Chapter 6 Part 4

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Core electrons vs. Valence Electrons

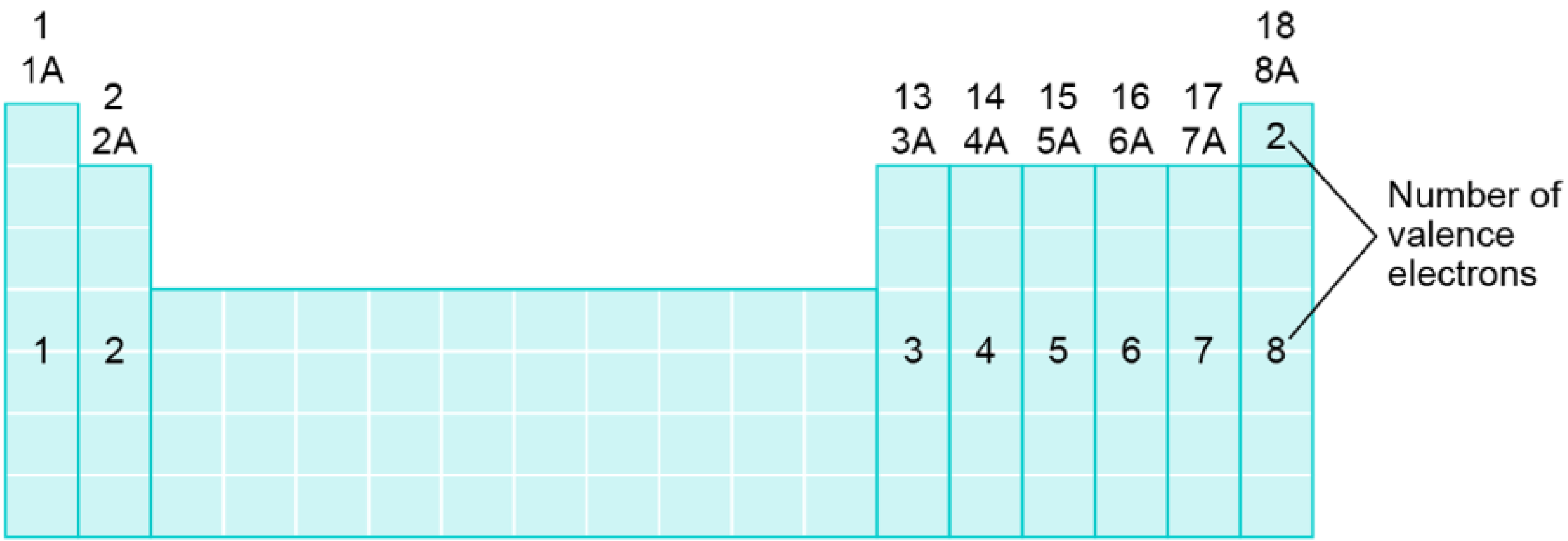
Core electrons

- Lower energy electrons, innermost electrons which shield the outermost electrons from the positive charge of the nucleus
- Rarely participate in bonding

Valence electrons

- Higher energy, outermost electrons that don't feel as much of the positive pull of the nucleus
- Regularly participate in bonding
- Equal to the sum of electrons in *s* and *p* orbitals of the highest occupied shell.

Valence Electrons



Valence Electrons

Identify the number of valence electrons in the following elements

A. Selenium
B. Lithium
C. Silicon

Atomic Radius

- Atomic radius size increases as one goes to the left and down the periodic table

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008																	2 He 4.0026
3 Li 6.94	4 Be 9.0122											5 B 10.81	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.06	17 Cl 35.45	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.38	31 Ga 69.723	32 Ge 72.630	33 As 74.922	34 Se 78.97	35 Br 79.904	36 Kr 83.798
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.95	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 *	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89-103 #	104 Rf (265)	105 Db (268)	106 Sg (271)	107 Bh (270)	108 Hs (277)	109 Mt (276)	110 Ds (281)	111 Rg (280)	112 Cn (285)	113 Nh (286)	114 Fl (289)	115 Mc (289)	116 Lv (293)	117 Ts (294)	118 Og (294)

* Lanthanide series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
# Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Atomic Radius

Which of the following shows the atoms C, Al, and Si in order of increasing atomic radius?

A. C, Al, Si
B. Si, Al, C
C. C, Si, Al
D. Al, Si, C

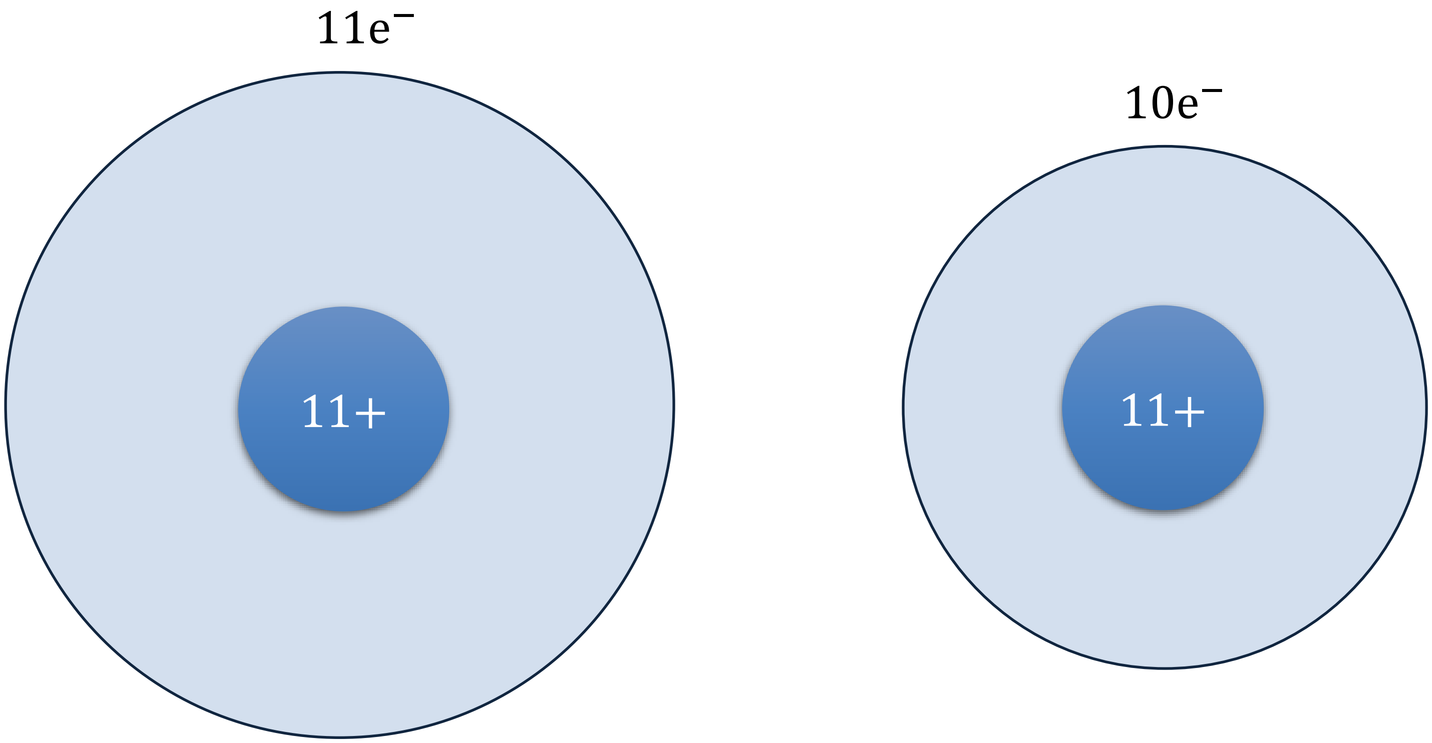
Ionic Size (Anions)

Fluorine Atom (F) → Fluoride Ion (F^-)

The added number of electrons increases the amount of electron repulsion. In response, the electron cloud gets bigger in order to accommodate the added repulsion

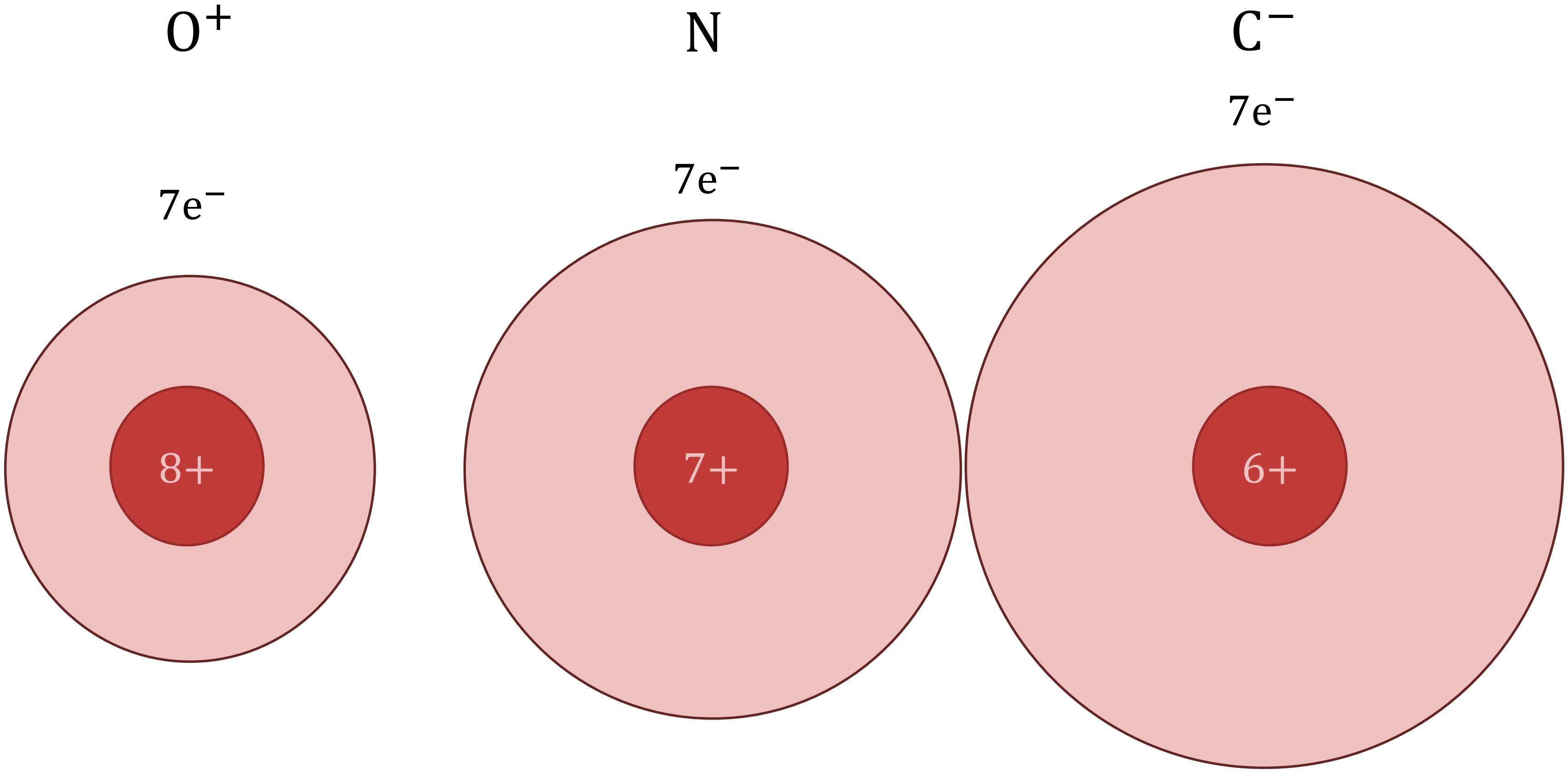
Ionic Size (Cations)

Sodium Atom (Na) → Sodium Ion (Na^+)



The decreased number of electrons decreases the amount of electron repulsion. In response, the electron cloud gets smaller.

Isoelectronic Series



Substances in isoelectronic series have the same number of electrons

Ionic Radius

Which of the following would you expect to have the largest ionic radius?

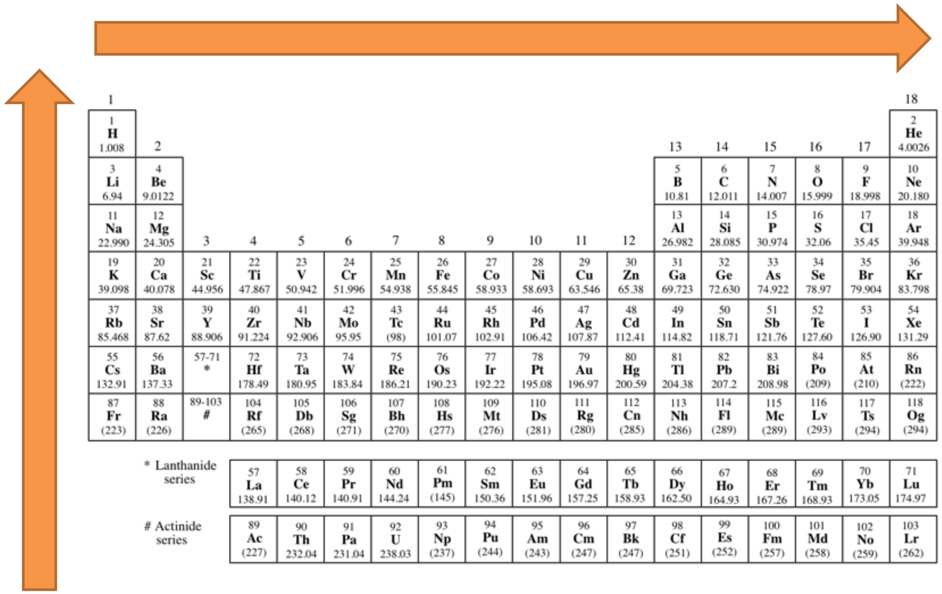
A. O^{+}
B. O
C. O^{-}
D. O^{2-}

First Ionization Energy

- The first ionization energy of an atom is the minimum energy needed by a photon to remove the highest-energy (that is, the outermost) electron from the neutral atom in the gaseous state

$$\mathrm{Li}(1s^2 2s^1) \rightarrow \mathrm{Li}^+(1s^2) + e^-$$

- First ionization energy generally increases as one goes to the right and up the periodic table



Ionization Energy

Using periodic trends, arrange the following elements by increasing first ionization energy: Ar, Na, Cl, Al

Second Ionization Energy

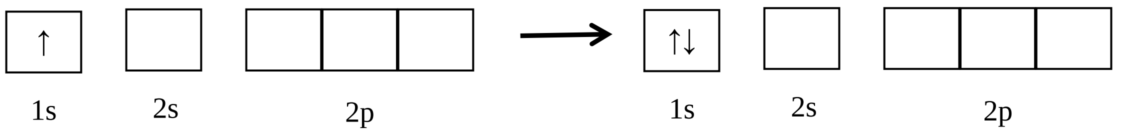
- Accordingly, the second ionization energy of atom is the minimum energy needed to remove the second-highest energy electron from the cation of the atom with a +1 charge

$$Li^{+}(1s^{2}) \rightarrow Li^{2+}(1s^{1}) + e^{-}$$

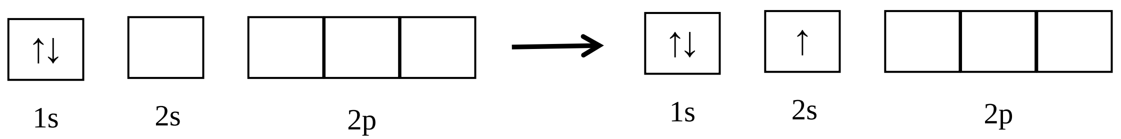
- The second ionization energy will always be larger than the first ionization energy since the second ionization energy involves pulling an electron from a cation

Key Information about Orbital Energies

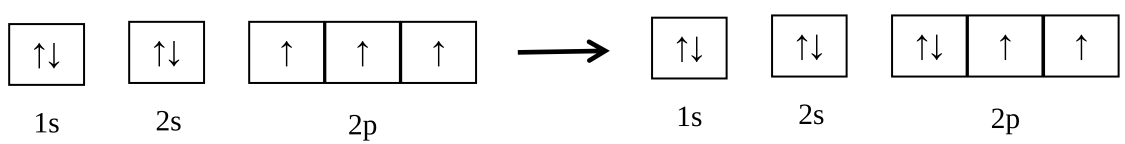
- Filling a subshell lowers the energy of all of the electrons in that subshell



- Adding a new electron to a subshell in a higher energy shell adds a significant amount of energy to the atom



- Adding an electron which results in the first spin pairing results in a lot of energy due to the new electron repulsion in that orbital



Ionization Energy Trend Deviations

First Ionization Energies of Some Elements (kJ/mol)

Period \ Group	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 1310																	He 2370
2	Li 520	Be 900											B 800	C 1090	N 1400	O 1310	F 1680	Ne 2080
3	Na 490	Mg 730											Al 580	Si 780	P 1060	S 1000	Cl 1250	Ar 1520
4	K 420	Ca 590	Sc 630	Ti 660	V 650	Cr 660	Mn 710	Fe 760	Co 760	Ni 730	Cu 740	Zn 910	Ga 580	Ge 780	As 960	Se 950	Br 1140	Kr 1350
5	Rb 400	Sr 550	Y 620	Zr 660	Nb 670	Mo 680	Tc 700	Ru 710	Rh 720	Pd 800	Ag 730	Cd 870	In 560	Sn 700	Sb 830	Te 870	I 1010	Xe 1170
6	Cs 380	Ba 500	La 540	Hf 700	Ta 760	W 770	Re 760	Os 840	Ir 890	Pt 870	Au 890	Hg 1000	Tl 590	Pb 710	Bi 800	Po 810	At . . .	Rn 1030
7	Fr . . .	Ra 510																

Figure 6.35 This version of the periodic table shows the first ionization energy of (IE_1), in kJ/mol, of selected elements.

Successive Ionization Energies and the Quantum Leap

Element	First	Second	Third	Fourth	Fifth	Sixth	Seventh
H	1312						
He	2372	5250					
Li	520	7298	11,815				
Be	899	1757	14,848	21,006			
B	801	2427	3660	25,025	32,826		
C	1086	2353	4620	6222	37,829	47,276	
N	1402	2857	4578	7475	9445	53,265	64,358
O	1314	3388	5306	7469	10,989	13,326	71,333

- Successive ionization energies get increasing large
- The quantum leap is a large increase in successive ionization energies that refers to pulling the first core electron from an atom

Ionization Energy

When trying to remove electrons from Be, which of the following sets of ionization energies makes the most sense going from first to third ionization energy?

A. First IE 900 kJ/mol, second IE 1,750 kJ/mol, third IE 15,000 kJ/mol
B. First IE 1,750 kJ/mol, second IE 900 kJ/mol, third IE 15,000 kJ/mol
C. First IE 15,000 kJ/mol, second IE 1,750 kJ/mol, third IE 900 kJ/mol
D. First IE 900 kJ/mol, second IE 15,000 kJ/mol, third IE 22,000 kJ/mol

Electron Affinity

- Electron affinity is the energy change for the process of adding an electron to a neutral atom in the gaseous state to form a negative atom

$$\mathrm{Cl}\ ([\mathrm{Ne}]3s^2 3p^5) + \mathrm{e}^- \rightarrow + \mathrm{Cl}^-([\mathrm{Ne}]3s^2 3p^6)$$

- If the negative ion is stable (does not spontaneously disintegrate into the natural atom and free electron), the energy change for its formation is a negative number.
- Larger negative numbers such as this indicate that a very stable negative ion is formed
- Smaller negative numbers indicate that a less stable ion is formed.

Electron Affinity

$$B\ (1s^2 2s^2 2p^1) + e^- \rightarrow B^- (1s^2 2s^2 2p^2)\ \text{E. A.} = -23\ \frac{\text{kJ}}{\text{mol}}$$

$$C\ (1s^2 2s^2 2p^2) + e^- \rightarrow C^- (1s^2 2s^2 2p^3)\ \text{E. A.} = -123\ \frac{\text{kJ}}{\text{mol}}$$

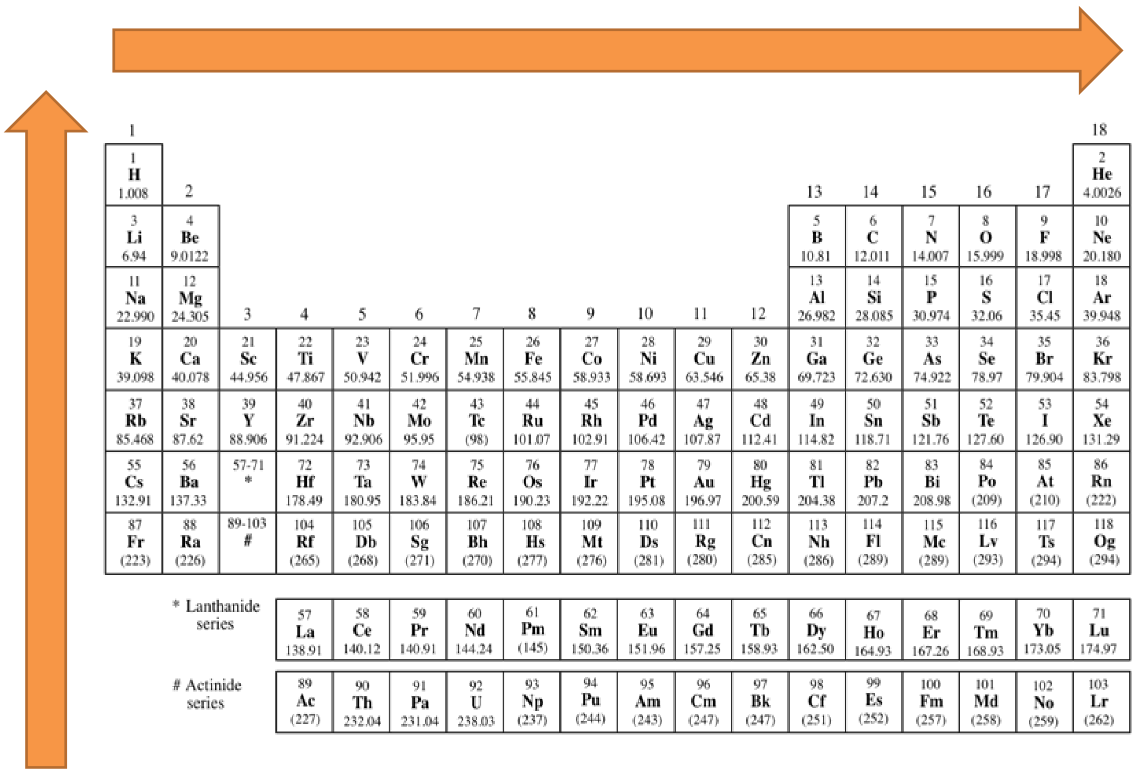
$$N\ (1s^2 2s^2 2p^3) + e^- \rightarrow N^- (1s^2 2s^2 2p^4)\ \text{E. A.} = 0\ \frac{\text{kJ}}{\text{mol}}$$

$$O\ (1s^2 2s^2 2p^4) + e^- \rightarrow O^- (1s^2 2s^2 2p^5)\ \text{E. A.} = -141\ \frac{\text{kJ}}{\text{mol}}$$

$$F\ (1s^2 2s^2 2p^5) + e^- \rightarrow F^- (1s^2 2s^2 2p^6)\ \text{E. A.} = -322\ \frac{\text{kJ}}{\text{mol}}$$

Electron Affinity

- Electron affinity generally increases as one goes to the right and up the periodic table



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.0026
3 **Li** 6.94	4 **Be** 9.0122											5 **B** 10.81	6 **C** 12.011	7 **N** 14.007	8 **O** 15.999	9 **F** 18.998	10 **Ne** 20.180
11 **Na** 22.990	12 **Mg** 24.305											13 **Al** 26.982	14 **Si** 28.085	15 **P** 30.974	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.948
19 **K** 39.098	20 **Ca** 40.078	21 **Sc** 44.956	22 **Ti** 47.867	23 **V** 50.942	24 **Cr** 51.996	25 **Mn** 54.938	26 **Fe** 55.845	27 **Co** 58.933	28 **Ni** 58.693	29 **Cu** 63.546	30 **Zn** 65.38	31 **Ga** 69.723	32 **Ge** 72.630	33 **As** 74.922	34 **Se** 78.97	35 **Br** 79.904	36 **Kr** 83.798
37 **Rb** 85.468	38 **Sr** 87.62	39 **Y** 88.906	40 **Zr** 91.224	41 **Nb** 92.906	42 **Mo** 95.95	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.76	52 **Te** 127.60	53 **I** 126.90	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57-71 *	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.84	75 **Re** 186.21	76 **Os** 190.23	77 **Ir** 192.22	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** (226)	89-103 #	104 **Rf** (265)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (270)	108 **Hs** (277)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **Cn** (285)	113 **Nh** (286)	114 **Fl** (289)	115 **Mc** (289)	116 **Lv** (293)	117 **Ts** (294)	118 **Og** (294)

* Lanthanide series	57 **La** 138.91	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.96	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.50	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.05	71 **Lu** 174.97
# Actinide series	89 **Ac** (227)	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)

Electron Affinities

From what you know in a general way about electron affinities, state which member of each of the following pairs has the more negative value:

A. As, Br
B. F, Li

Electron Affinity Trend Deviations

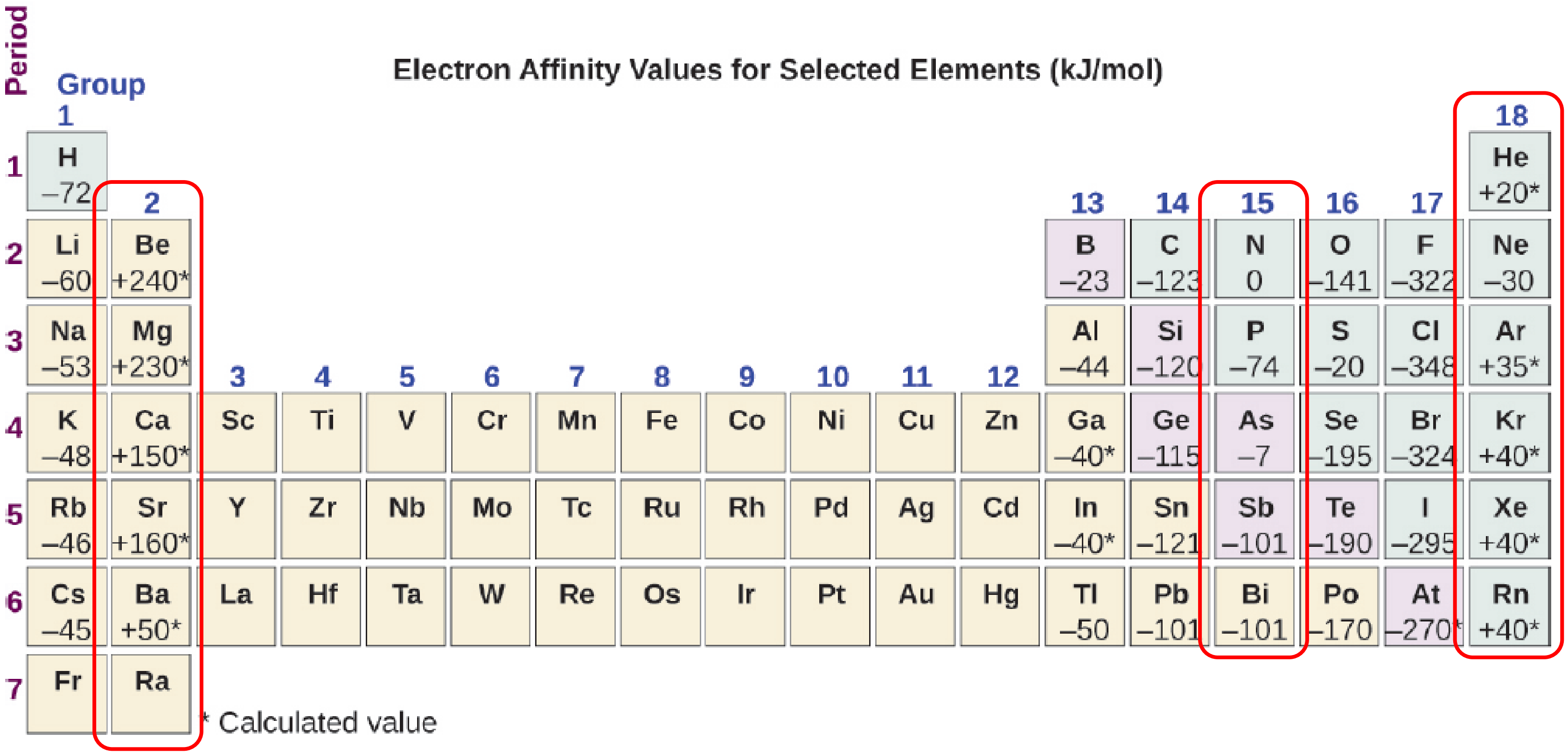


Figure 6.36 This version of the periodic table displays the electron affinity values (in kJ/mol) for selected elements.

Ionization Energy and Electron Affinity

Ionization Energy

- Corresponds to loosing e^-
- Values become greater <u>positive</u> values as one goes up and to the right in the periodic table

Both

- Affected by going from one subshell to another
- Affected by spin pairing e^- in an orbital

Electron Affinity

- Corresponds to gaining e^-
- Values become more <u>negative</u> values as one goes up and to the right in the periodic table

Trend Deviations

Sodium has a greater affinity for electrons than magnesium. Which of the following explains sodium's greater affinity for electrons?

A. The electron affinity of magnesium corresponds to adding the first spin pairing to the 2p subshell
B. The electron affinity of magnesium corresponds to adding the first electron to the 2p subshell
C. Electron affinity generally increases as one goes towards the left and down the periodic table
D. Electron affinity generally increases as one goes towards the right and up the periodic table

Metals vs. Nonmetals

Metals

- Have small positive values for ionization energy
- Thus, it doesn't take much energy for them to lose an e^- and form a cation
- Have small negative values for electron affinity
- Means the system doesn't loose a lot of energy when gaining an e^-
- Thus, the system doesn't get a lot of stabilization by gaining an e^- and forming an anion

Nonmetals

- Have large positive values for ionization energy
- Thus, it takes a lot of energy for them to lose an e^- and form a cation
- Have large negative values for electron affinity
- Means the system looses a lot of energy when gaining an e^-
- Thus, the system gets a lot of stabilization by gaining an e^- and forming an anion

Ion Formation

Cations

- Group 1 atoms normally form ions with a +1 charge since they have low first ionization energies
- Group 2 atoms normally form ions with a +2 charge since they have low first and second ionization energies
- Group 13 atoms normally form ions with a +3 charge since they have low first, second, and third ionization energies.

Ion Formation

Cations

- Group 1 atoms normally form ions with a +1 charge since they have low first ionization energies
- Group 2 atoms normally form ions with a +2 charge since they have low first and second ionization energies
- Group 13 atoms normally form ions with a +3 charge since they have low first, second, and third ionization energies.

Anions

- Group 17 atoms normally form ions with a -1 since gaining one e^- fills the subshell
- Group 16 atoms normally form ions with a -2 charge since gaining two e^- fills the subshell
- Group 15 atoms normally form ions with a -3 charge since gaining three e^- fills the subshell

Ion Formation

Cations

- Group 1 atoms normally form ions with a +1 charge since they have low first ionization energies
- Group 2 atoms normally form ions with a +2 charge since they have low first and second ionization energies
- Group 13 atoms normally form ions with a +3 charge since they have low first, second, and third ionization energies.

Anions

- Group 17 atoms normally form ions with a -1 since gaining one e^- fills the subshell
- Group 16 atoms normally form ions with a -2 charge since gaining two e^- fills the subshell
- Group 15 atoms normally form ions with a -3 charge since gaining three e^- fills the subshell

Group 14 atoms normally don't form ions since it requires a lot of energy to gain or loose 4 e^-

Ionic Bonds

- Atoms held together by electrostatic attraction between cations and anions
- Primarily between metals and nonmetals
- Involves electron transfer

Ionic Lattice Formation

$$Na(1s^2 2s^2 2p^6 3s^1) \rightarrow Na^+(1s^2 2s^2 2p^6) + e^-$$

$$Cl(1s^2 2s^2 2p^6 3s^2 3p^5) + e^- \rightarrow Cl^-(1s^2 2s^2 2p^6 3s^2 3p^6)$$