

Development of the Periodic Table (Section 8.1 and 8.2)

- In 2002, there were 115 elements known.
- The majority of the elements were discovered between 1735 and 1843.
- How do we organize 115 different elements in a meaningful way that will allow us to make predictions about undiscovered elements?
- Arrange elements to reflect the trends in chemical and physical properties.
- First attempt (Mendeleev and Meyer) arranged the elements in order of increasing atomic weight.
- Certain elements were missing from this scheme.

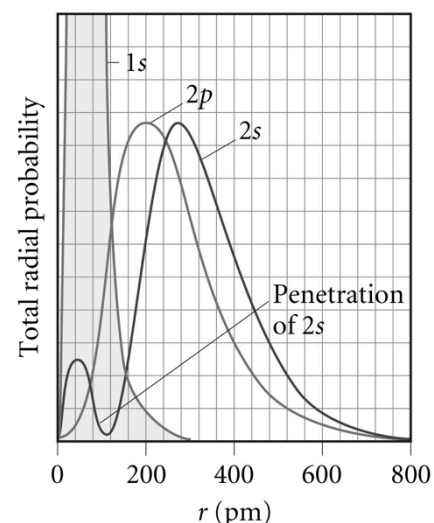
Example: In 1871, Mendeleev noted that As properly belonged underneath P and not Si, which left a missing element underneath Si. He predicted a number of properties for this element. In 1886 Ge was discovered. The properties of Ge match Mendeleev's predictions well.

- Modern periodic table: arrange elements in order of increasing atomic number.

TABLE 7.8 Some Properties of the Noble Gases

Element	Electron Configuration	Boiling Point (K)	Density (g/L)	Atomic Radius* (Å)	I_1 (kJ/mol)
Helium	$1s^2$	4.2	0.18	0.32	2372
Neon	$[\text{He}]2s^22p^6$	27.1	0.90	0.69	2081
Argon	$[\text{Ne}]3s^23p^6$	87.3	1.78	0.97	1521
Krypton	$[\text{Ar}]3d^{10}4s^24p^6$	120	3.75	1.10	1351
Xenon	$[\text{Kr}]4d^{10}5s^25p^6$	165	5.90	1.30	1170
Radon	$[\text{Xe}]4f^{14}5d^{10}6s^26p^6$	211	9.73	—	1037

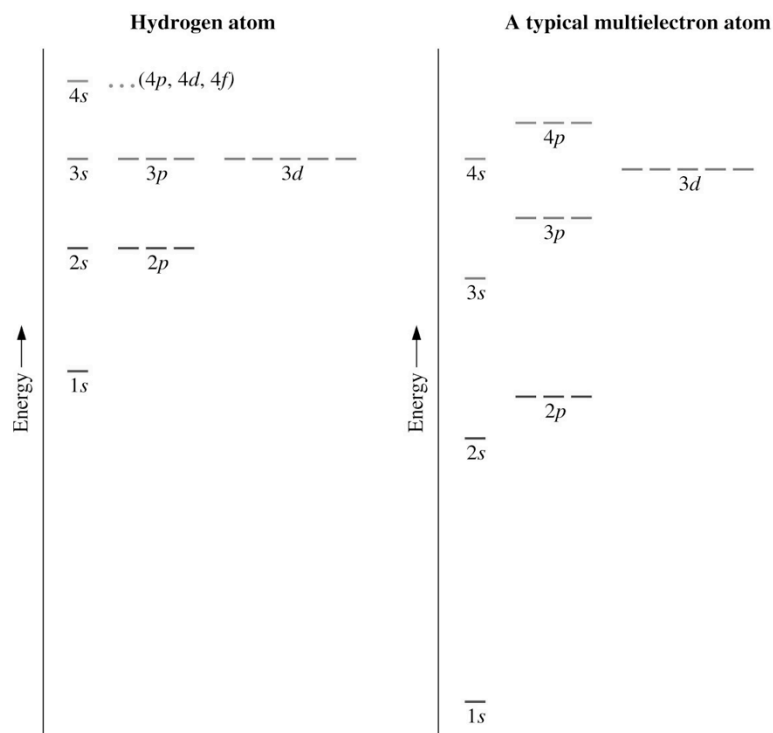
* Only the heaviest of the noble-gas elements form chemical compounds. Thus, the atomic radii for the lighter noble-gas elements are predicted, estimated values.



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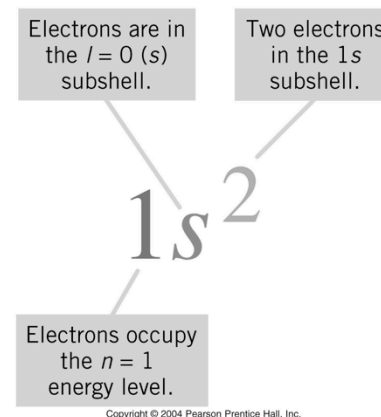
Electron Configurations (Section 8.3)

- Electrons ordinarily occupy orbitals of the lowest energy available.
- No two electrons in the same atom may have all four quantum numbers alike.
- Pauli exclusion principle:** one atomic orbital can accommodate no more than **two** electrons, and these electrons must have **opposing** spins.
- Of a group of orbitals of **identical** energy, electrons enter **empty** orbitals whenever possible (**Hund's rule**).
- Electrons in half-filled orbitals have **parallel** spins (same direction).
- Core electrons:** electrons in [Noble Gas].
- Valence electrons:** electrons outside of [Noble Gas].

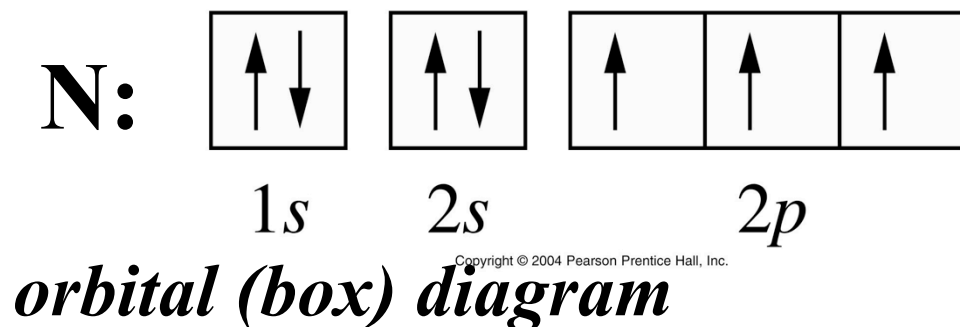


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The *spdf notation* uses numbers to designate a principal shell and letters (*s*, *p*, *d*, *f*) to identify a subshell; a superscript indicates the number of electrons in a designated subshell.



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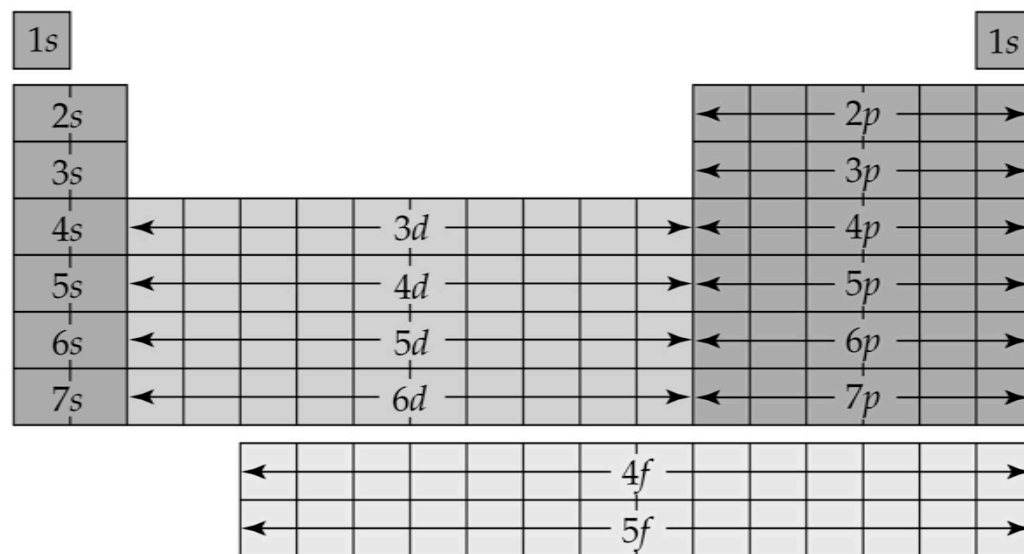
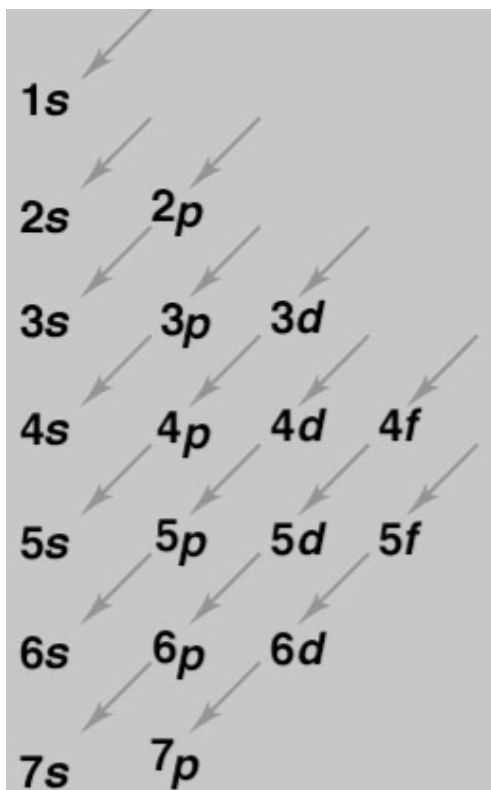


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Electron Configurations and the Periodic Table (Section 8.4)

- The periodic table can be used as a guide for electron configurations.
- The period number is the value of n .
- Groups 1A and 2A have the s -orbital filled.
- Groups 3A - 8A have the p -orbital filled.
- Groups 3B - 2B have the d -orbital filled.
- The lanthanides and actinides have the f -orbital filled.

• **Noble-gas-core abbreviation:** we can replace the portion that corresponds to the electron configuration of a noble gas with a bracketed chemical symbol. It's easier to write ...



Representative s-block elements

Transition metals

Representative p-block elements

f-Block metals

Orbital Blocks of the Periodic Table

Groups																		18	
		1A	2A											3A	4A	5A	6A	7A	8A
Periods	1	1 H $1s^1$																	2 He $1s^2$
	2	3 Li $2s^1$	4 Be $2s^2$											5 B $2s^2 2p^1$	6 C $2s^2 2p^2$	7 N $2s^2 2p^3$	8 O $2s^2 2p^4$	9 F $2s^2 2p^5$	10 Ne $2s^2 2p^6$
	3	11 Na $3s^1$	12 Mg $3s^2$	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 Al $3s^2 3p^1$	14 Si $3s^2 3p^2$	15 P $3s^2 3p^3$	16 S $3s^2 3p^4$	17 Cl $3s^2 3p^5$	18 Ar $3s^2 3p^6$
	4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $4s^2 3d^1$	22 Ti $4s^2 3d^2$	23 V $4s^2 3d^3$	24 Cr $4s^1 3d^5$	25 Mn $4s^2 3d^5$	26 Fe $4s^2 3d^6$	27 Co $4s^2 3d^7$	28 Ni $4s^2 3d^8$	29 Cu $4s^1 3d^{10}$	30 Zn $4s^2 3d^{10}$	31 Ga $4s^2 4p^1$	32 Ge $4s^2 4p^2$	33 As $4s^2 4p^3$	34 Se $4s^2 4p^4$	35 Br $4s^2 4p^5$	36 Kr $4s^2 4p^6$
	5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $5s^2 4d^1$	40 Zr $5s^2 4d^2$	41 Nb $5s^1 4d^4$	42 Mo $5s^1 4d^5$	43 Tc $5s^2 4d^5$	44 Ru $5s^1 4d^7$	45 Rh $5s^1 4d^8$	46 Pd $4d^{10}$	47 Ag $5s^1 4d^{10}$	48 Cd $5s^2 4d^{10}$	49 In $5s^2 5p^1$	50 Sn $5s^2 5p^2$	51 Sb $5s^2 5p^3$	52 Te $5s^2 5p^4$	53 I $5s^2 5p^5$	54 Xe $5s^2 5p^6$
	6	55 Cs $6s^1$	56 Ba $6s^2$	57 La $6s^2 5d^1$	72 Hf $6s^2 5d^2$	73 Ta $6s^2 5d^3$	74 W $6s^2 5d^4$	75 Re $6s^2 5d^5$	76 Os $6s^2 5d^6$	77 Ir $6s^2 5d^7$	78 Pt $6s^1 5d^9$	79 Au $6s^1 5d^{10}$	80 Hg $6s^2 5d^{10}$	81 Tl $6s^2 6p^1$	82 Pb $6s^2 6p^2$	83 Bi $6s^2 6p^3$	84 Po $6s^2 6p^4$	85 At $6s^2 6p^5$	86 Rn $6s^2 6p^6$
	7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac $7s^2 6d^1$	104 Rf $7s^2 6d^2$	105 Db $7s^2 6d^3$	106 Sg $7s^2 6d^4$	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112		114		116		

Lanthanides

Actinides

58 Ce $6s^2 4f^2$	59 Pr $6s^2 4f^3$	60 Nd $6s^2 4f^4$	61 Pm $6s^2 4f^5$	62 Sm $6s^2 4f^6$	63 Eu $6s^2 4f^7$	64 Gd $6s^2 4f^7 5d^1$	65 Tb $6s^2 4f^9$	66 Dy $6s^2 4f^{10}$	67 Ho $6s^2 4f^{11}$	68 Er $6s^2 4f^{12}$	69 Tm $6s^2 4f^{13}$	70 Yb $6s^2 4f^{14}$	71 Lu $6s^2 4f^{14} 6d^1$
90 Th $7s^2 6d^2$	91 Pa $7s^2 5f^2 6d^1$	92 U $7s^2 5f^3 6d^1$	93 Np $7s^2 5f^4 6d^1$	94 Pu $7s^2 5f^6$	95 Am $7s^2 5f^7$	96 Cm $7s^2 5f^7 6d^1$	97 Bk $7s^2 5f^9$	98 Cf $7s^2 5f^{10}$	99 Es $7s^2 5f^{11}$	100 Fm $7s^2 5f^{12}$	101 Md $7s^2 5f^{13}$	102 No $7s^2 5f^{14}$	103 Lr $7s^2 5f^{14} 6d^1$

- The **valence shell** is the outermost occupied principal shell. The valence shell contains the **valence electrons**.
- For main group elements, the number of valence shell electrons is the same as the periodic table group number (2A elements: two valence electrons, etc.)
- The period number is the same as the principal quantum number n of the electrons in the valence shell.
- Electrons in inner shells are called **core electrons**.

Half-filled d subshell plus half-filled s subshell has slightly lower in energy than $s^2 d^4$.

Five valence electrons, for which $n = 4$

Example: As $[\text{Ar}] 4s^2 3d^{10} 4p^3$

28 core electrons

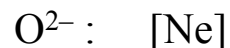
		$3d$					$4s$	
Sc	[Ar]	↑					↑↓	$[\text{Ar}] 3d^1 4s^2$
Ti	[Ar]	↑	↑				↑↓	$[\text{Ar}] 3d^2 4s^2$
V	[Ar]	↑	↑	↑			↑↓	$[\text{Ar}] 3d^3 4s^2$
Cr	[Ar]	↑	↑	↑	↑	↑	↑	$[\text{Ar}] 3d^5 4s^1$
Mn	[Ar]	↑	↑	↑	↑	↑	↑↓	$[\text{Ar}] 3d^5 4s^2$
Fe	[Ar]	↑↓	↑	↑	↑	↑	↑↓	$[\text{Ar}] 3d^6 4s^2$
Co	[Ar]	↑↓	↑↓	↑	↑	↑	↑↓	$[\text{Ar}] 3d^7 4s^2$
Ni	[Ar]	↑↓	↑↓	↑↓	↑	↑	↑↓	$[\text{Ar}] 3d^8 4s^2$
Cu	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$[\text{Ar}] 3d^{10} 4s^1$
Zn	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	$[\text{Ar}] 3d^{10} 4s^2$

Filled d subshell plus half-filled s subshell has slightly lower in energy than $s^2 d^9$.

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Electron Configurations of Ions (Section 8.5 and 8.7)

- To obtain the electron configuration of an **anion** by the aufbau process, we simply **add** the additional electrons to the valence shell of the neutral nonmetal atom.
- The number added usually completes the shell.
- A nonmetal monatomic ion usually attains the electron configuration of a noble gas atom.



- A metal atom loses electrons to form a **cation**.
- Electrons are **removed** from the configuration of the atom.
- The first electrons lost are those of the **highest principal** quantum number.
- If there are two subshells with the same highest principal quantum number, electrons are lost from the subshell with the higher l .

Atom	Ion	(or)
F $1s^2 2s^2 2p^5$	F ⁻ $1s^2 2s^2 2p^6$	[Ne]
S [Ne] $3s^2 3p^4$	S ²⁻ [Ne] $3s^2 3p^6$	[Ar]
Sr [Kr] $5s^2$	Sr ²⁺ [Kr] $5s^2$	[Kr]
Ti [Ar] $4s^2 3d^2$	Ti ⁴⁺ [Ar] $4s^2 3d^2$	[Ar]
Fe [Ar] $4s^2 3d^6$	Fe ²⁺ [Ar] $4s^2 3d^6$	[Ar] $3d^6$

Table 8.3 Electron Configurations of Some Metal Ions

Noble Gas			Pseudo-Noble Gas ^a		18 + 2 ^b	Various
Li ⁺	Be ²⁺	Al ³⁺	Cu ⁺	Zn ²⁺	In ⁺	Cr ²⁺ : [Ar]3d ⁴
Na ⁺	Mg ²⁺		Ag ⁺	Cd ²⁺	Tl ⁺	Cr ³⁺ : [Ar]3d ³
K ⁺	Ca ²⁺		Au ⁺	Hg ²⁺	Sn ²⁺	Mn ²⁺ : [Ar]3d ⁵
Rb ⁺	Sr ²⁺				Pb ²⁺	Mn ³⁺ : [Ar]3d ⁴
Cs ⁺	Ba ²⁺				Sb ³⁺	Fe ²⁺ : [Ar]3d ⁶
					Bi ³⁺	Fe ³⁺ : [Ar]3d ⁵
						Co ²⁺ : [Ar]3d ⁷
						Co ³⁺ : [Ar]3d ⁶
						Ni ²⁺ : [Ar]3d ⁸

^a In the pseudo-noble gas configuration, all valence electrons are lost and the remaining $(n - 1)$ shell has 18 electrons in the configuration $(n - 1)s^2(n - 1)p^6(n - 1)d^{10}$.

^b In the 18 + 2 configuration, $(n - 1)s^2(n - 1)p^6(n - 1)d^{10}ns^2$, two valence electrons remain.

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Magnetic Properties (Section 8.7)

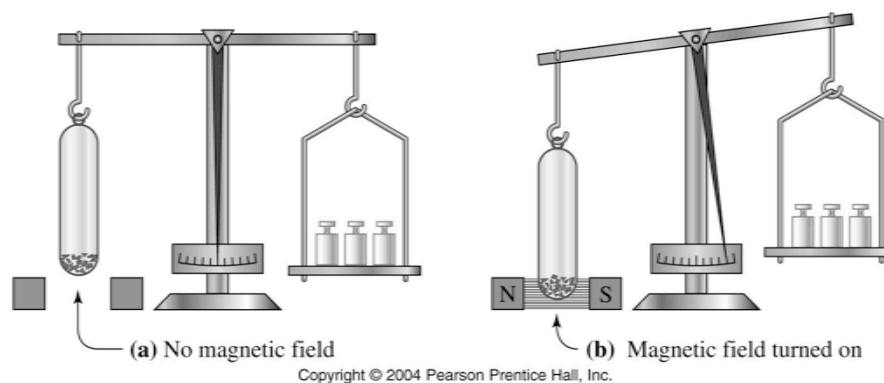
• **Diamagnetism** is the weak repulsion associated with *paired* electrons.

• **Paramagnetism** is the attraction associated with *unpaired* electrons.

– This produces a much stronger effect than the weak diamagnetism of paired electrons.

• **Ferromagnetism** is the exceptionally strong attractions of a magnetic field for iron and a few other substances.

A sample of chlorine gas is found to be diamagnetic.
Can this gaseous sample be composed of individual Cl atoms?

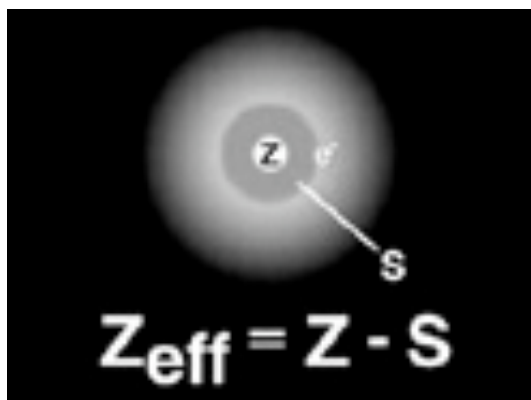


Elements That Form Ions with Predictable Charges

	1A	2A														3A	4A	5A	6A	7A	8A
1	Li ⁺																	N ³⁻	O ²⁻	F ⁻	
2	Na ⁺	Mg ²⁺														Al ³⁺			S ²⁻	Cl ⁻	
3	K ⁺	Ca ²⁺																	Se ²⁻	Br ⁻	
4	Rb ⁺	Sr ²⁺																	Te ²⁻	I ⁻	
5	Cs ⁺	Ba ²⁺																			

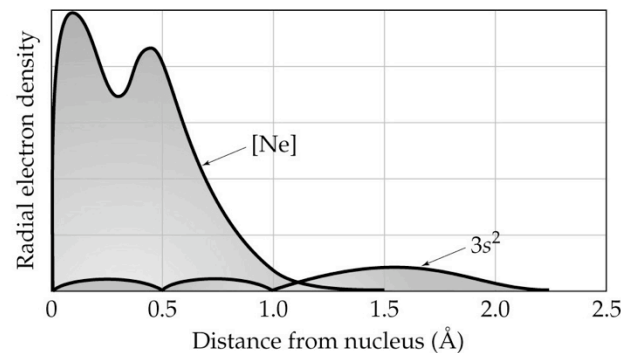
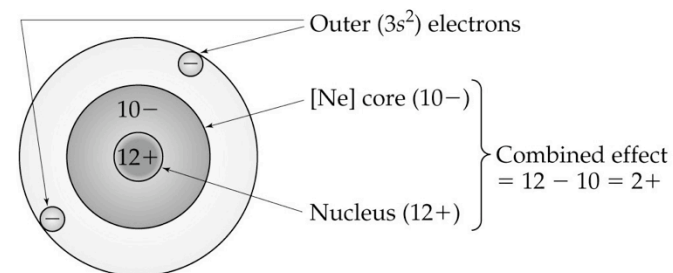
Effective Nuclear Charge (8.3 and 8.6)

- Effective nuclear charge is the charge experienced by an electron on a many-electron atom.
- The effective nuclear charge is not the same as the charge on the nucleus because of the effect of the inner electrons.
- Electrons are attracted to the nucleus, but repelled by the electrons that screen it from the nuclear charge.
- The nuclear charge experienced by an electron depends on its distance from the nucleus and the number of core electrons.
- As the average number of screening electrons (S) increases, the effective nuclear charge (Z_{eff}) decreases.
- As the distance from the nucleus increases, S increases and Z_{eff} decreases.
- The ns orbitals all have the same shape, but have different sizes and different numbers of nodes.
- Consider: He: $1s^2$, Ne: $1s^2 2s^2 2p^6$, and Ar: $1s^2 2s^2 2p^6 3s^2 3p^6$.
- The radial electron density is the probability of finding an electron at a given distance.



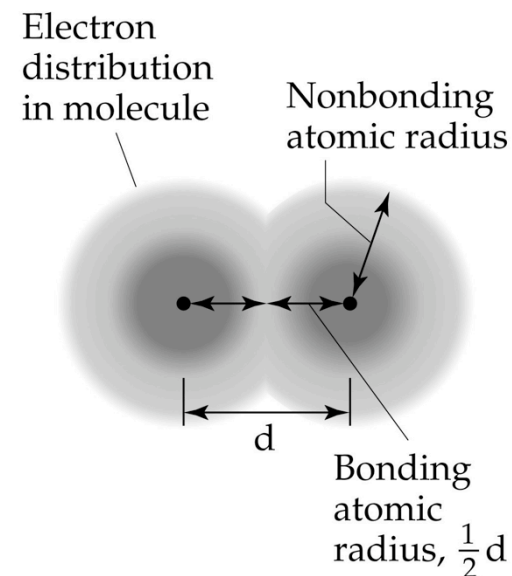
Z = # of protons

S = # of core electrons



Sizes of Atoms and Ions (Sections 8.6 and 8.7)

- Consider a simple diatomic molecule.
- The distance between the two nuclei is called the bond distance.
- If the two atoms which make up the molecule are the same, then half the bond distance is called the covalent radius of the atom.
- As the principal quantum number increases, the size of the orbital increases.
- Consider the s orbitals.
- All s orbitals are spherical and increase in size as n increases.
- The spherical symmetry of the orbitals can be seen in the contour plots.
- Contour plots are connecting points of equal electron density.

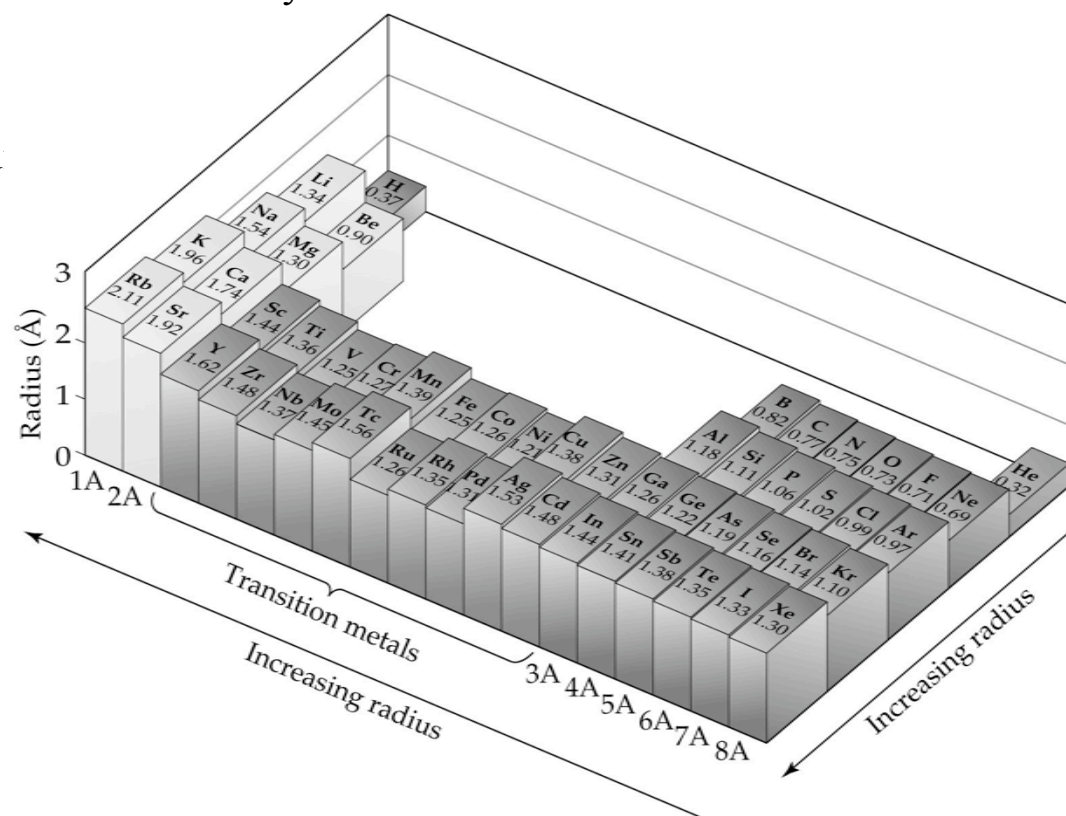


Periodic Trends in Atomic Radii

- As a consequence of the ordering in the periodic table, properties of elements vary periodically.
- Atomic size varies consistently through the periodic table.
- As we move down a group, the atoms become larger.
- As we move across a period, atoms become smaller.

There are two factors at work:

- principal quantum number, n , and
- the effective nuclear charge, Z_{eff}



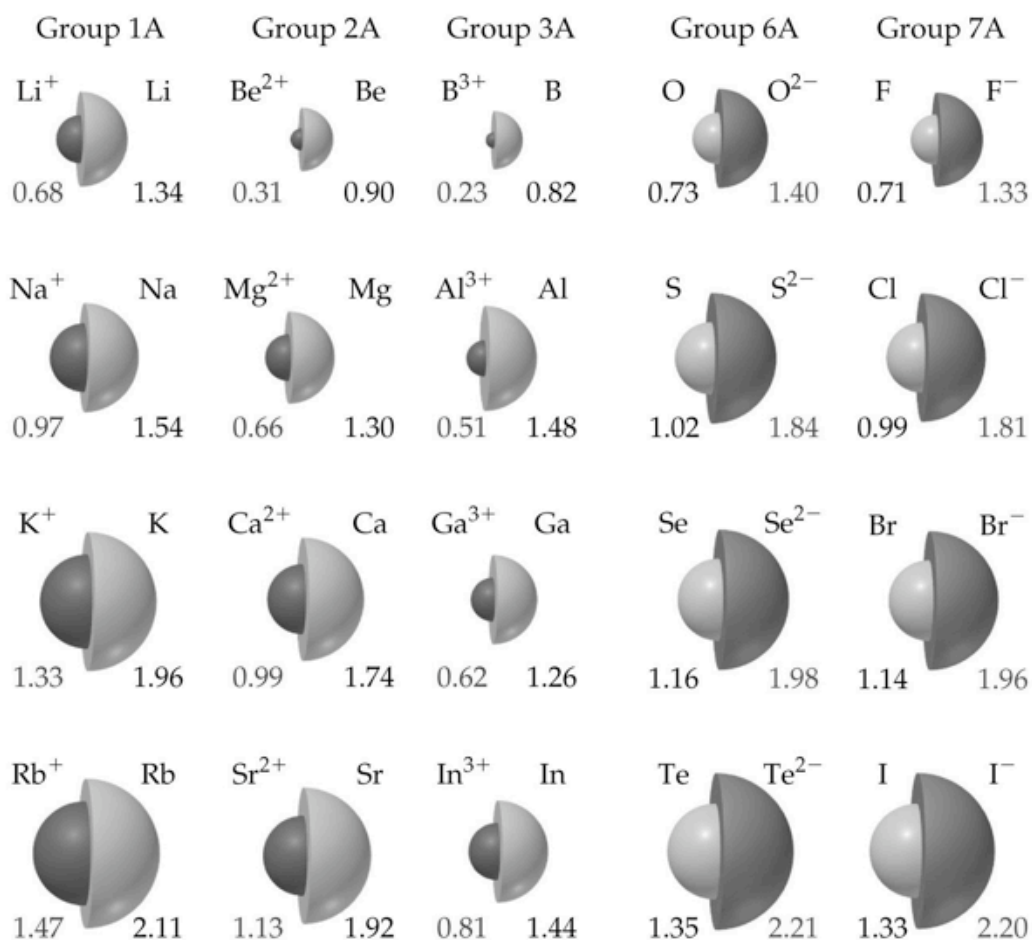
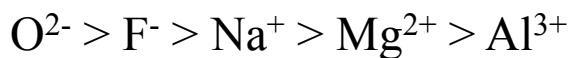
Trends in the Sizes of Ions (Section 8.7)

- Ion size is the distance between ions in an ionic compound.
- Ion size also depends on nuclear charge, number of electrons, and orbitals that contain the valence electrons.
- Cations vacate the most spatially extended orbital and are **smaller than the parent ion**.
- Anions add electrons to the most spatially extended orbital and are **larger than the parent ion**.

• For ions of the same charge, ion size increases down a group.

• All the members of an **isoelectronic series** have the same number of electrons.

• As nuclear charge increases in an isoelectronic series the ions become smaller:

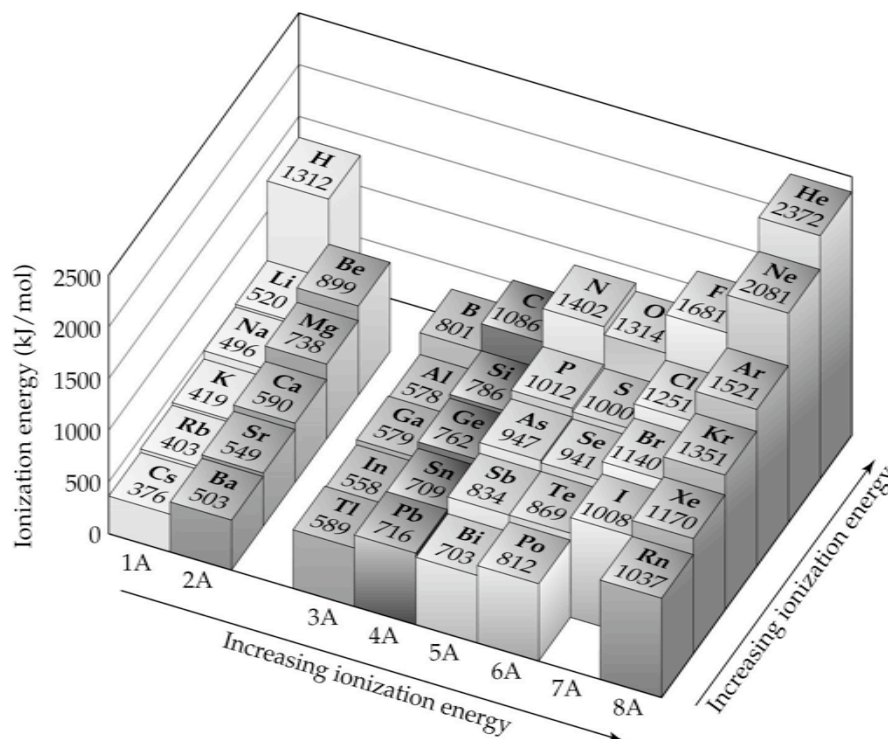


Ionization Energy (Section 8.7)

- The first ionization energy, I_1 , is the amount of energy required to remove an electron from a gaseous atom: $\text{Na}(g) \rightarrow \text{Na}^+(g) + e^-$.
- The second ionization energy, I_2 , is the energy required to remove an electron from a gaseous ion:
 $\text{Na}^+(g) \rightarrow \text{Na}^{2+}(g) + e^-$.
- The larger ionization energy, the more difficult it is to remove the electron.

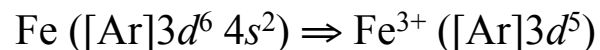
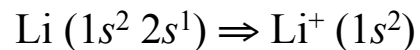
Periodic Trends in Ionization Energies

- Ionization energy decreases down a group.
- This means that the outermost electron is more readily removed as we go down a group.
- As the atom gets bigger, it becomes easier to remove an electron from the most spatially extended orbital.
 - Ionization energy generally increases across a period.
 - As we move across a period, Z_{eff} increases. Therefore, it becomes more difficult to remove an electron.
- Two exceptions: removing the first p electron and removing the fourth p electron.
- The s electrons are more effective at shielding than p electrons. Therefore, forming the s^2p^0 becomes more favorable.
- When a second electron is placed in a p orbital, the electron-electron repulsion increases. When this electron is removed, the resulting s^2p^3 is more stable than the starting s^2p^4 configuration. Therefore, there is a decrease in ionization energy.

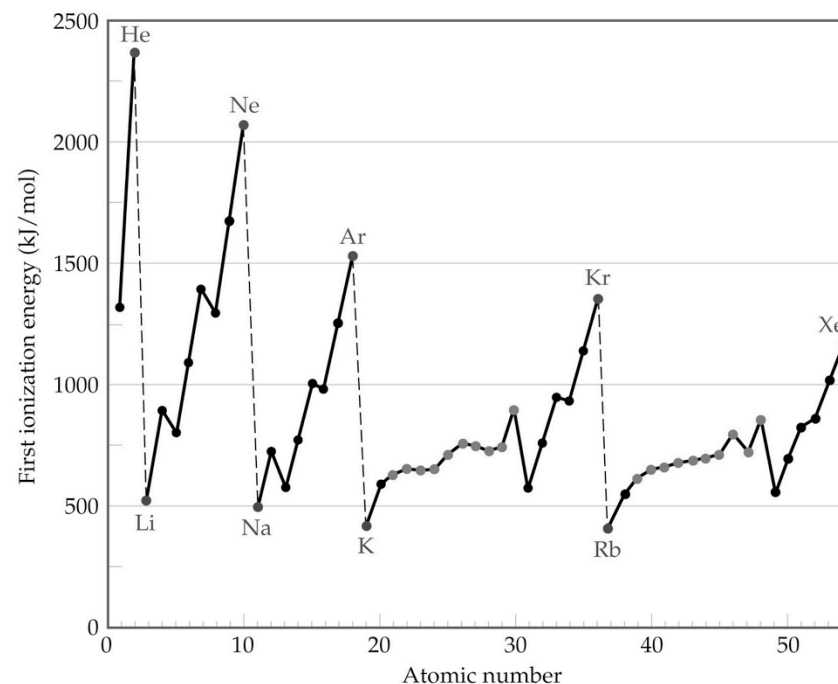
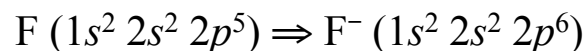


Electron Configuration of Ions

- Cations:** electrons removed from orbital with highest principle quantum number, n , first:



- Anions:** electrons added to the orbital with highest n :



Variations in Successive Ionization Energies

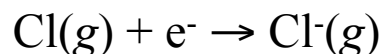
- There is a sharp increase in ionization energy when a core electron is removed.

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

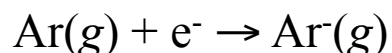
Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	496	4560	(inner-shell electrons)				
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,100		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,000
Ar	1521	2670	3930	5770	7240	8780	12,000

Electron Affinities (Section 8.8)

- Electron affinity is the opposite of ionization energy.
- Electron affinity is the energy change when a gaseous atom gains an electron to form a gaseous ion:



- Electron affinity can either be exothermic (as the above example) or endothermic:



- Look at electron configurations to determine whether electron affinity is positive or negative.
 - The extra electron in Ar needs to be placed in the 4s orbital which is significantly higher in energy than the 3p orbital.

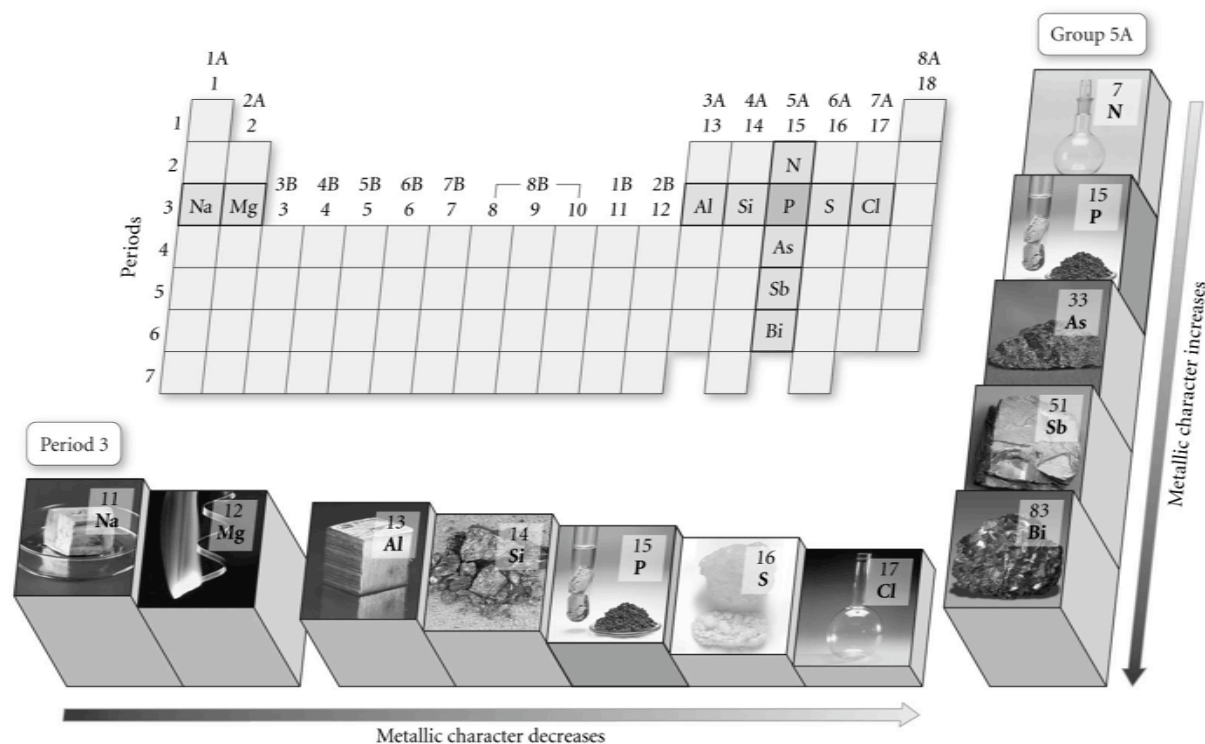
H -73							He >0
Li -60	Be >0	B -27	C -122	N >0	O -141	F -328	Ne >0
Na -53	Mg >0	Al -43	Si -134	P -72	S -200	Cl -349	Ar >0
K -48	Ca -2	Ga -30	Ge -119	As -78	Se -195	Br -325	Kr >0
Rb -47	Sr -5	In -30	Sn -107	Sb -103	Te -190	I -295	Xe >0
1A	2A	3A	4A	5A	6A	7A	8A

Metals, Nonmetals, and Metalloids (Section 8.8)

Metals

- Metallic character refers to the properties of metals (shiny or lustrous, malleable and ductile, oxides form basic ionic solids, and tend to form cations in aqueous solution).
- Metallic character increases down a group.
- Metallic character decreases across a period.
- Metals have low ionization energies.
- Most neutral metals are oxidized rather than reduced.

Trends in Metallic Character II



A Summary of Trends

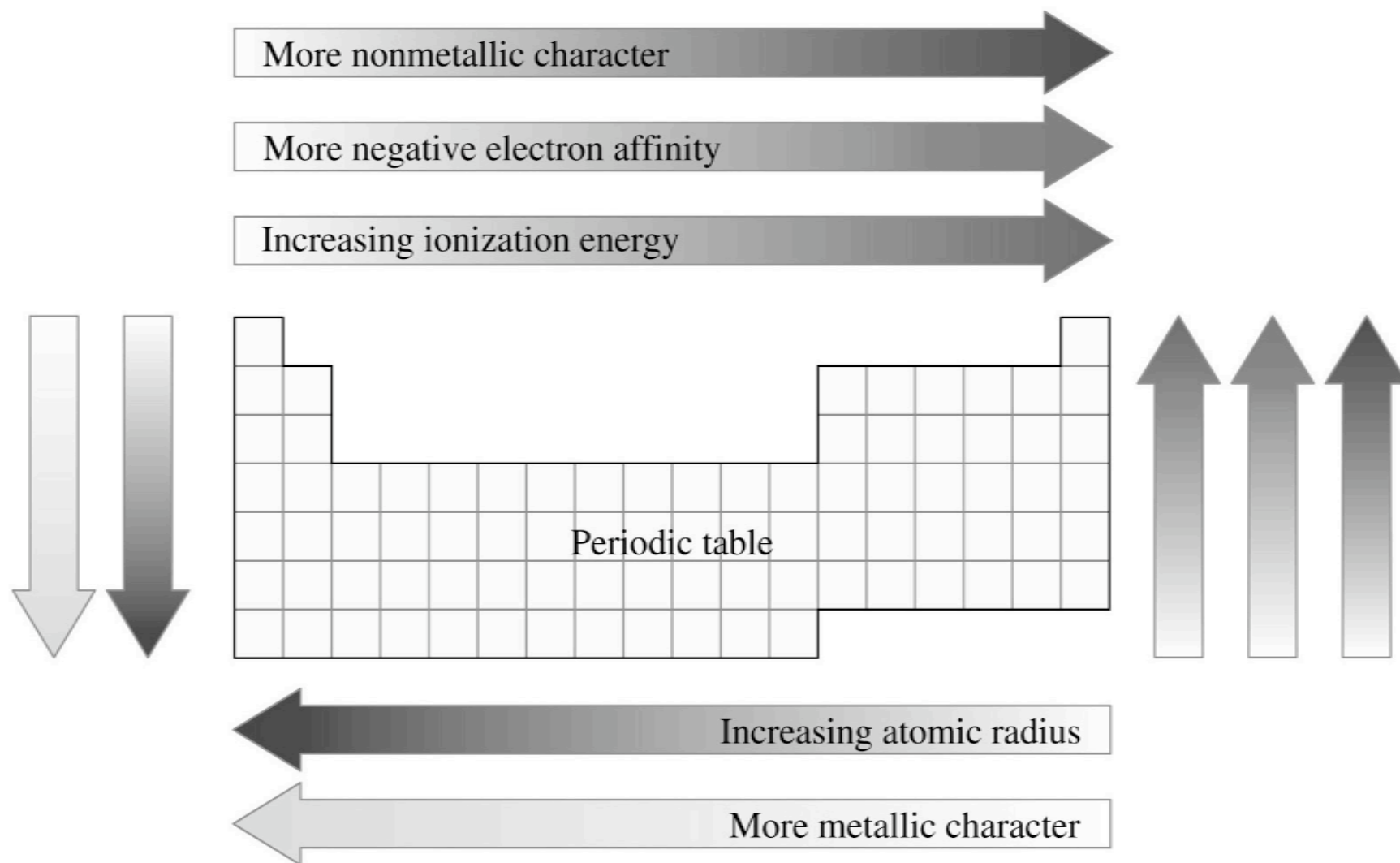


TABLE 7.4 Some Properties of the Alkali Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm ³)	Atomic Radius (Å)	<i>I</i> ₁ (kJ/mol)
Lithium	[He]2s ¹	181	0.53	1.34	520
Sodium	[Ne]3s ¹	98	0.97	1.54	496
Potassium	[Ar]4s ¹	63	0.86	1.96	419
Rubidium	[Kr]5s ¹	39	1.53	2.11	403
Cesium	[Xe]6s ¹	28	1.88	2.60	376

Group 2A: The Alkaline Earth Metals

TABLE 7.5 Some Properties of the Alkaline Earth Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm ³)	Atomic Radius (Å)	<i>I</i> ₁ (kJ/mol)
Beryllium	[He]2s ²	1287	1.85	0.90	899
Magnesium	[Ne]3s ²	650	1.74	1.30	738
Calcium	[Ar]4s ²	842	1.54	1.74	590
Strontium	[Kr]5s ²	777	2.63	1.92	549
Barium	[Xe]6s ²	727	3.51	2.15	503

Group 7A: The Halogens

TABLE 7.7 Some Properties of the Halogens

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	<i>I</i> ₁ (kJ/mol)
Fluorine	[He]2s ² 2p ⁵	−220	1.69 g/L	0.71	1681
Chlorine	[Ne]3s ² 3p ⁵	−102	3.21 g/L	0.99	1251
Bromine	[Ar]3d ¹⁰ 4s ² 4p ⁵	−7.3	3.12 g/cm ³	1.14	1140
Iodine	[Kr]4d ¹⁰ 5s ² 5p ⁵	114	4.93 g/cm ³	1.33	1008