General Chemistry

Principles and Modern Applications

Petrucci • Harwood • Herring

8^{th} Edition

Chapter 9: The Periodic Table and Some Atomic Properties

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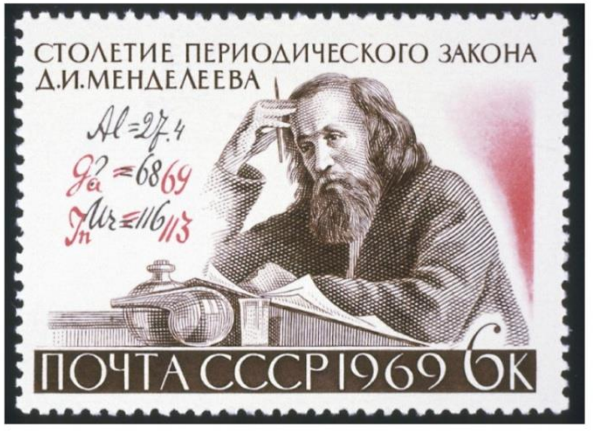
Contents

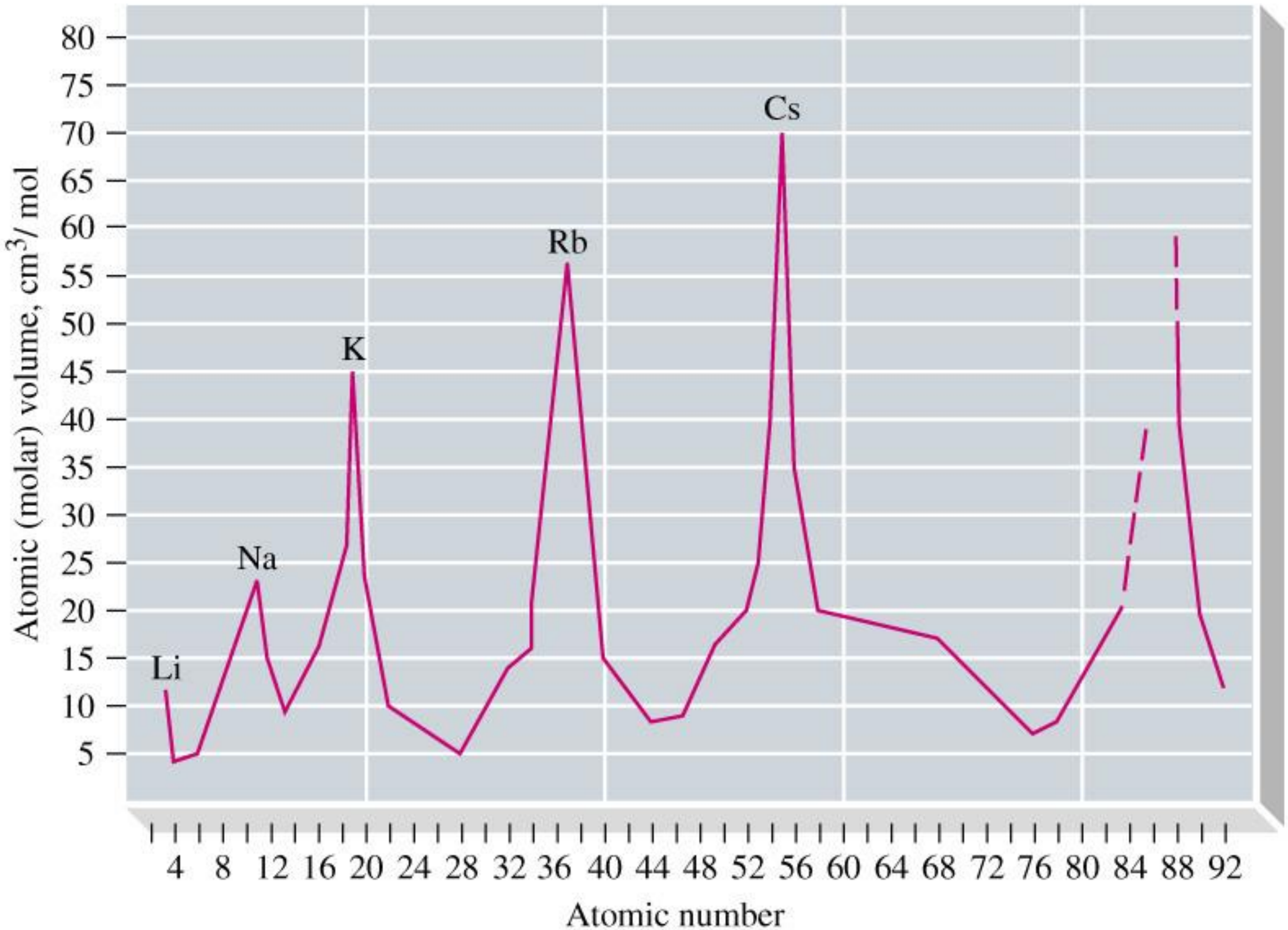
9-1 Classifying the Elements: The Periodic Law and the Periodic Table

- 1869, Dimitri Mendeleev
  Lothar Meyer



When the elements are arranged in order of increasing atomic mass, certain sets of properties recur periodically.

Periodic Law

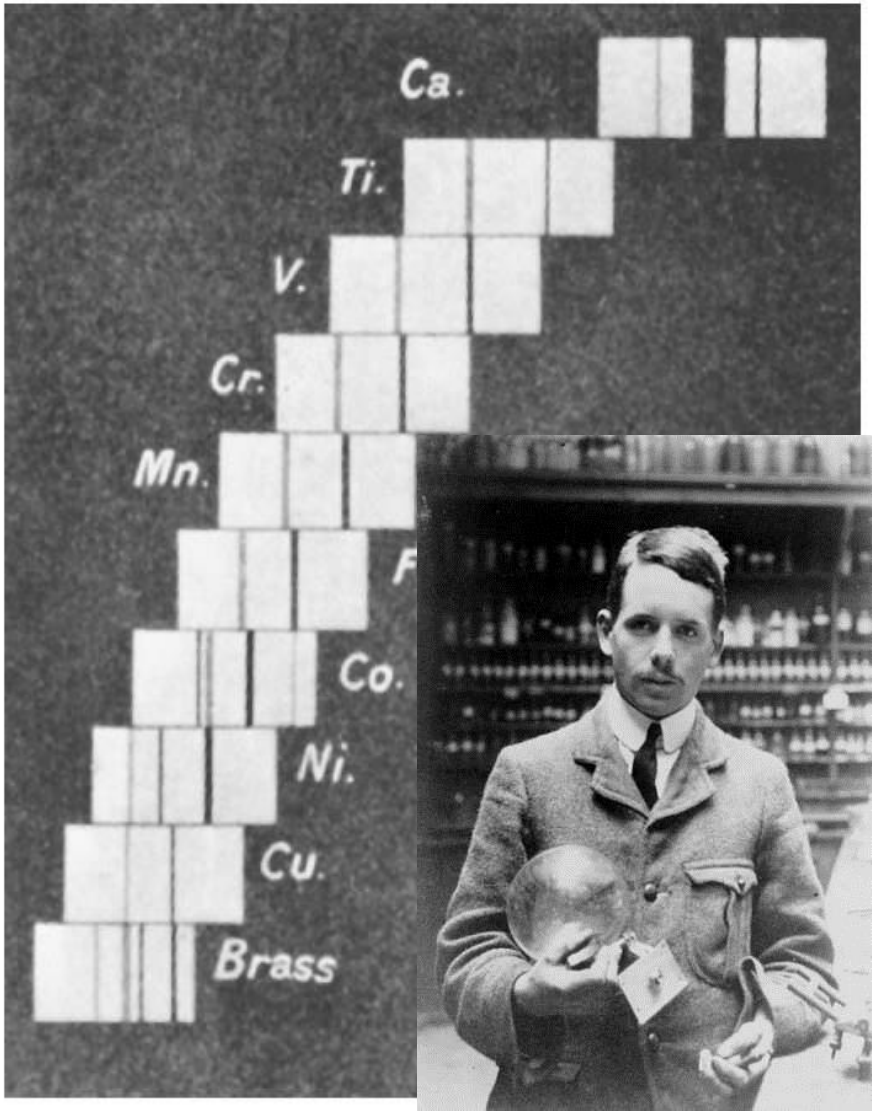


Mendeleev's Periodic Table

1871

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	**— = 44**	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	**— = 68**	**— = 72**	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	**— = 100**	Ru = 104, Rh = 104, Pd = 106, Ag = 108
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	–	–	–	– – – –
9	(–)	–	–	–	–	–	–	
10	–	–	?Er = 178	?La = 180	Ta = 182	W = 184	–	Os = 195, Ir = 197, Pt = 198, Au = 199
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208			
12	–	–	–	Th = 231	–	U = 240		

The Periodic Table



- Moseley found a relationship between the frequency of X-rays and the number of charges in the nuclei of elements and Mendeleev's periodic table.
- According to Moseley; Similar properties recur periodically when elements are arranged according to increasing atomic number.
- According to this, the atomic number increases from left to right and from top to bottom in the periodic table.
- Often, the relative atomic mass increases in parallel with this.
- Horizontal rows in the table are called periods, and vertical rows are called groups.

The Periodic table

Alkali Metals

Alkaline Earths

Transition Metals

Halogens

Main Group

Noble Gases

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 131.29
55 Cs 132.905	56 Ba 137.327	57 *La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	89 †Ac 227.028	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (272)		114 (287)		116 (289)		118 (293)

*Lanthanide series	58 Ce 140.115	59 Pr 140.908	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
†Actinide series	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np 237.048	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 (260)

Main Group

Lanthanides and Actinides

The Periodic Table

- In the periodic table, the vertical groups bring together elements with similar properties.
- The horizontal periods of the table are arranged in order of increasing atomic number from left to right.
- The first two groups the *s* block and the last six groups the *p* block together constitute the *main-group elements*.
- Because they come between the *s* block and the *p* block, the *d* block elements are known as the *transition elements*.
- The *f* block elements, sometimes called the *inner transition elements*, would extend the table to a width of 32 members if incorporated in the main body of the table.
- The table would generally be too wide to fit on a printed page, and so the *f* block elements are extracted from the table and placed at the bottom. The 15 elements following barium are called the *lanthanides*, and the 15 following radon are called the *actinides.*

The period number of an element indicates the highest energy level of electrons that element has.

The group number of an element shows the number of electrons in the final orbital of that element, that is, the valence electrons.

Horizontal Column is Period: 7 periods

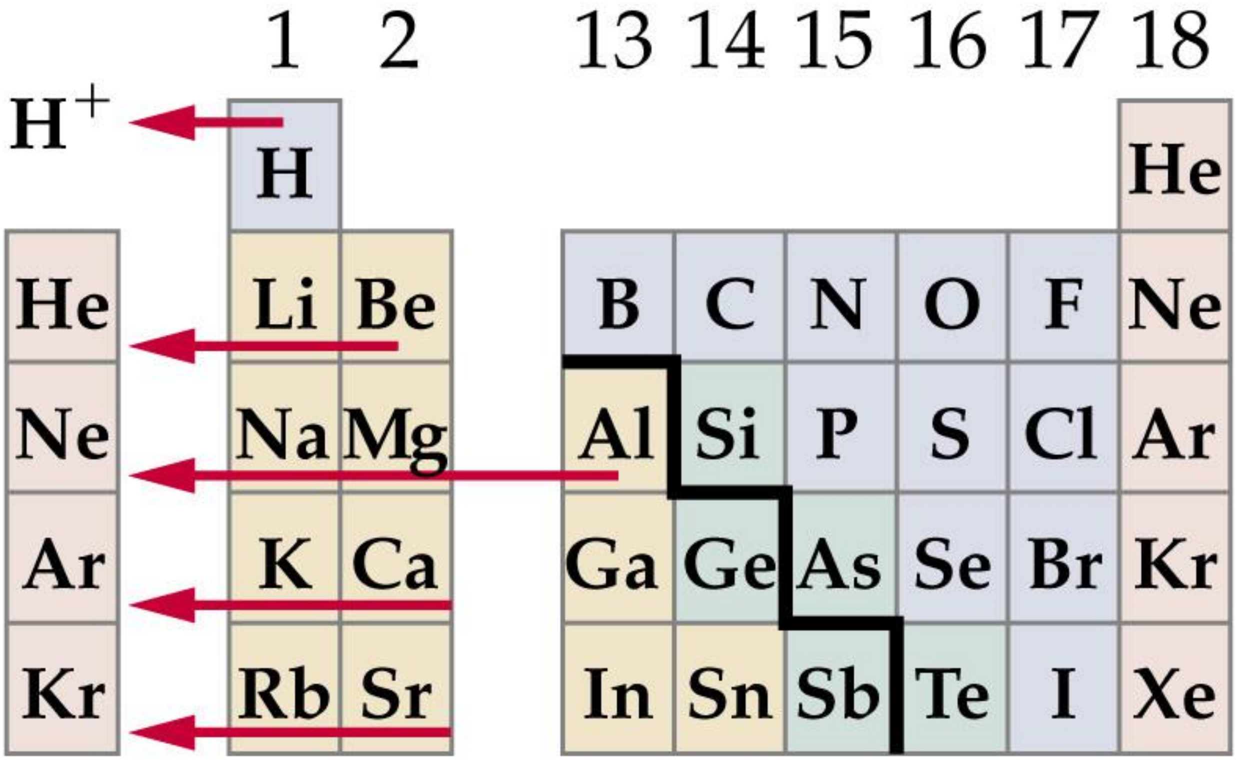
Vertical Column is Group : 8 A groups
8 B groups

B group elements are called transition elements.

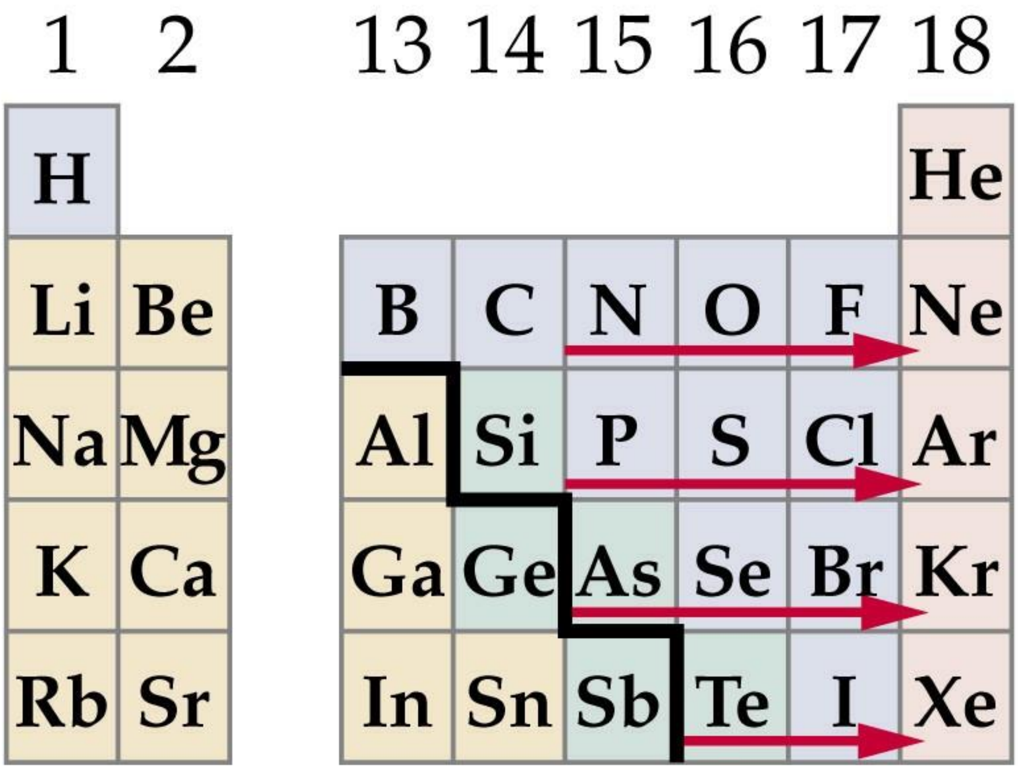
9-2 Metals and Nonmetals and Their Ions

- Metals
  - Good conductors of heat and electricity.
  - Malleable and ductile.
  - Moderate to high melting points.
- Nonmetals
  - Nonconductors of heat and electricity.
  - Brittle solids.
  - Some are gases at room temperature.

Metals Tend to Lose Electrons



Nonmetals Tend to Gain Electrons



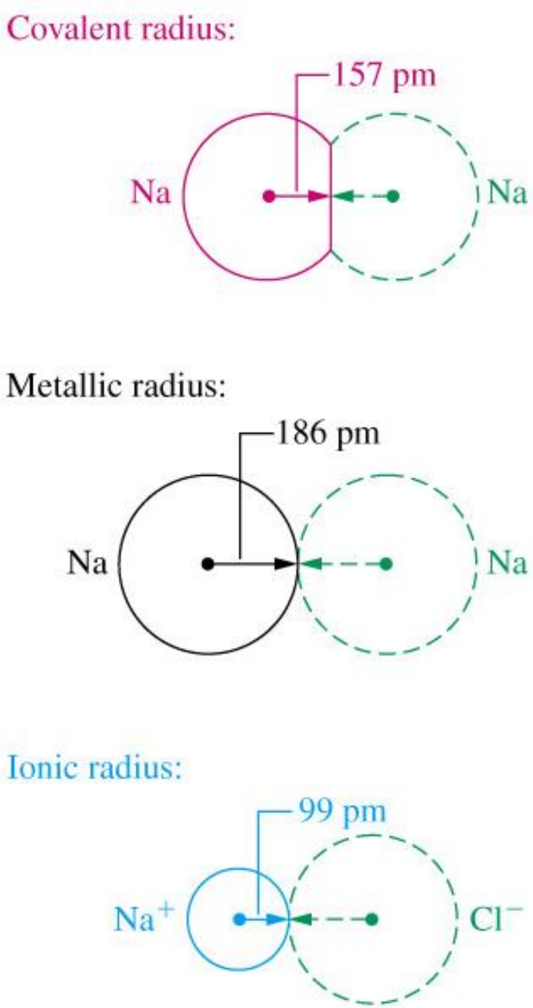
Finding Periods and Groups on the Periodic Table

- To find the period and group of an element in the periodic table, the configuration of the electrons of that element is done.

- In the electron configuration, the last orbital specifies the period. The number of electrons in the last orbital indicates its group.

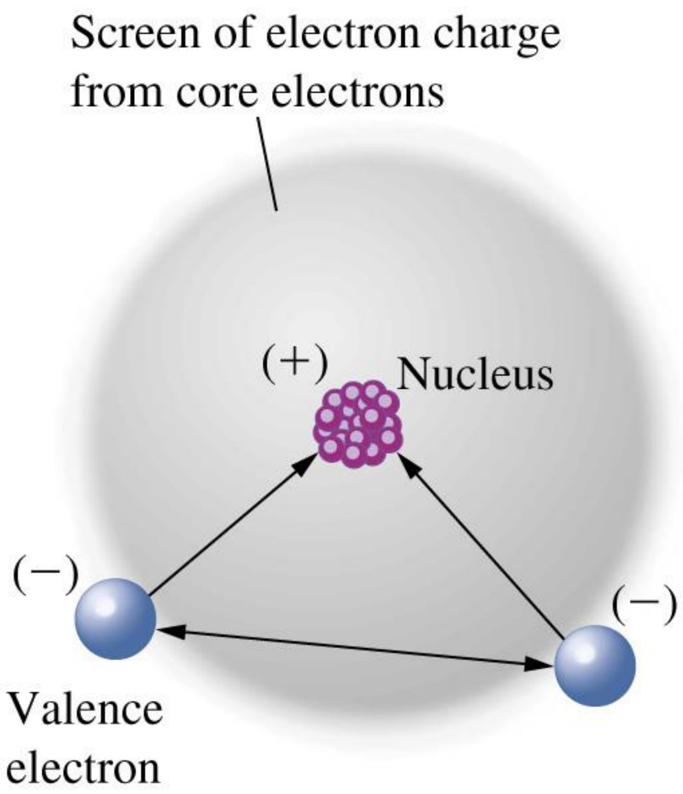
1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
S^1	S^2	d^1	d^2	d^3	d^4	d^5	d^6	d^7	d^8	d^9	d^{10}	P^1	P^2	P^3	P^4	P^5	P^6

9-3 The Sizes of Atoms and Ions



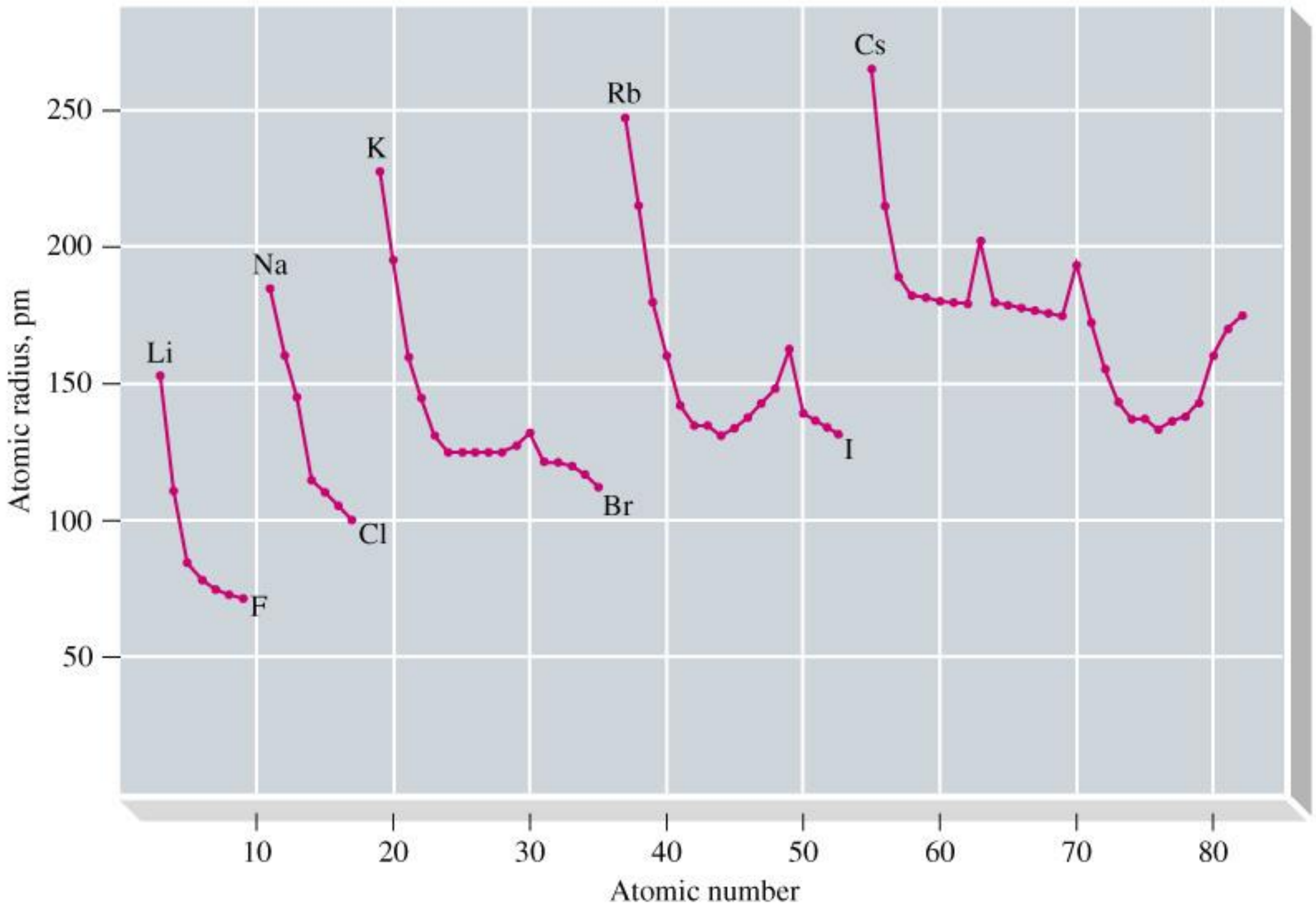
- Atomic Radius: It is determined by the distance between two atoms bonded by chemical bonds.
- Covalent Radius: Half the distance between the nuclei of two atoms bonded by a single covalent bond.
- Ionic Radius: It is the distance between the nuclei of the ionic bonded ions. Since the ions are not of equivalent size, the distance between them must be appropriately divided between the cation and anion.

Screening and Penetration



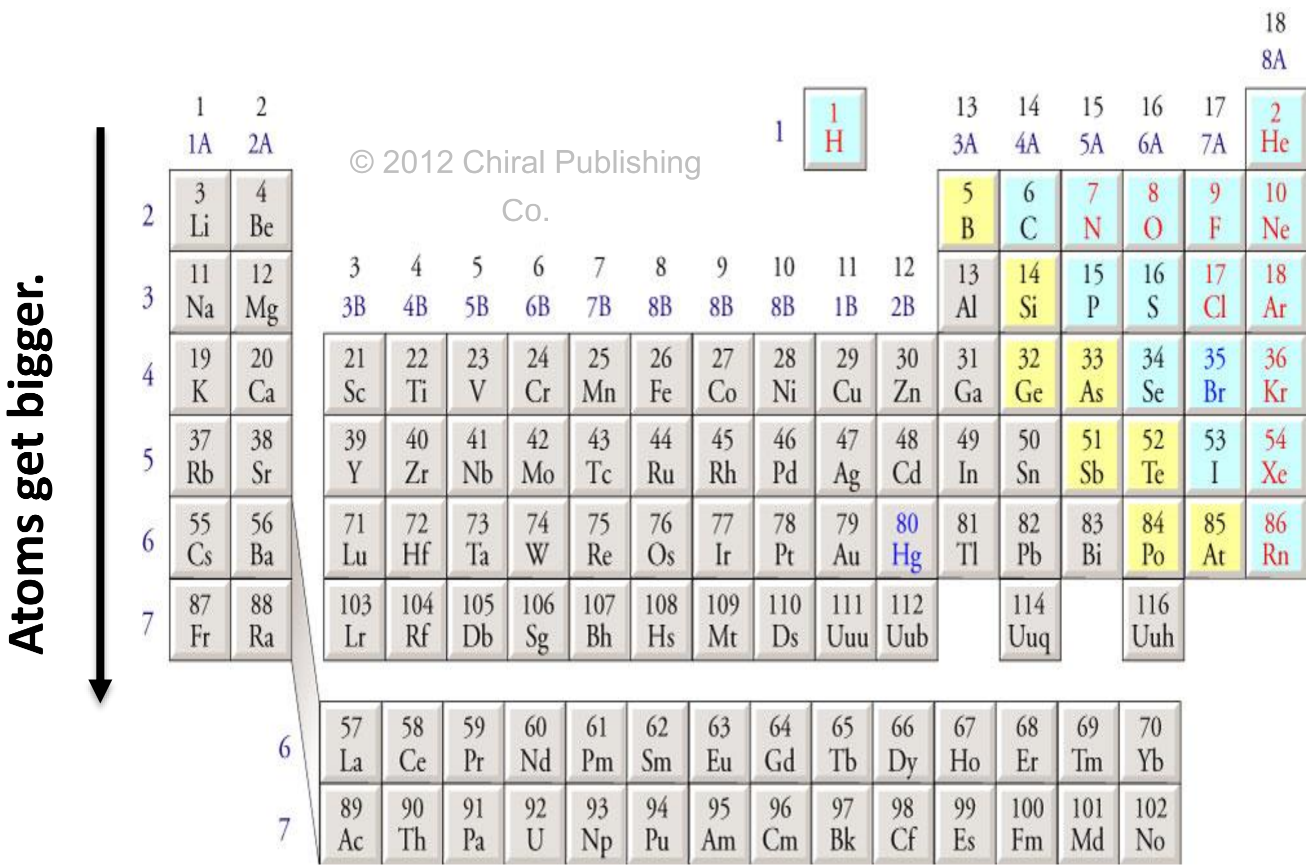
- Penetration was described as a gauge of how close an electron gets to the nucleus. When interpreting the radial probability distributions, we saw that *s* electrons, by virtue of their extra humps of probability close to the nucleus, penetrate better than *p* electrons, which in turn penetrate better than *d* electrons.
- Screening, or shielding, reflects how an outer electron is blocked from the nuclear charge by inner electrons.

Atomic Radius



Sizes of Atoms

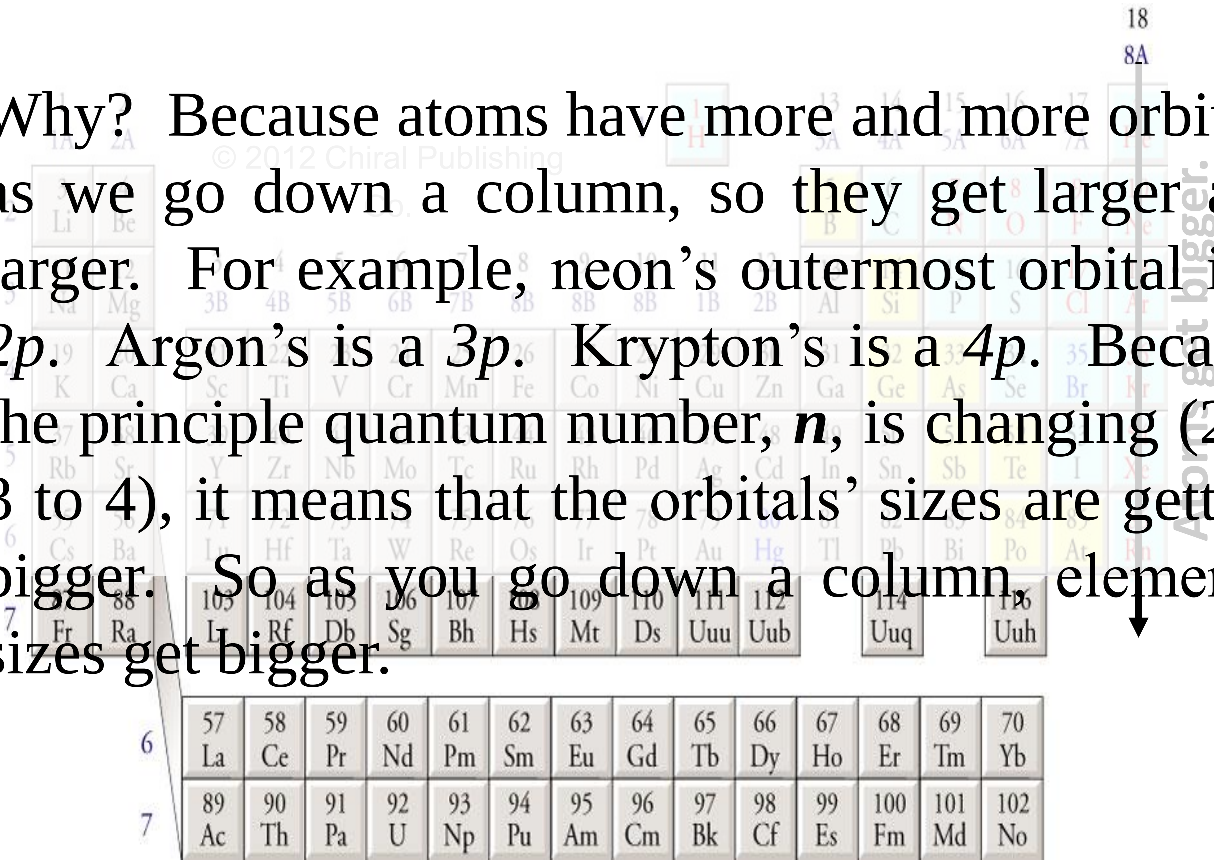
Atoms get bigger as you go down a column (group) on the periodic table:



Sizes of Atoms

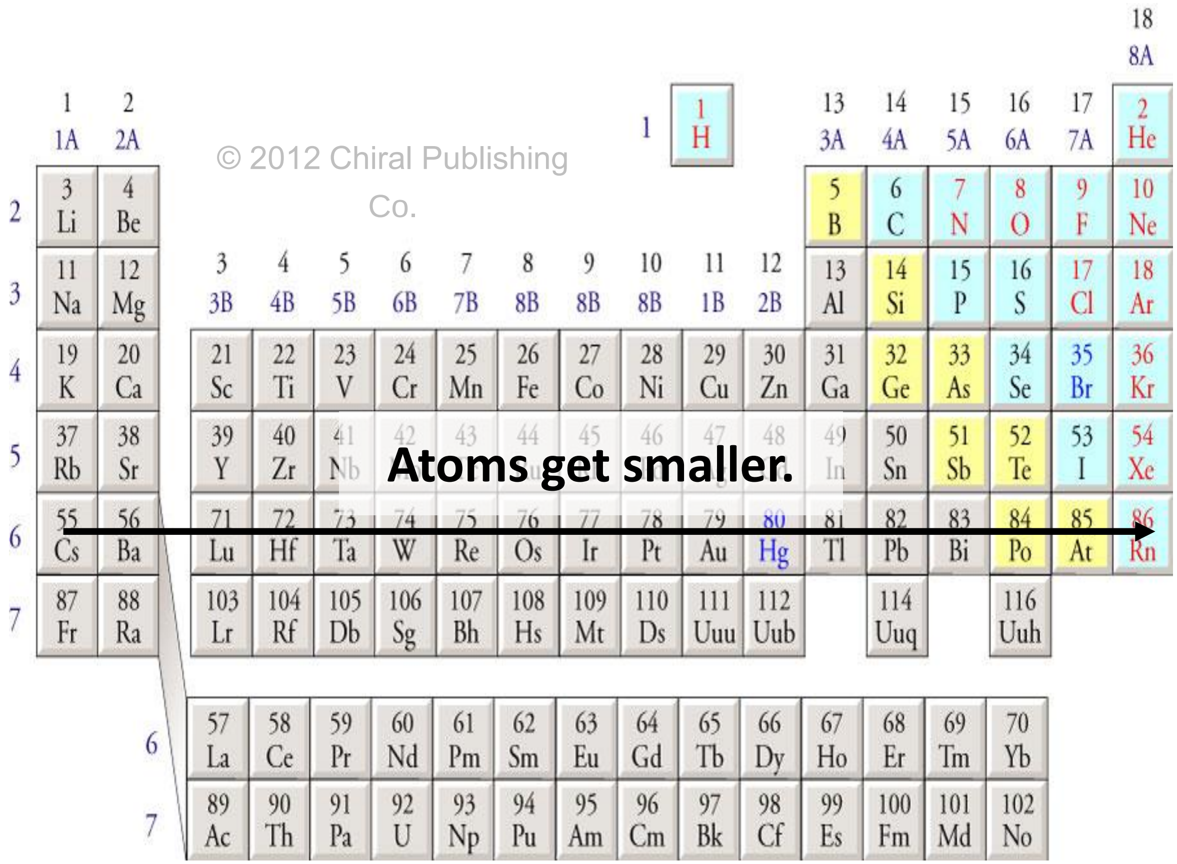
Atoms get bigger as you go down a column (group) on the periodic table:

Why? Because atoms have more and more orbitals as we go down a column, so they get larger and larger. For example, neon's outermost orbital is a *2p*. Argon's is a *3p*. Krypton's is a *4p*. Because the principle quantum number, ***n***, is changing (2 to 3 to 4), it means that the orbitals' sizes are getting bigger. So as you go down a column, elements' sizes get bigger.



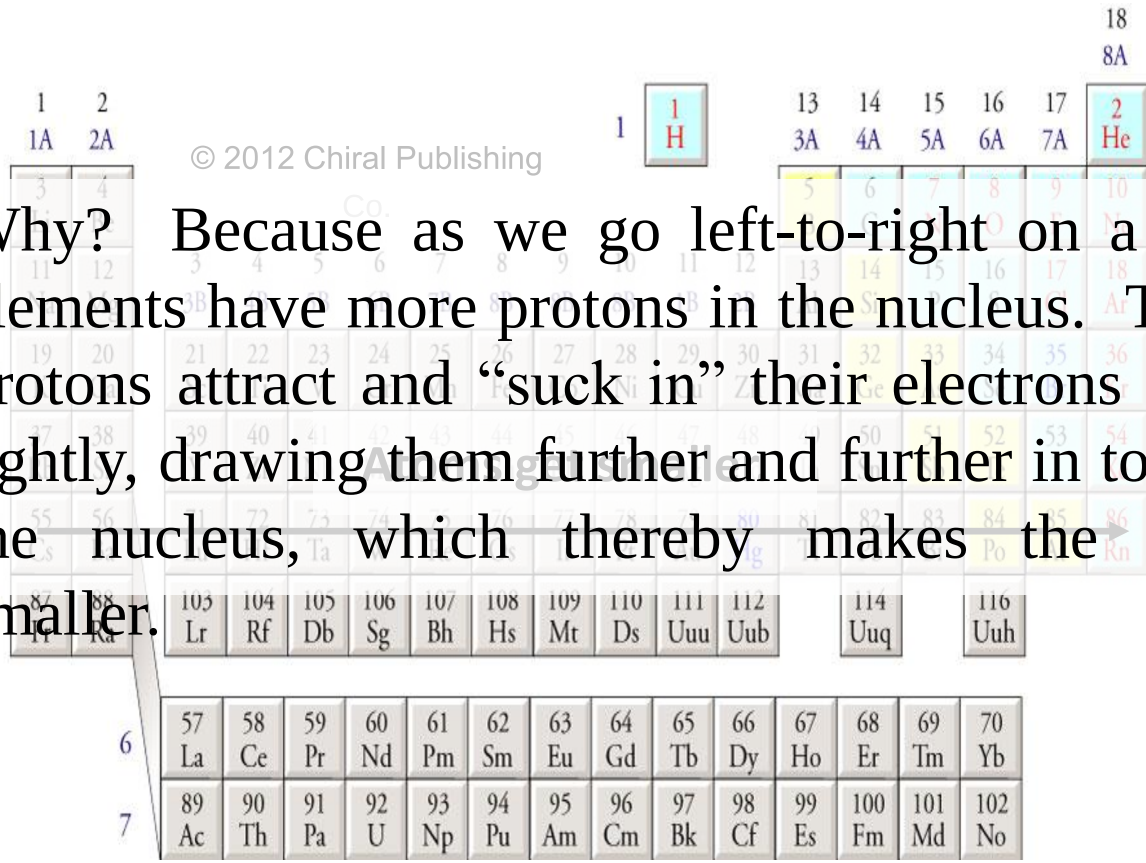
Sizes of Atoms

Atoms get smaller as you go left-to-right across a row (period) on the periodic table:



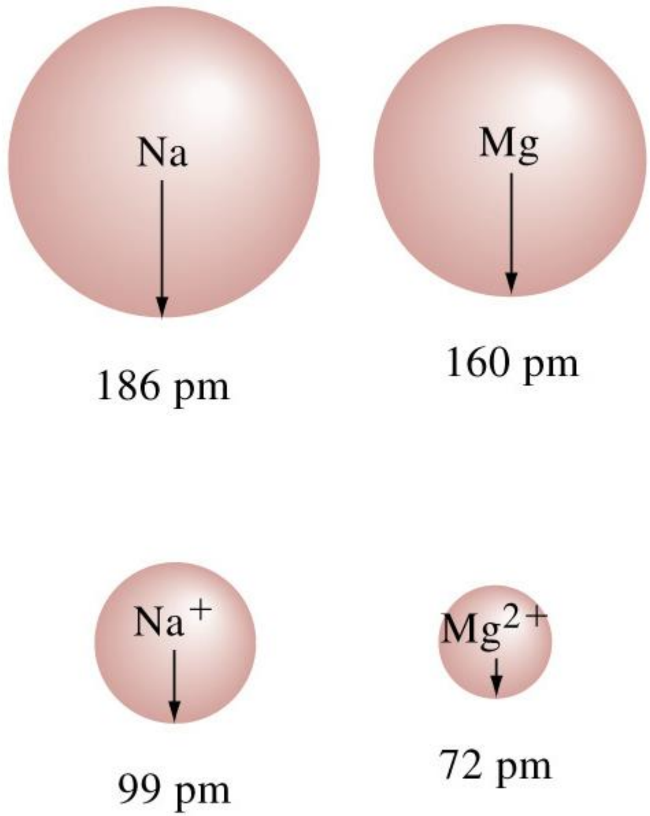
Sizes of Atoms

Atoms get smaller as you go left-to-right across a row (period) on the periodic table:



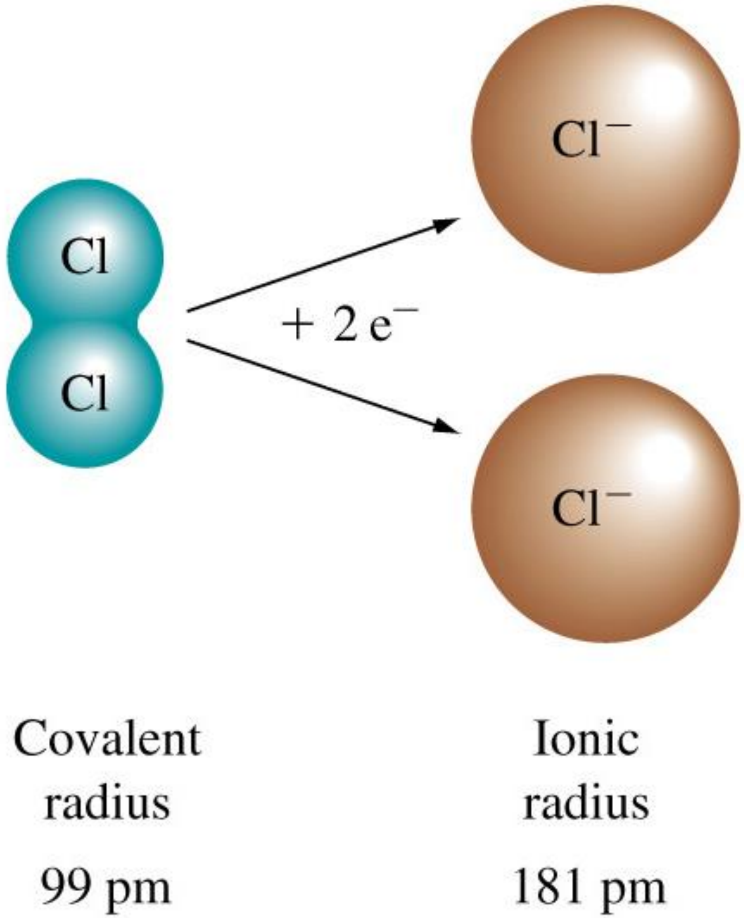
Why? Because as we go left-to-right on a row, elements have more protons in the nucleus. These protons attract and “suck in” their electrons more tightly, drawing them further and further in toward the nucleus, which thereby makes the atom smaller.

Cationic Radius



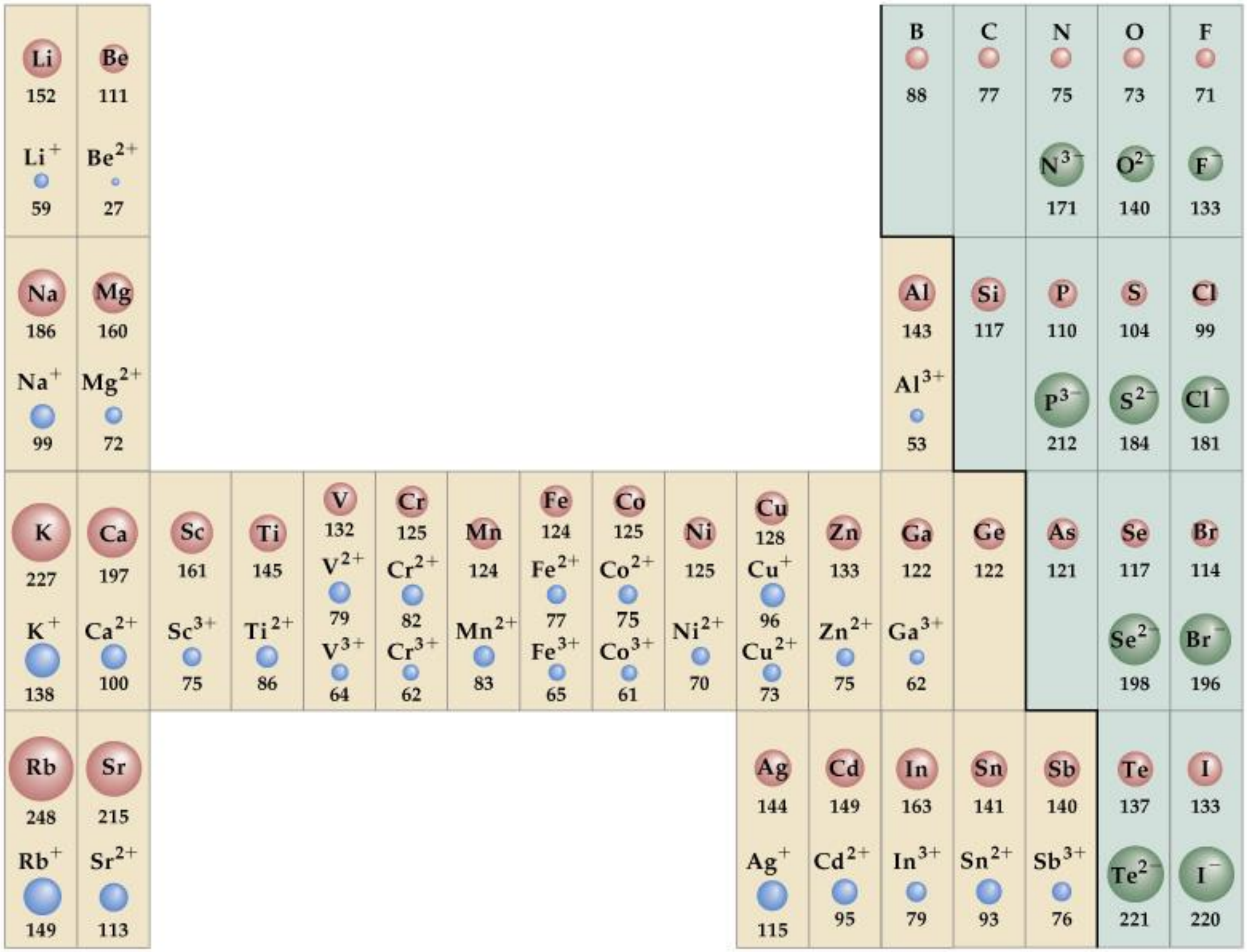
- Cations are smaller than the atoms from which they are formed.
- In cations containing the same number of electrons (isoelectronic), the cation with a larger ionic charge is smaller in size

Anionic Radius



- Anions are larger than the atoms from which they are formed.
- For anions containing equal numbers of electrons, the ion radius increases as the ion charge increases.

Atomic and Ionic Radii



9-4 Ionization Energy

The **ionization energy**, ***I***, is the quantity of energy a *gaseous* atom must absorb to be able to expel an electron. The electron that is lost is the one that is most loosely held.

The *first ionization energy* is the energy required to remove the first electron from a neutral atom.

The *second* ionization energy the energy to strip an electron from a gaseous ion with a charge of 1+.

The larger the ionization energy, the harder it is to remove an electron.

Electrons in higher-energy orbitals are easier to remove, because they're further from the nucleus. For example, it's easier to remove an electron from a 3s orbital than from a 2s orbital.

$Mg(g) \rightarrow Mg^{+}(g) + e^{-}$ $\quad I_1 = 738$ kJ

$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-}$ $\quad I_2 = 1451$ kJ

First Ionization Energy

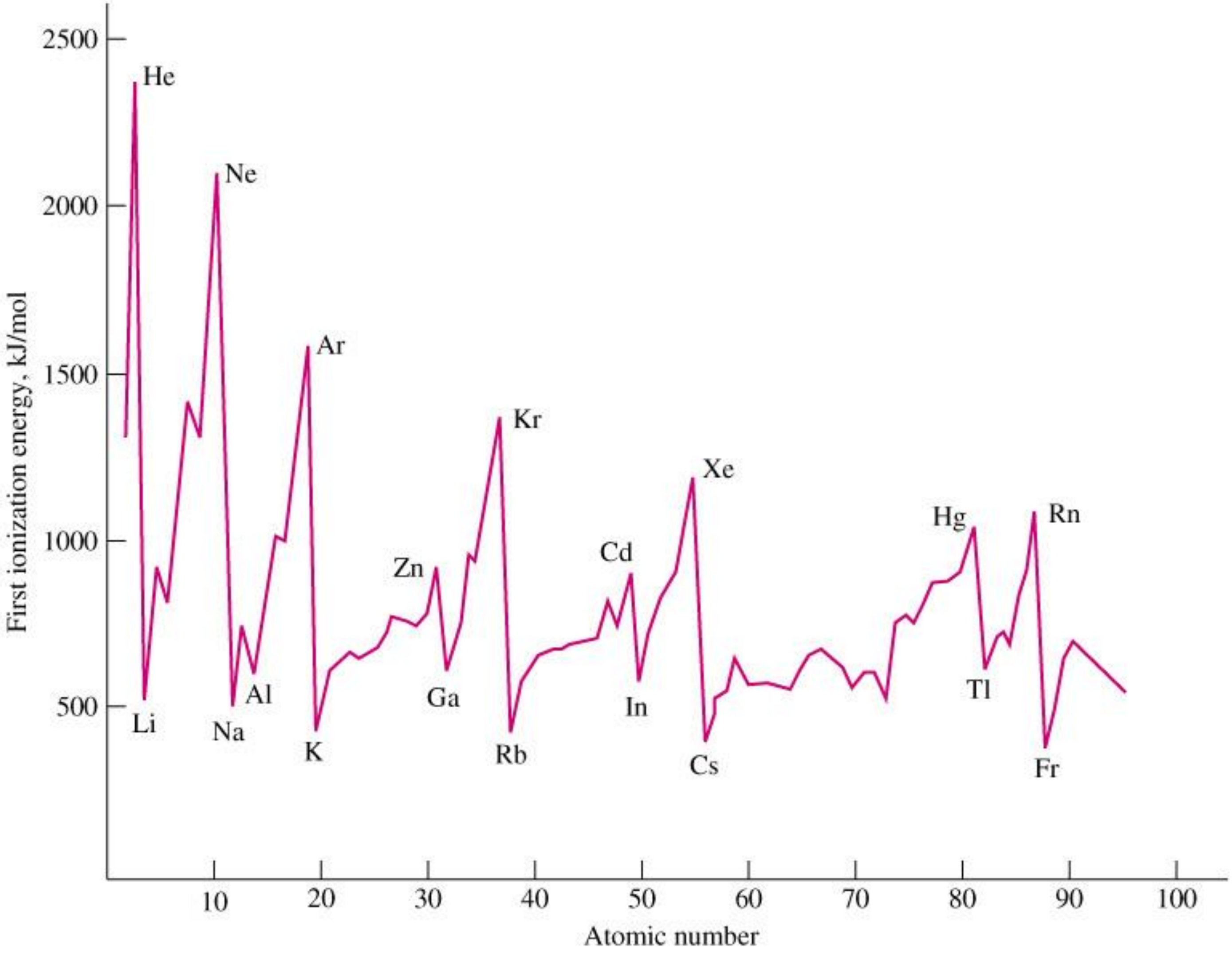


TABLE 10.4 Ionization Energies of the Third-Period Elements (in kJ/mol)

	Na	Mg	Al	Si	P	S	Cl	Ar
I_1	495.8	**737.7**	**577.6**	786.5	**1012**	**999.6**	1251.1	1520.5
I_2	4562	**1451**	1817	1577	1903	2251	2297	2666
I_3		**7733**	2745	3232	2912	3361	3822	3931
I_4			11580	4356	4957	4564	5158	5771
I_5				16090	6274	7013	6542	7238
I_6					21270	8496	9362	8781
I_7						27110	11020	12000

I_2 (Mg) vs. I_3 (Mg) I_1 (Mg) vs. I_1 (Al) I_1 (P) vs. I_1 (S)

9-5 Electron Affinity

The energy change of *adding* an electron to an atom is called the atom's **electron affinity**.

For most elements, energy is given off when an electron is added.

$$F(g) + e^{-} \rightarrow F^{-}(g) \qquad EA = -328\ kJ$$

$$F(1s^22s^22p^5) + e^{-} \rightarrow F^{-}(1s^22s^22p^6)$$

$$Li(g) + e^{-} \rightarrow Li^{-}(g) \qquad EA = -59.6\ kJ$$

Electron Affinity

- In the periodic table, Electron affinity increases from left to right and decreases from top to bottom.

- The electron affinity of a positively charged atom is equal to the ionization energy of a neutral atom.

Electron Affinity

- More than one electron can also be added to atoms. But the second electron addition is endothermic.

$O(g) + e^- \rightarrow O^-(g)$ $\quad$ $EA_1 = -141$ kJ/mol

$O^-(g) + e^- \rightarrow O^{2-}(g)$ $\quad$ $EA_2 = +744$ kJ/mol

- Electron affinities are also related to the size of the atom, as are the ionization energies. This is because the nuclear charge increases as the electron approach the atom.
- For this reason, the electron affinity of the elements in the periodic table increases as you go up and to the right.

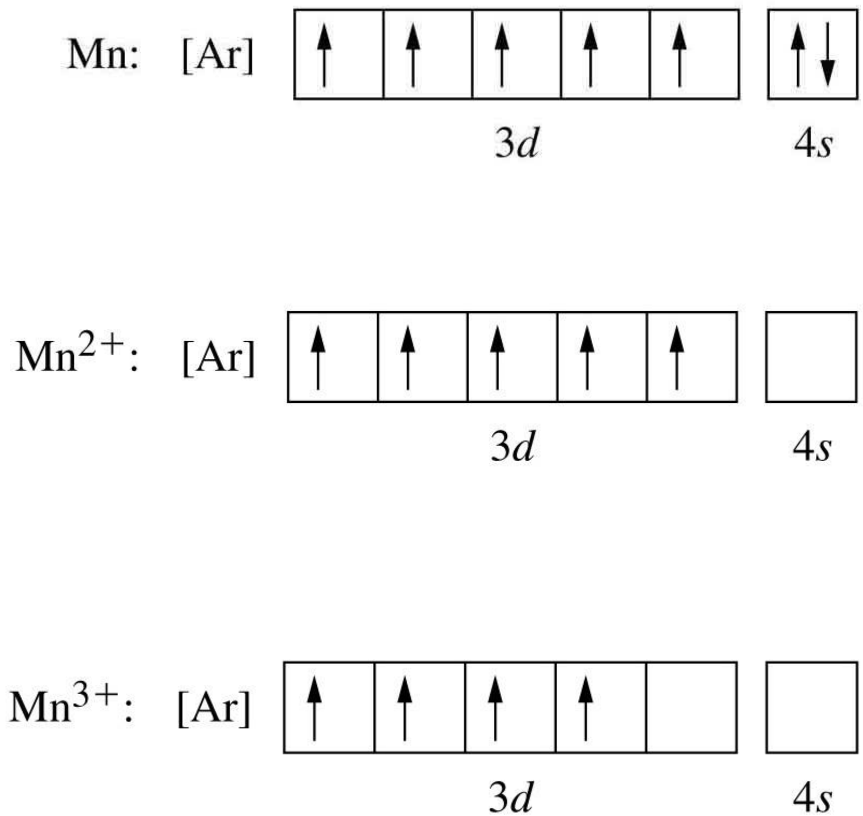
First Electron Affinities

1	2	13	14	15	16	17	18
H −72.8							He - -
Li −59.6	Be - -	B −26.7	C −153.9	N −7	O −141.0	F −328.0	Ne - -
Na −52.9	Mg - -	Al −42.5	Si −133.6	P −72	S −200.4	Cl −349.0	Ar - -
K −48.4	Ca - -	Ga −28.9	Ge −119.0	As −78	Se −195.0	Br −324.6	Kr - -
Rb −46.9	Sr - -	In −28.9	Sn −107.3	Sb −103.2	Te −190.2	I −295.2	Xe - -
Cs −45.5	Ba - -	Tl −19.2	Pb −35.1	Bi −91.2	Po −186	At −270	Rn - -

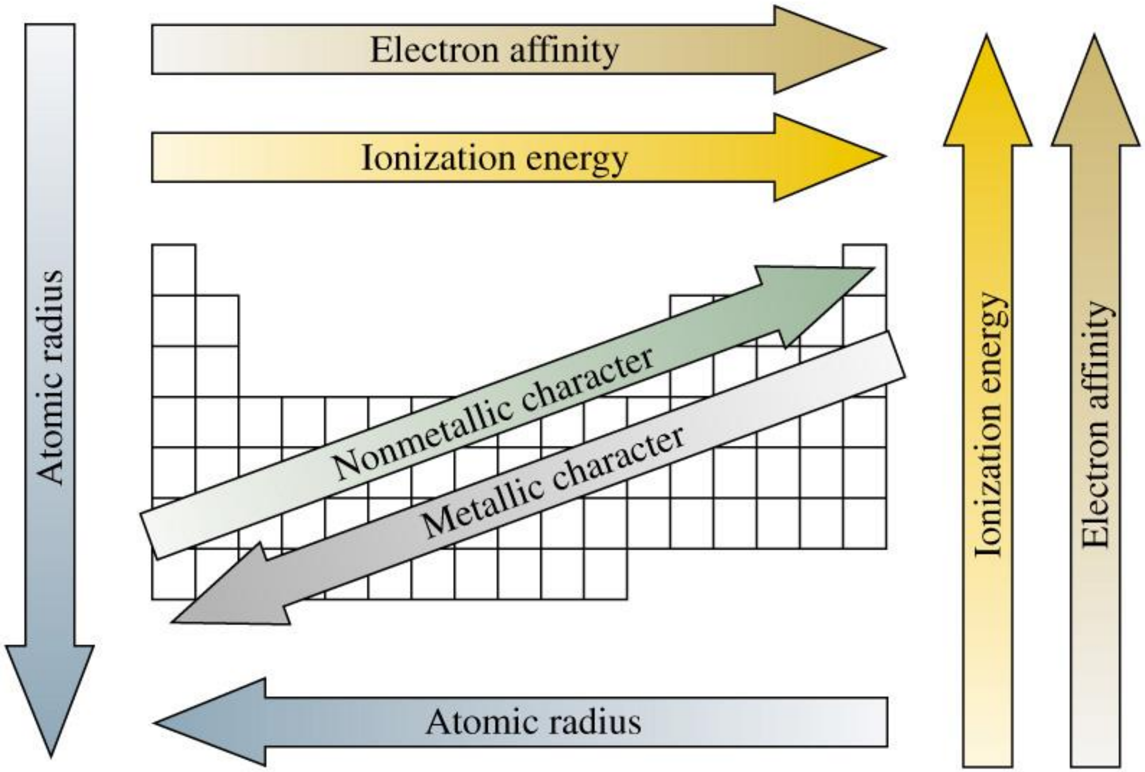
9-6 Magnetic Properties

- Diamagnetic atoms or ions:
  - All e^- are paired.
  - Weakly repelled by a magnetic field.
- Paramagnetic atoms or ions:
  - *Unpaired* e^-.
  - Attracted to an external magnetic field.

Paramagnetism



9-7 Periodic Properties of the Elements



EXAMPLE 9-1

Determine which is the largest atom: $_{21}Sc$, $_{56}Ba$, or $_{34}Se$.

Sc: 4.P 3B

Ba: 6.P 2A

Se: 4.P 6A

Sc and Se are both in the fourth period, and we would expect Sc to be larger than Se because atomic sizes decrease from left to right in a period. Ba is in the sixth period and so has more electronic shells than either Sc or Se. Furthermore, it lies even closer to the left side of the table (group 2) than does Sc (group 3). We can say with confidence that the Ba atom should be the largest of the three.

Ba > Sc > Se

EXAMPLE 9-2

Arrange the following species in order of increasing size:

$_{18}Ar$, $_{19}K^{+}$, $_{17}Cl^{-}$, $_{16}S^{2-}$, $_{20}Ca^{2+}$

The key lies in recognizing that the four species are *isoelectronic*, having the electron configuration of argon.

When considering isoelectronic cations, the higher the charge on the ion, the smaller the ion.

The larger charge on the calcium ion means that Ca^{2+} is smaller than K^{+}. Because K^{+} has a higher nuclear charge than Cl^{-} ($Z = 19$, compared with $Z = 17$), it is smaller than Cl^{-}. For isoelectronic anions, the higher the charge, the larger the ion. S^{2-} is larger than Cl^{-}. The order of increasing size is

$Ca^{2+} > K^{+} > Ar > Cl^{-} > S^{2-}$

EXAMPLE 9-3

Arrange the following species in order of increasing first ionization energy: I_1: As, Sn, Br, Sr.

As: 4.P 5A   Sn: 5.P 4A   Br: 3.P 7A   Sr: 5.P 2A

Of the four atoms, the one that best fits the large-atom category is Sr. Although none of the four atoms is particularly close to the top of the table, Br is the rightest. This fixes the two extremes: Sr with the lowest ionization energy and Br with the highest. A tin atom should be larger than an arsenic atom, and thus Sn should have a lower ionization energy than As. The expected order of *increasing* ionization energies is

Sr < Sn < As < Br

EXAMPLE 9-4

Which of the following would you expect to be diamagnetic and which paramagnetic?

(a) $_{11}Na$ atom **(b)** $_{12}Mg$ atom **(c)** $_{17}Cl^-$ ion **(d)** $_{47}Ag$ atom

(a) $_{11}Na$: ....$3s^1$ Paramagnetic. It has unpaired electron

(b) $_{12}Mg$: ....$3s^2$ Diamagnetic.

(c) $_{17}Cl^-{}_{\mathbf{18}}$: ....$3p^6$ Diamagnetic. All electrons paired

(d) $_{47}Ag$: Paramagnetic. We do not need to work out the exact electron configuration of Ag. Because the atom has 47 electrons an odd number at least one of the electrons must be unpaired.