Chem 103, Section F0F
Unit II - Quantum Theory and Atomic Structure
Lecture 9

- The periodic trends for three key properties of the elements
- How the electronic structure of the elements affects their chemical reactivity.

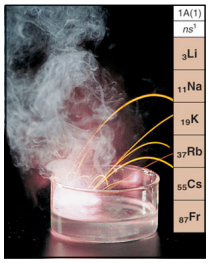
Lecture 9 - Electron Configuratioin

- Reading in Silberberg
  - Chapter 8, Section 4 *Trends in Three Key atomic Properties*
  - Chapter 8, Section 5 *Atomic Structure and Chemical Reactivity*

Lecture 9 - Introduction

In lab we compared the chemical reactivity for elements at different locations on the periodic table.

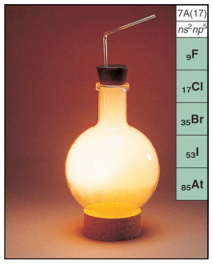
- We defined chemical reactivity for metals as a willingness to give up electrons.



Lecture 9 - Introduction

In lab we compared the chemical reactivity for elements at different locations on the periodic table.

- We defined chemical reactivity for nonmetals as a willingness to accept electrons.



Lecture 9 - Introduction

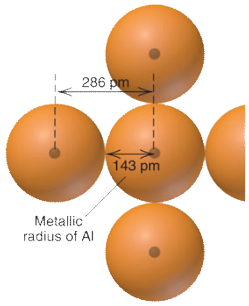
Today we will discuss the question: *What does the electron configuration of a element tell us about the its physical and chemical properties?*

- We will focus on trends for three properties:
  - Trends in atomic size
  - Trends in ionization energy
    - The energy required to remove an electron form an atom in the gaseous state.
  - Trends in electron affinity
    - The energy change when an electron is added to an atom in the gaseous state.

Lecture 9 - Trends in Atomic Size

There are different ways to measure size, we will use the following definitions:

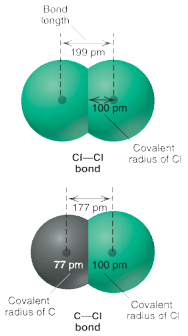
- For metals
  - 1/2 the distance between one atom and its neighbor in a metallic crystal



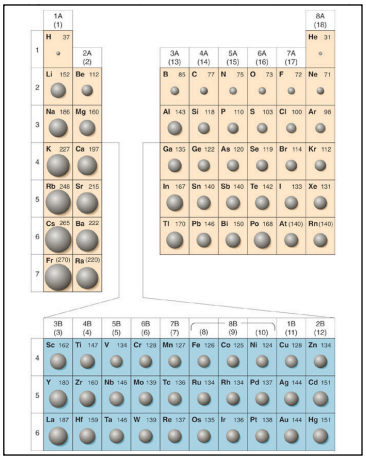
Lecture 9 - Trends in Atomic Size

There are different ways to measure size, we will use the following definitions:

- For nonmetals, which typically form covalent bonds with other nonmetals
  - 1/2 the distance between one one atom and the neighbor it is bonded to.



Atomic Size



Lecture 9 - Trends in Atomic Size

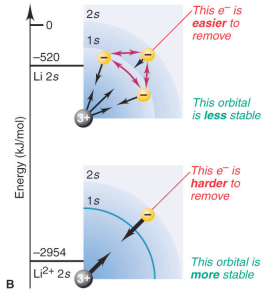
Trends among the main-group elements:

- As *n* (the period) increases, the radius increases.
  - Each shell places the electrons further from the nucleus.
  - There are also more inner electrons, which shield the outer electrons, lowering the Z_{eff}.

Lecture 9 - Many-Electron Atoms

The effect of electron repulsions and shielding on orbital energy

- The electrons feel not only the attraction of the nucleus, but also the repulsion of the other electrons.
- Shielding by inner elecrons greatly lowers the effective nuclear charge (Z_{eff}).



Electron shielding

Lecture 9 - Trends in Atomic Size

Trends among the main-group elements:

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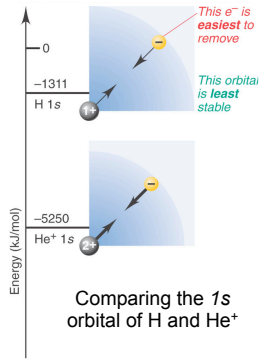
- As you move across a period, the size decreases.
  - The added electrons are in the same shell, but the Z_{eff} increases, pulling these electrons closer to the nucleus.

Lecture 9 - Many-Electron Atoms

The effect of nuclear charge (Z)

- Higher nuclear charge lowers the energy of an energy level.



Comparing the *1s* orbital of H and He^+

Lecture 9 - Trends in Atomic Size

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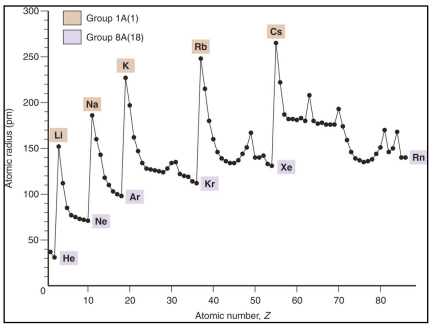
- As you move across a period, the size decreases.
  - The added electrons are in the same shell, but the Z_{eff} increases, pulling these electrons closer to the nucleus.
- We will focus on the main group elements and not worry about the subtleties of the transitions metals (pp318-320).

Lecture 9 - Trends in Atomic Size

Moving through the periodic table, we see the atomic size is periodic



Lecture 9 - Trends in Ionization Energy

The ionization energy (IE) is the energy required to completely remove 1 electron from an atom in its gaseous state.

- Because energy is entering the atom (system), it is positive.

$$Atom_{(g)} \longrightarrow Ion^{+}_{(g)} + e^{-} \quad \Delta E = IE_1 > 0$$

- The energy for removing a second electron (IE_2) is always greater than the energy for removing the first electron (IE_1)

$$Ion^{+}_{(g)} \longrightarrow Ion^{+2}_{(g)} + e^{-} \quad \Delta E = IE_2 > IE_1$$

Lecture 9 - Trends in Ionization Energy

- For the hydrogen atom, the ionization energy (IE) can be calculated by combining *Rydberg's equation* with *Planck's equation* and finding the energy it takes to move the electron from $n = 1$ to $n = \infty$.

$$\frac{1}{\lambda} = R\left(\frac{1}{n_1^2} - \frac{1}{n_\infty^2}\right) \text{ (Rydberg's Equation)}$$

$$E = h\nu \text{ (Planck's Equation)}$$

$$= \frac{hc}{\lambda}$$

$$= hcR\left(\frac{1}{n_1^2} - \frac{1}{n_\infty^2}\right)$$

Lecture 9 - Trends in Ionization Energy

- For the hydrogen atom, the ionization energy (IE) can be calculated by combining *Rydberg's equation* with *Planck's equation* and finding the energy it takes to move the electron from $n = 1$ to $n = \infty$.

$$E = hcR\left(\frac{1}{n_1^2} - \frac{1}{n_\infty^2}\right)$$

$$= \left(6.626 \times 10^{-34}\ \text{J•s}\right)\left(2.998 \times 10^{8}\ \text{m/s}\right)\left(1.097 \times 10^{7}\ \text{1/m}\right)\left(\frac{1}{1^2} - \frac{1}{\infty}\right)$$

$$= 2.179 \times 10^{-18}\ \text{J}$$

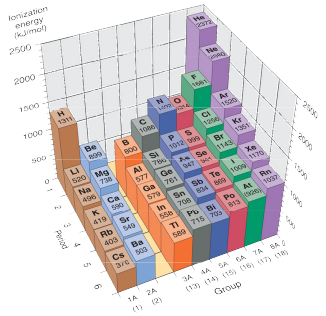
For 1 mole of atoms

$$= \left(2.179 \times 10^{-18}\ \text{J}\right)\left(6.022 \times 10^{23}\ \text{1/mol}\right) = 1{,}312{,}00\ \text{J/mol}$$

$$E = 1{,}312\ \text{kJ/mol}$$

Lecture 9 - Trends in Ionization Energy

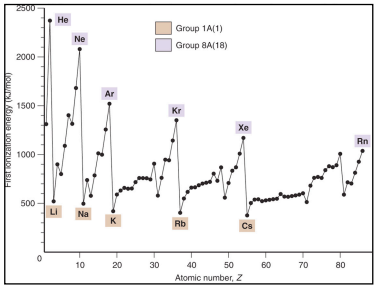
Atoms with low IE's tend to form cations (lose electrons),
Atoms with high IE's tend to form anions (gain electrons)



Lecture 9 - Trends in Ionization Energy

Atoms with low IE's tend to form cations (lose electrons),
Atoms with high IE's tend to form anions (gain electrons)

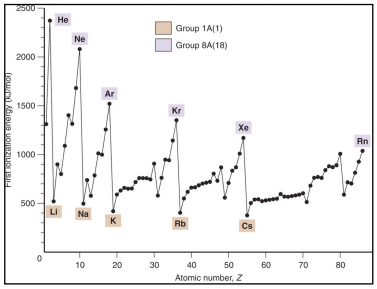
- Like the atomic sizes, the IE's vary periodically:



Lecture 9 - Trends in Ionization Energy

Trends

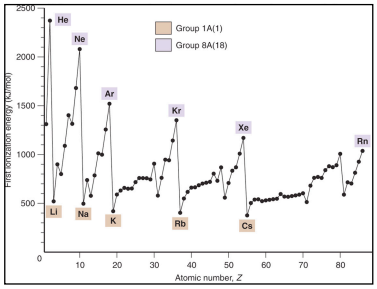
- As *n* increases ionization energies decrease
  - As the atoms get larger, the electrons removed are located further from the nucleus.



Lecture 9 - Trends in Ionization Energy

Trends

- From left to right in a period, ionization energies increase.
  - The Z_{eff} is increasing while the atoms are getting smaller, so the electrons are held more strongly.



Lecture 9 - Clicker Question 1

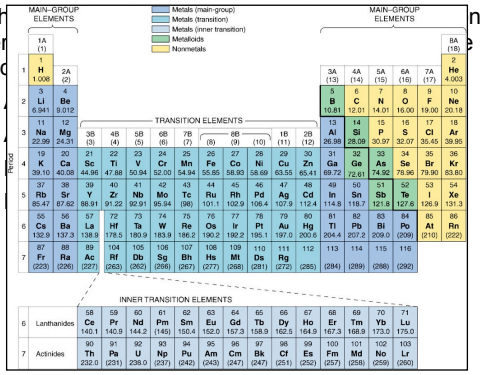
Rank th n (Ne) in order e periodic

A)

B)

C)

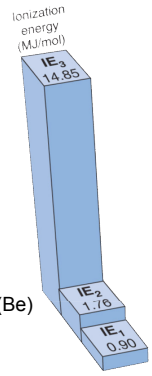
D)



Lecture 9 - Trends in Ionization Energy

Successive ionizations

- Successive ionizations show increasing ionization energies
  - This is because the resulting ion becomes increasingly positively charged.
- There is also a substantial jump in IE when dropping down to a lower energy level to remove one of the inner (core) electrons.



Berylium (Be)

Lecture 9 - Trends in Ionization Energy

Successive ionizations

- There is also a substantial jump in IE when dropping down to a lower energy level to remove one of the inner (core) electrons.

Table 8.5 Successive Ionization Energies of the Elements Lithium Through Sodium

Z	Element	Number of Valence Electrons	Ionization Energy (MJ/mol)*									
			IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
3	Li	1	0.52	7.30	11.81							
4	Be	2	0.90	1.76	14.85	21.01			CORE ELECTRONS			
5	B	3	0.80	2.43	3.66	25.02	32.82					
6	C	4	1.09	2.35	4.62	6.22	37.83	47.28				
7	N	5	1.40	2.86	4.58	7.48	9.44	53.27	64.36			
8	O	6	1.31	3.39	5.30	7.47	10.98	13.33	71.33	84.08		
9	F	7	1.68	3.37	6.05	8.41	11.02	15.16	17.87	92.04	106.43	
10	Ne	8	2.08	3.95	6.12	9.37	12.18	15.24	20.00	23.07	115.38	131.43
11	Na	1	0.50	4.56	6.91	9.54	13.35	16.61	20.11	25.49	28.93	141.37

*MJ/mol, or megajoules per mole = 10^3 kJ/mol.

Lecture 9 - Clicker Question 2

Element Q is in Period 3 and has the following ionization energies (in kJ/mol):

IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
577	1816	2744	11,576	14,829	18,375

Element Q is

A) Al
B) Be
C) B
D) C
E) N

Lecture 9 - Electron Affinity

Electron affinity (EA) is the energy that is released or gained when an atom adds an electron.

$$Atom_{(g)} + e^- \longrightarrow Ion^-_{(g)} \quad \Delta E = EA_1$$

- Usually energy is released when an electron is gained ($\Delta E < 0$), but there are some exceptions.
  - Unlike ionization energy, EA can be either positive or negative.

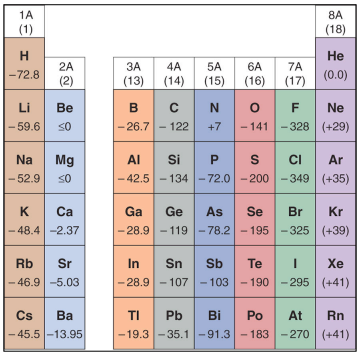
Lecture 9 - Electron Affinity

The electron affinity for adding a second electron (EA_2) is always positive because energy must be added to bind a negatively charged electron to an already negatively charged ion.

$$Ion^-_{(g)} + e^- \longrightarrow Ion^{2-}_{(g)} \quad \Delta E = EA_2 > 0$$

The trends for electron affinities are not as easy to predict as those for the ionization energies.

Lecture 9 - Electron Affinity



1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							He (0.0)
Li −59.6	Be ≤0	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)
Na −52.9	Mg ≤0	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)
K −48.4	Ca −2.37	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr −5.03	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)
Cs −45.5	Ba −13.95	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

Lecture 9 - Summarizing Trends

Trends in Ionization Energies (IE) and Electron Affinities (EA):

- *Reactive nonmetals*, in particular the halogens (7A)
  - Have high (positive) ionization energies (IE)
    - This makes it unfavorable for them to lose an electron
  - At the same time, they have highly negative electron affinities (EA)
    - This makes them more likely to gain an electron
  - Together, these properties make it more favorable for these elements to become negatively charge ions (**anions**).

Lecture 9 - Electron Affinity

Trends in Ionization Energies (IE) and Electron Affinities (EA):

- *Reactive metals,* in particular alkali and alkaline earth metals (1A and 2A)
  - Have low ionization energies (IE)
    - This makes it easy for them to lose an electron.
  - At the same time they have low electron affinities (EA)
    - This makes it less likely for them to gain an electron.
  - Together, these properties make it more favorable for these elements to become positively charged ions (**cations**).

Lecture 9 - Electron Affinity

Trends in Ionization Energies (IE) and Electron Affinities (EA):

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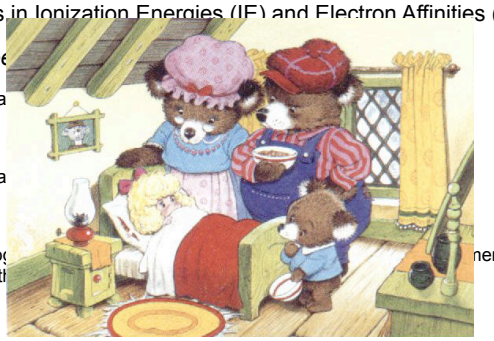
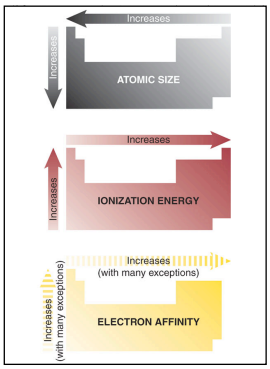


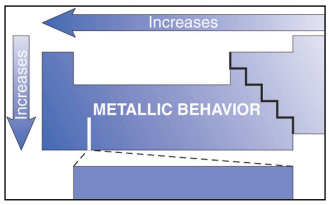
Lecture 9 - Summarizing Trends



Lecture 9 - Summarizing Trends

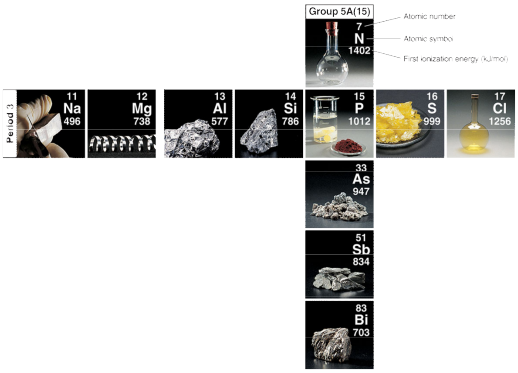
Atomic Structure and Chemical Reactivity

- Trends in Metallic *vs* Nonmetallic Behaviors
  - The further down and to the left on the periodic table, the more metallic an elements properties become.



- You can skip the discussion of the acid-base behavior of the element oxides (pp 326-327)

Lecture 9 - Summarizing Trends



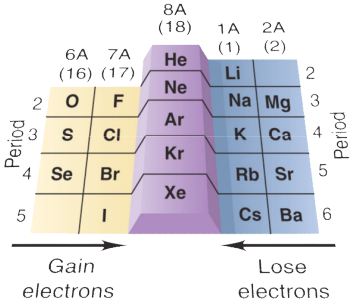
Lecture 9 - Summarizing Trends

- Trends in Metallic *vs* Nonmetallic Behaviors

Metal	*Nonmetal*
Shiny	*Dull*
High melting points	*Low melting points*
Electrical conductors	*Electrical insulators*
Thermal conductors	*Thermal insulators*
Malleable solids	*Brittle solids*
Lose electrons in chemical reactions	*Gain electrons in chemical reactions*

Lecture 9 -Properties of Monoatomic Ions

Main group elements (1A, 2A, 6A, 7A) lose (1A & 2A) or gain (6A & 7A) electrons to become isoelectronic (same number of electrons) with the nearest noble gas.



Lecture 9 -Properties of Monoatomic Ions

For larger metals, with *d* orbitals, only the *s* and p electrons are lost

- Giving them **pseudo-noble gas configurations**.

Lecture 9 - Clicker Question 3

Write the electron configuration for Aluminum (Al). Write the electron configuration for an aluminum ion (Al^{3+}).

Which noble gas is Al^{3+} isoelectronic with?

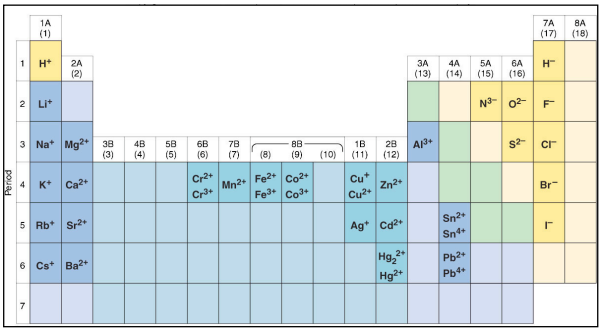
A) He
B) Ne
C) Ar
D) Kr
E) Xe

Lecture 9 -Properties of Monoatomic Ions

Transition metals have too many electrons to loose become isoelectronic with a noble gas

- They tend to loose their *ns* electrons, plus some of their (*n-1*)*d* electrons.
  - Typically the *ns* electrons are lost before the (*n-1*) electrons.
- They often form ions with different charge states one charge.

Lecture 9 -Properties of Monoatomic Ions

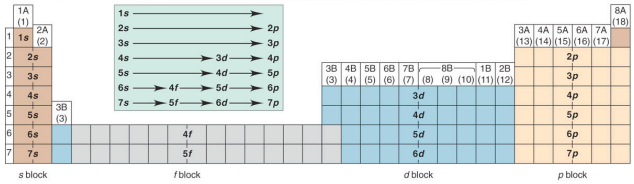


Lecture 9 -Properties of Monoatomic Ions

Generals rules for predicting the electronic configurations of ions

- For main group, *s-block* metals (1A & 2A)
  - Remove all electrons with the highest *n* value.
- For main-group, *p-block* metals (3A & 4A)
  - Remove *np* electrons before *ns* electrons.
- For transition, *d-block* metals
  - Remove *ns* electrons before (*n-1*)*d* electrons.
- For nonmetals (5A, 6A & 7A)
  - Add electrons to the *p* orbitals of the highest *n* value.

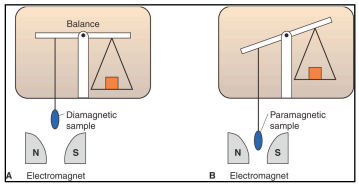
Lecture 9 -Properties of Monoatomic Ions



Lecture 9 -Properties of Monoatomic Ions

Magnetic properties of ions

- If an ion leaves some of the electrons unpaired, the ion will be attracted by a magnetic field
  - The ion is said to be **paramagnetic**.
- If an ion has all its electrons paired, the ion is weakly repelled by a magnetic field
  - The ion is said to be **diamagnetic**.



Lecture 9 - Clicker Question 4

Write a condensed electronic configuration for Cr^{3+}.

Is Cr^{3+} *paramagnetic*?

A) Yes
B) No

Lecture 9 - Clicker Question 5

Write a condensed electronic configuration for Cu^{+}.

Is Cu^{+} paramagnetic?

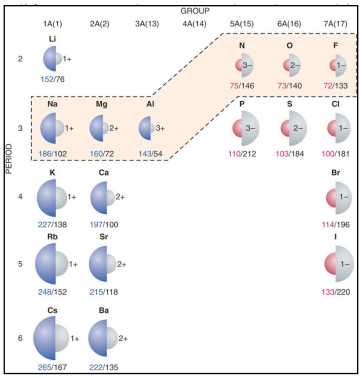
A) Yes
B) No

Lecture 9 -Properties of Monoatomic Ions

Ionic Size *vs* Atomic Size

- Cations are smaller than their parent atoms.
  - The outer electrons have been removed leaving only the inner (core) electrons.
    - The resulting net positive charge, along with reduced electron repulsion, leads to a contraction in the size.
- Anions are larger then their parent atoms.
  - Electrons are added to the outer level
    - The increased the electron repulsion leads to a swelling in the size.

Lecture 9 -Properties of Monoatomic Ions



Unit II - Up Next

Unit III - Combining Atoms to Make Compounds, Part I: Ionic Compounds

- Atomic properties and chemical bonds
- The ionic bonding model.

The End