
Tend to form cations in aqueous solution	Tend to form anions or oxyanions in aqueous solution

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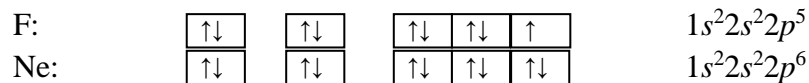
THE AUFBAU PRINCIPLE AND THE PERIODIC TABLE SECTION #11

- German for “building up”
 - As protons are added to the nucleus to build up the elements, electrons are added to the hydrogen like orbitals
 - Use Z to represent atomic number
 - Below are the orbital diagrams and electron configurations for the first 8 elements.

	1s	2s	2p	
H:	\uparrow			$1s^1$
He:	$\uparrow\downarrow$			$1s^2$
Li:	$\uparrow\downarrow$	\uparrow		$1s^2 2s^1$
Be:	$\uparrow\downarrow$	$\uparrow\downarrow$		$1s^2 2s^2$
B:	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow	$1s^2 2s^2 2p^1$

- Hund’s Rule
 - Lowest E for an atom is one having the maximum number of unpaired electrons allowed by the Pauli exclusion principal
 - By convention the unpaired electrons are represented as having parallel spins

C:	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow \uparrow \square	$1s^2 2s^2 2p^2$
N:	$\uparrow\downarrow$	$\uparrow\downarrow$	\uparrow \uparrow \uparrow	$1s^2 2s^2 2p^3$
O:	$\uparrow\downarrow$	$\uparrow\downarrow$	$\uparrow\downarrow$ \uparrow \uparrow	$1s^2 2s^2 2p^4$



- Valence electrons
 - Outer *s* and *p* electrons
 - For group I metals 1 valence electrons ns^1
 - For group II metals 2 valence electrons ns^2
 - For transition metals 2 valence electrons $ns^2(n-1)d^{1-10}$
 - For first *p* metals 3 valence electrons $ns^2(n-1)^{10}np^1$
 - For last *p* elements 8 valence electrons $ns^2(n-1)^{10}np^6$
 - Magic number for stability in an atom/ion is 8 valence electrons (ns^2np^6) called octet rule
- Electron configurations
 - Use the above rules for writing (Hund's, Pauli, and Aufbau)
 - Sequence goes as follows from lowest *E* level to highest *E* level

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^{14} 6d^{10}$$
 - Determine *Z* – atomic number of the element
 - This is the number of e^- needed to place in configuration
 - Use powers above to find where your *Z* number falls
 - Example: Fe *Z*=26 so you last e^- is in $3d^{10}$.
 - So the electron configuration for Fe is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
 - Add up the powers to get *Z*
 - $2+2+6+2+6+2+6 = 26$
 - Exceptions to Aufbau principle
 - These are elements that “borrow” from one subshell to fill or half-fill another half shell
 - Cr, Mo, and W borrow from their *s* to half fill their *d*
 - Cu, Ag, and Au borrow from their *s* to fill their *d*
 - Half filled or filled subshells give the atom/ion more stability
 - Abbreviated electron configuration (Noble gas configuration)
 - Use the symbol of a Noble gas to represent a string of filled orbitals

[Ne]	represents	$1s^2 2s^2 2p^6$
[Ar]	represents	$1s^2 2s^2 2p^6 3s^2 3p^6$
[Kr]	represents	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

 And so on
 - Add the subshells with correct number of electrons in the row the element is in until the element is reached
 - Abbreviated configuration for Fe: [Ar] $4s^2 3d^6$
 - **Sample Exercise #7** – Electron Configurations

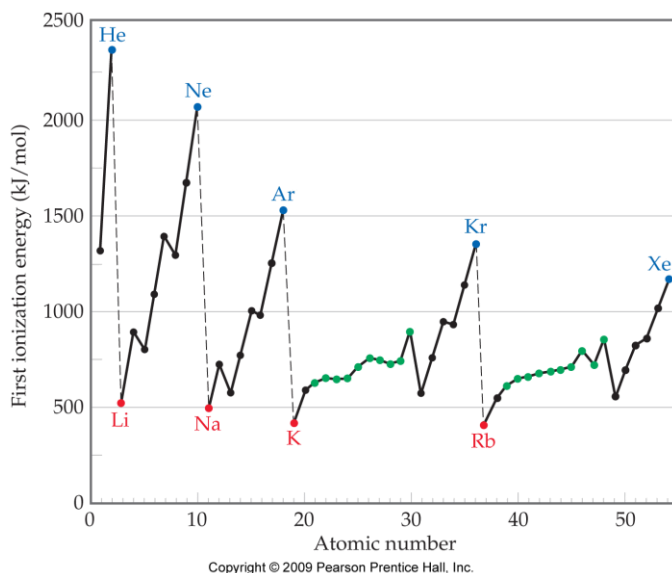
Write the electron configuration, orbital diagram, abbreviated configuration, valence electrons, and electron dot diagrams for the following.

(a) sulfur
(b) cadmium
(c) hafnium
(d) radium

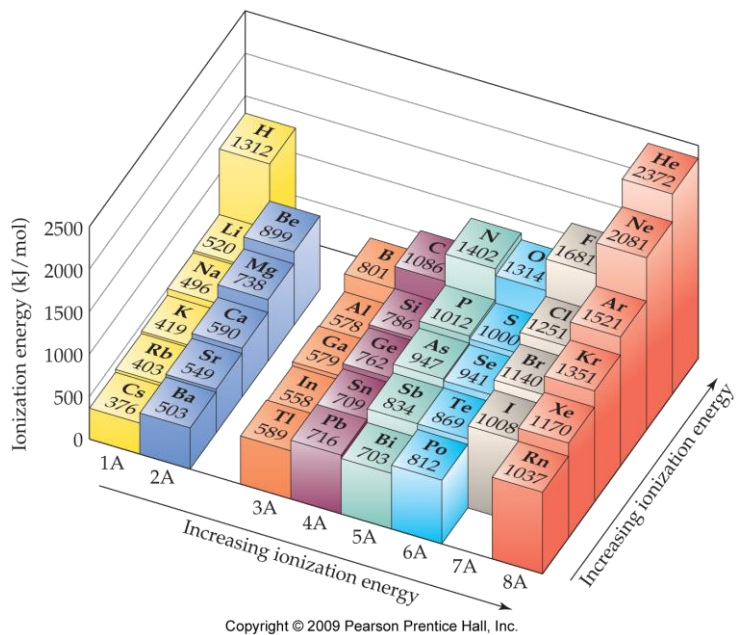
PERIODIC TRENDS IN ATOMIC PROPERTIES SECTION #12

- Ionization energy
 - E needed to remove an electron from a gaseous atom or ion
 - Each electron removed increases the I_E .

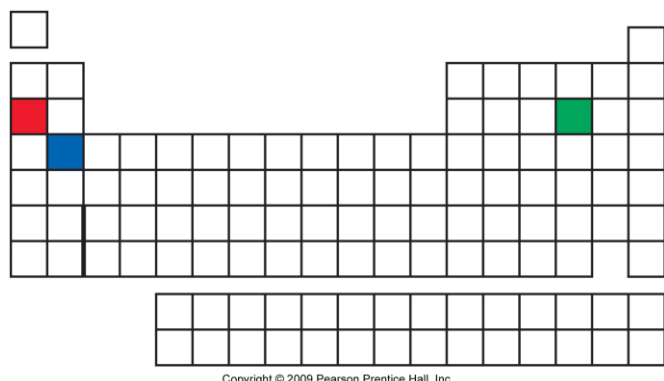
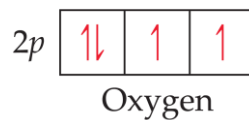
$\text{Al}(g) \rightarrow \text{Al}^+ + e^-$	$I_1 = 580 \text{ kJ/mol}$
$\text{Al}(g) \rightarrow \text{Al}^{2+} + e^-$	$I_2 = 1815 \text{ kJ/mol}$
$\text{Al}(g) \rightarrow \text{Al}^{3+} + e^-$	$I_3 = 2740 \text{ kJ/mol}$
$\text{Al}(g) \rightarrow \text{Al}^{4+} + e^-$	$I_4 = 11,600 \text{ kJ/mol}$
 - First ionization energy (I_1) is the E needed to remove the highest E electron of an atom
 - Why does it take more E to remove subsequent electrons? _____
 - Trend:
 - across a period from left to right the I_E increases
 - down a group I_E decreases
 - **Sample Exercise #8** – Trends in Ionization Energies
The first ionization E for phosphorus is 1060kJ/mol, and that for sulfur is 1005kJ/mol. Why?
 - **Sample Exercise #9** – Ionization Energies
Consider atoms with the following electron configurations:
(a) $1s^22s^22p^6$ (b) $1s^22s^22p^63s^1$ (c) $1s^22s^22p^63s^2$
What atom has the largest I_1 , and which one has the smallest I_2 ? Explain your choices.



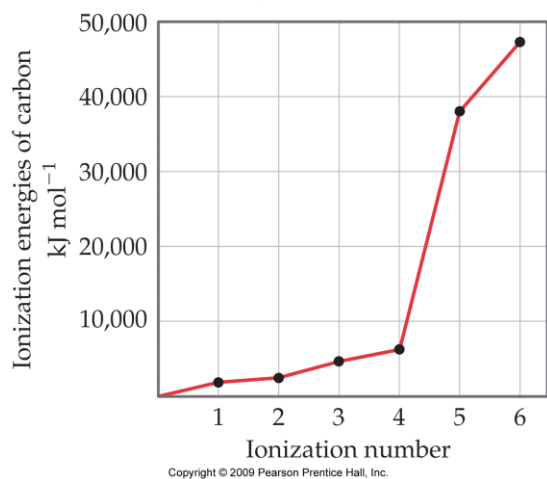
The red dots mark the beginning of a period (alkali metals), the blue dots mark the end of a period (noble gases), and the black dots indicate other representative elements. Green dots are used for the transition metals.



The presence of a fourth electron in the $2p$ orbitals of oxygen leads to an extra repulsion associated with putting two electrons in a single orbital. This repulsion is responsible for the lower first ionization energy of oxygen.



Three elements are indicated on the periodic table in the figure. Based on their locations, predict the one with the largest second ionization energy.



Explain the variation in ionization energies of carbon, as displayed in the graph.

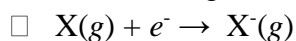
TABLE 7.2 Successive Values of Ionization Energies, I , for the Elements Sodium through Argon (kJ/mol)

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4562	(inner-shell electrons)				
Mg	738	1451	7733				
Al	578	1817	2745	11,577			
Si	786	1577	3232	4356	16,091		
P	1012	1907	2914	4964	6274	21,267	
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018
Ar	1521	2666	3931	5771	7238	8781	11,995

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- Electron affinity

- The E change associated with the addition of an e^- to a gaseous atom (making it an anion).
- Associated with the production of a negative ion
- The more apt an atom is towards forming a negative ion the more E is released during the formation of the negative ion



- Positive sign on E means endothermic
- Negative sign on E means exothermic
 - Measuring change in E
- Trend
 - Electron affinities become more negative left to right across a period
 - Exceptions: N^- is not stable, while C^- and O^- are stable
 - Why?
 - Let's look at the electron configuration

$N^-(g)$	$1s^2 2s^2 2p^4$ (unstable)
$C^-(g)$	$1s^2 2s^2 2p^3$ (stable)
$O^-(g)$	$1s^2 2s^2 2p^4$ (stable)

 - Extra e^- has to occupy an orbital with an e^- already in it leading to repulsion
 - Extra e^- in unoccupied orbital
 - Extra e^- in an orbital with an electron, but has a greater nuclear charge pulling on e^- enough to overcome the repulsion between the e^- 's
 - Note: O^{2-} is not stable as a (g) too much repulsion among e^- 's
 - Stable in ionic compounds because of large attractions involved between positive ions and oxide ions
 - Down a group electron affinity becomes more positive
 - due to electrons being added at increasing distances from the nucleus
 - changes going down a group are relatively small
 - Take the halogens for example:
 - A_e for halogens

F	-327.8 kJ/mol
Cl	-348.7 kJ/mol
Br	-324.5 kJ/mol
I	-295.2 kJ/mol
 - Cl thru I show the trend, but F to Cl does not
 - Difference is due to F small $2p$ orbitals
 - e^- 's are very close, large $e^- - e^-$ repulsions
 - other halogens have larger p orbitals mitigating the repulsion effect

H -73							He > 0
Li -60	Be > 0	B -27	C -122	N > 0	O -141	F -328	Ne > 0
Na -53	Mg > 0	Al -43	Si -134	P -72	S -200	Cl -349	Ar > 0
K -48	Ca -2	Ga -30	Ge -119	As -78	Se -195	Br -325	Kr > 0
Rb -47	Sr -5	In -30	Sn -107	Sb -103	Te -190	I -295	Xe > 0
1A	2A	3A	4A	5A	6A	7A	8A

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Electron affinities in for the *s*- and *p*-block elements in the first five rows of the periodic table. The more negative the electron affinity, the greater the attraction of the atom for an electron. An electron affinity > 0 indicates that the negative ion is higher in energy than the separated atom and electron.

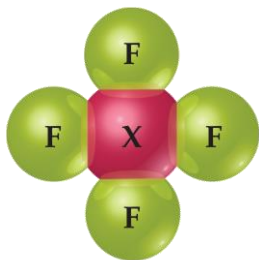
Cu -119	Zn > 0
Ag -126	Cd > 0
Au -223	Hg > 0

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The following table gives the electron affinities, in kJ/mol, for the group 1B and group 2B metals: (a) Why are the electron affinities of the group 2B elements greater than zero? (b) Why do the electron affinities of the group 1B elements become more negative as we move down the group? [*Hint*: Examine the trends in the electron affinity of other groups as we proceed down the periodic table.]

- Atomic radius

- Half the distance between the nuclei in a molecule consisting of identical atoms
 - Sometimes called covalent atomic radius
 - For nonmetallic atoms that do not form diatomic molecules the radius is estimated from their various covalent compounds.
 - For metals the radii is determined to be half way in between two atoms in a metallic crystal
 - Trends
 - Increases down a group
 - Outer e^- are “shielded” from the pull of the protons in the nucleus by the inner e^-
 - Also, each new period has a corresponding increase in orbital sizes for each new principal quantum number
 - Decrease going from left to right across a period
 - Explained by increasing effective nuclear charge (Z_{eff})
 - Valence e^- are drawn closer to the nucleus, decreasing the size of the atom
 - e^- s are in same E level across a period and the number of protons in nucleus is increasing so pull on e^- greater.



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An element X reacts with $F_2(g)$ to form the molecular product shown below. (a) Write a balanced equation for this reaction (do not worry about the phases for X and the product). (b) Do you think that X is a metal or nonmetal? Explain.

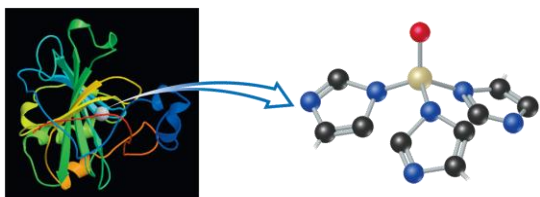
THE PROPERTIES OF A GROUP: THE ALKALI METALS SECTION #13

- It is the number and type of valence electrons that primarily determine an atom's chemistry
- Use electron configuration
 - Memorize exceptions: Cu and Cr
- Some groups have special names

Group 1	Alkali metals
Group 2	Alkaline earth metals
Group 3 – 12	Transition metals
Group 17	Halogens
Group 18	Noble gases
Row 4f	Lanthanides
Row 5f	Actinides (all are radioactive)
Metals	left of Zhenzil line
Metalloids	on the line (except Po and As)
Nonmetals	right of line
- Basic breakup is into metals and nonmetals
 - Most important chemical property to consider is metals (low ionization E) form cations and nonmetals (large ionization E) form anions
 - Most active metals: lower left of periodic table
 - Where ionization E are the lowest
 - Most active nonmetals: upper right of periodic table
 - High electron affinities
 - Highest ionization E
- We will use the Alkali metals group to illustrate
 - Lithium, sodium, potassium, rubidium, cesium, and francium
 - Fr only occurs very minimally in nature, so it is not considered

Properties of Five Alkali Metals

Element	Val elect config	Density @ 25°C (g/cm ³)	mp (°C)	bp (°C)	I_1 (kJ/mol)	Atomic Radius (pm)	Ionic (M ⁺) Radius (pm)
Li	2s ¹	0.53	180	1330	520	152	60
Na	3s ¹	0.97	98	892	495	186	95
K	4s ¹	0.86	64	760	419	227	133
Rb	5s ¹	1.53	39	668	409	247	148
Cs	6s ¹	1.87	29	690	382	265	169



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The enzyme called carbonic anhydrase (left) catalyzes the reaction between CO_2 and water to form HCO_3^- . The ribbon represents the folding of the protein chain. The "active site" of the enzyme (represented by the ball-and-stick model) is where the reaction occurs. (H atoms have been excluded from this model for clarity.) The red sphere represents the oxygen of a water molecule that is bound to the zinc ion (gold sphere) at the center of the active site. The water molecule is replaced by CO_2 in the reaction. The bonds coming off the five-member rings attach the active site to the protein.



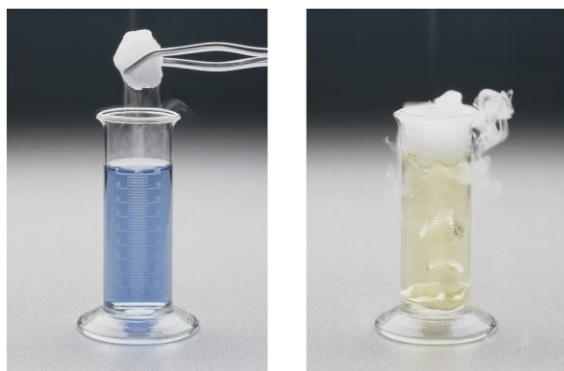
NiO

(a)

(b)

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(a) Nickel oxide (NiO), nitric acid (HNO_3), and water. (b) NiO is insoluble in water, but reacts with HNO_3 to give a green solution of the salt $\text{Ni}(\text{NO}_3)_2$.



(a)

(b)

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(a) The water has been made slightly basic and contains a few drops of bromthymol blue, an acid–base indicator that is blue in basic solution. (b) Upon the addition of a piece of solid carbon dioxide, $\text{CO}_2(s)$, the color changes to yellow, indicating an acidic solution. The mist is due to water droplets condensed from the air by the cold CO_2 gas.



(a)

(b)

(c)

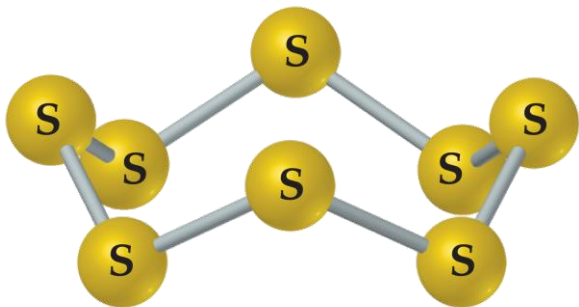
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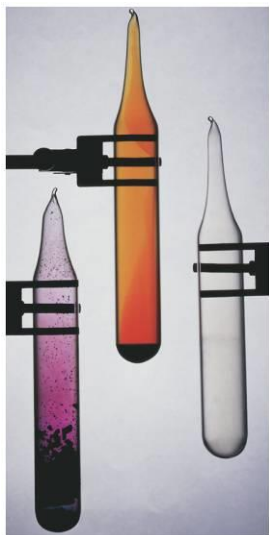
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- a) lithium
- b) sodium
- c) potassium
- d) calcium



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At room temperature, the most common allotropic form of sulfur is an eight-member ring, S_8 .



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All three of these—from left to right, iodine (I_2), bromine (Br_2), and chlorine (Cl_2)—exist as diatomic molecules.

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