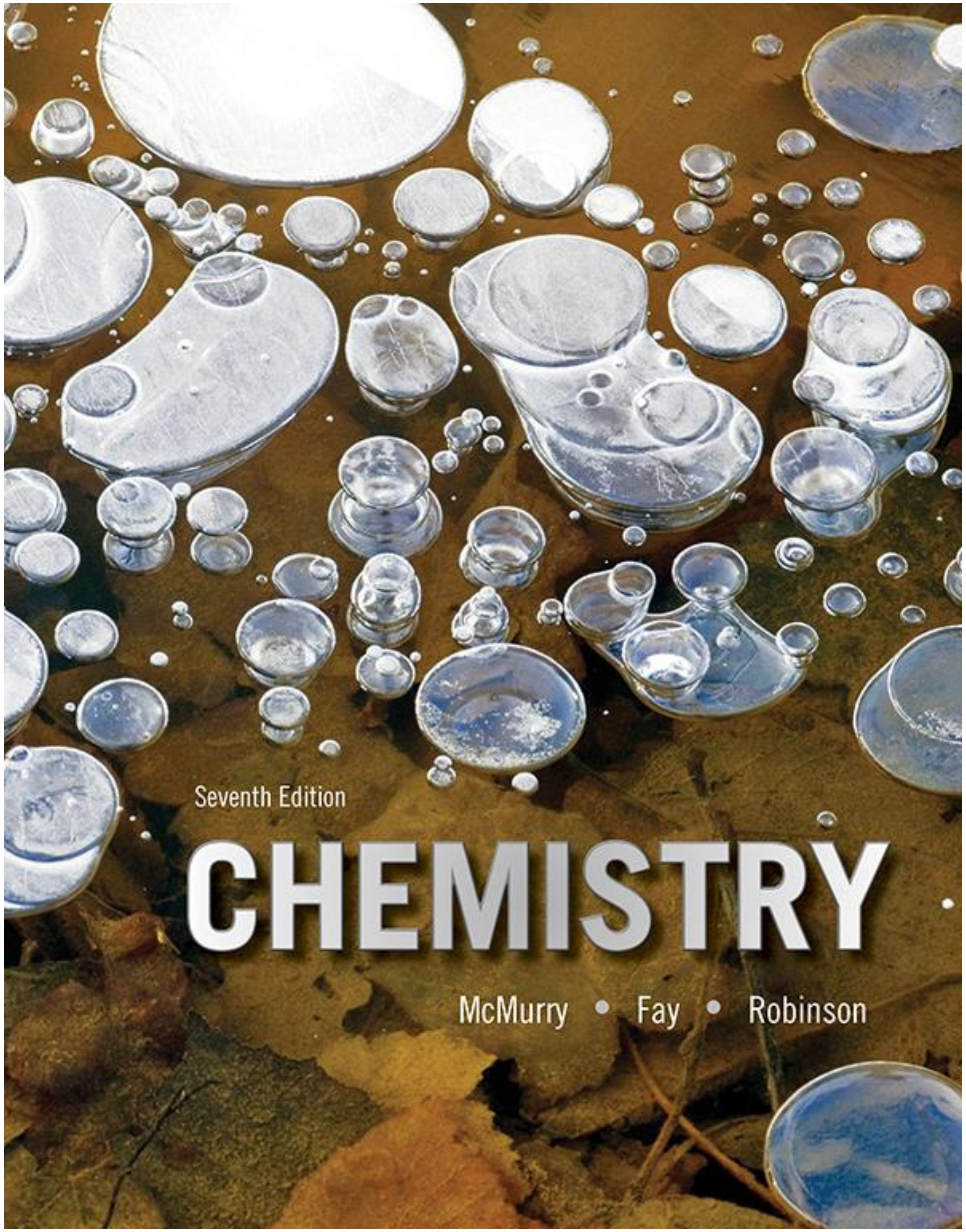

Lecture Presentation

Chapter 6

Ionic Compounds: Periodic Trends and Bonding Theory

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Robert C. Fay

Electron Configurations of Ions

Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$ $-1\ e^-$ → **Na^+:** $1s^2\ 2s^2\ 2p^6$

Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ $+1\ e^-$ → **Cl^-:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Group 1a atom: [Noble gas] ns^1 $-1\ e^-$ →

Group 1a ion$^+$: [Noble gas]

Group 2a atom: [Noble gas] ns^2 $-2\ e^-$ →

Group 2a ion^{2+}: [Noble gas]

Electron Configurations of Ions

Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$ $-1\ e^-$ → **Na⁺:** $1s^2\ 2s^2\ 2p^6$

Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ $+1\ e^-$ → **Cl⁻:** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Group 6a atom: [Noble gas] $ns^2\ np^4$ $+2\ e^-$ →

Group 6a ion^{2-}: [Noble gas] $ns^2\ np^6$

Group 7a atom: [Noble gas] $ns^2\ np^5$ $+1\ e^-$ →

Group 7a ion^{-}: [Noble gas] $ns^2\ np^6$

Electron Configurations of Ions

TABLE 6.1 Some Common Main-Group Ions and Their Noble-Gas Electron Configurations

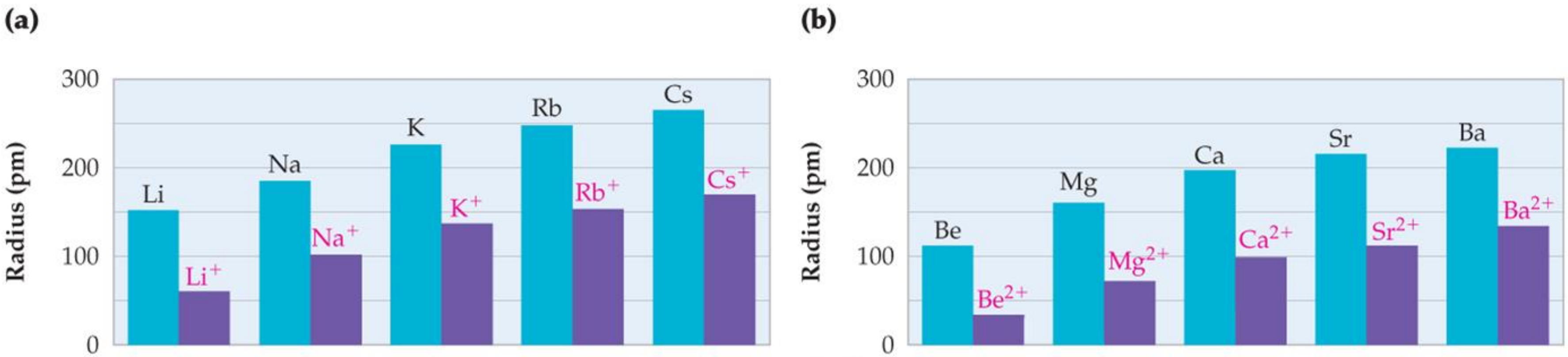
Group 1A	Group 2A	Group 3A	Group 6A	Group 7A	Electron Configuration
H^+					[None]
H^-					[He]
Li^+	Be^{2+}				[He]
Na^+	Mg^{2+}	Al^{3+}	O^{2-}	F^-	[Ne]
K^+	Ca^{2+}	*Ga^{3+}	S^{2-}	Cl^-	[Ar]
Rb^+	Sr^{2+}	*In^{3+}	Se^{2-}	Br^-	[Kr]
Cs^+	Ba^{2+}	*Tl^{3+}	Te^{2-}	I^-	[Xe]

*These ions don't have a true noble-gas electron configuration because they have an additional filled *d* subshell.

Electron Configurations of Ions

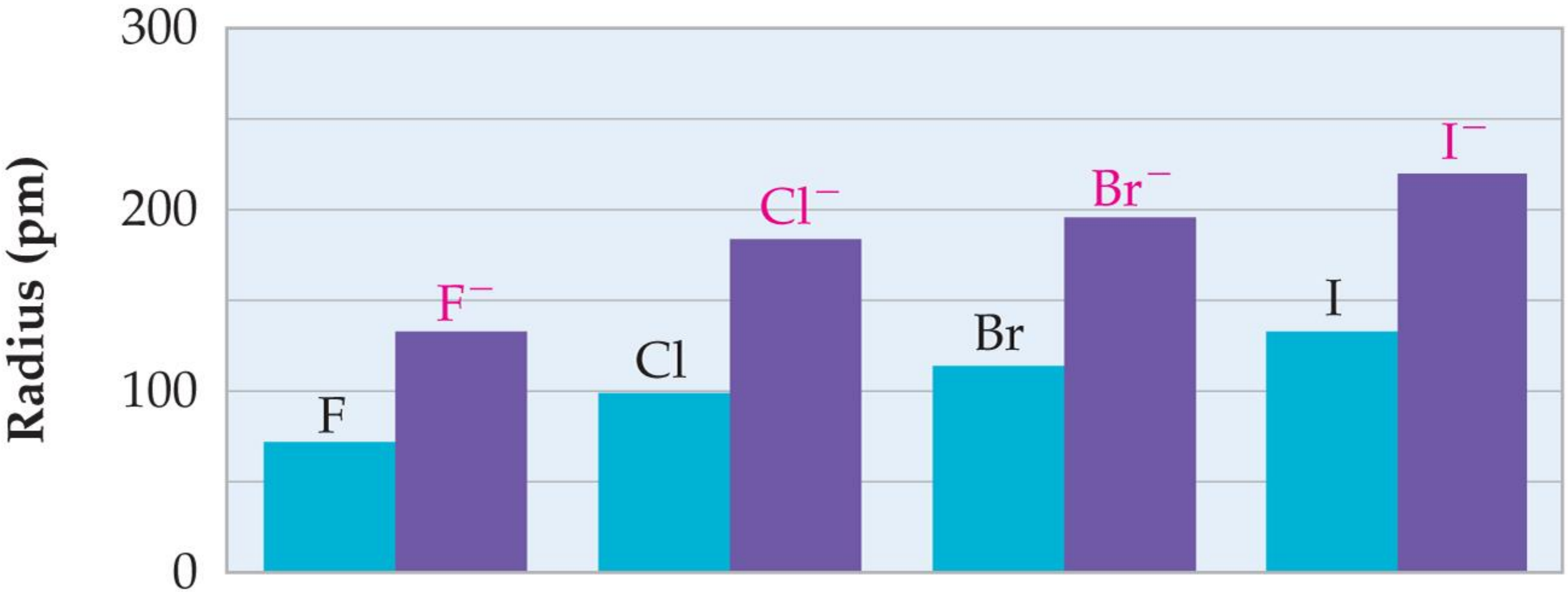
Atoms			Ions
Fe: [Ar] $4s^2\ 3d^6$	$-2\ e^-$	→	**Fe^{2+}:** [Ar] $3d^6$
Fe: [Ar] $4s^2\ 3d^6$	$-3\ e^-$	→	**Fe^{3+}:** [Ar] $3d^5$

Ionic Radii



Cations are smaller than the corresponding **neutral atoms**, both because the principal quantum number of the valence-shell electrons is smaller for the cations than it is for the neutral atoms and because Z_{eff} is larger.

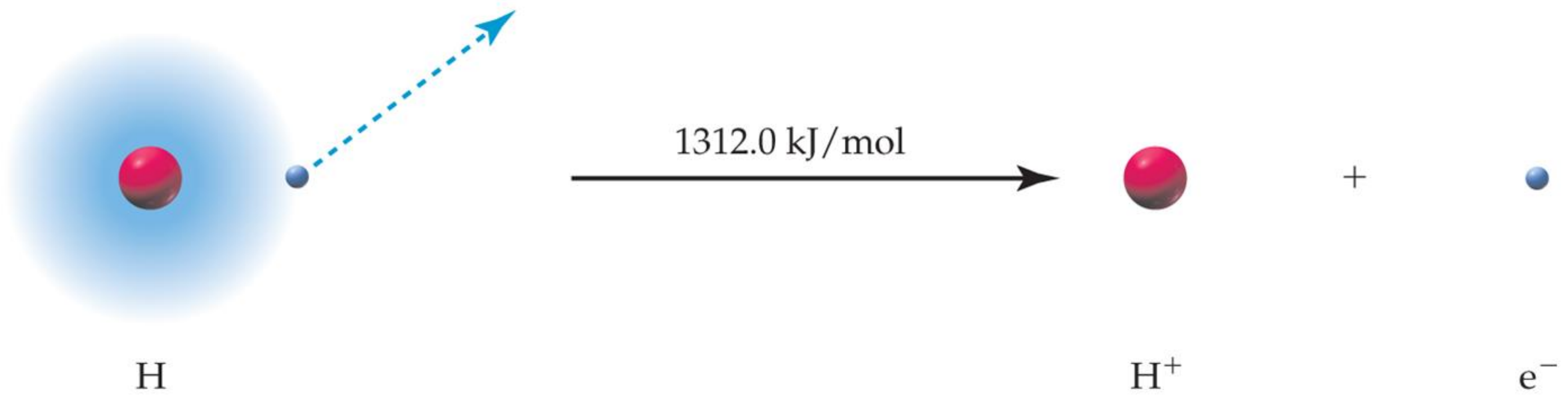
Ionic Radii



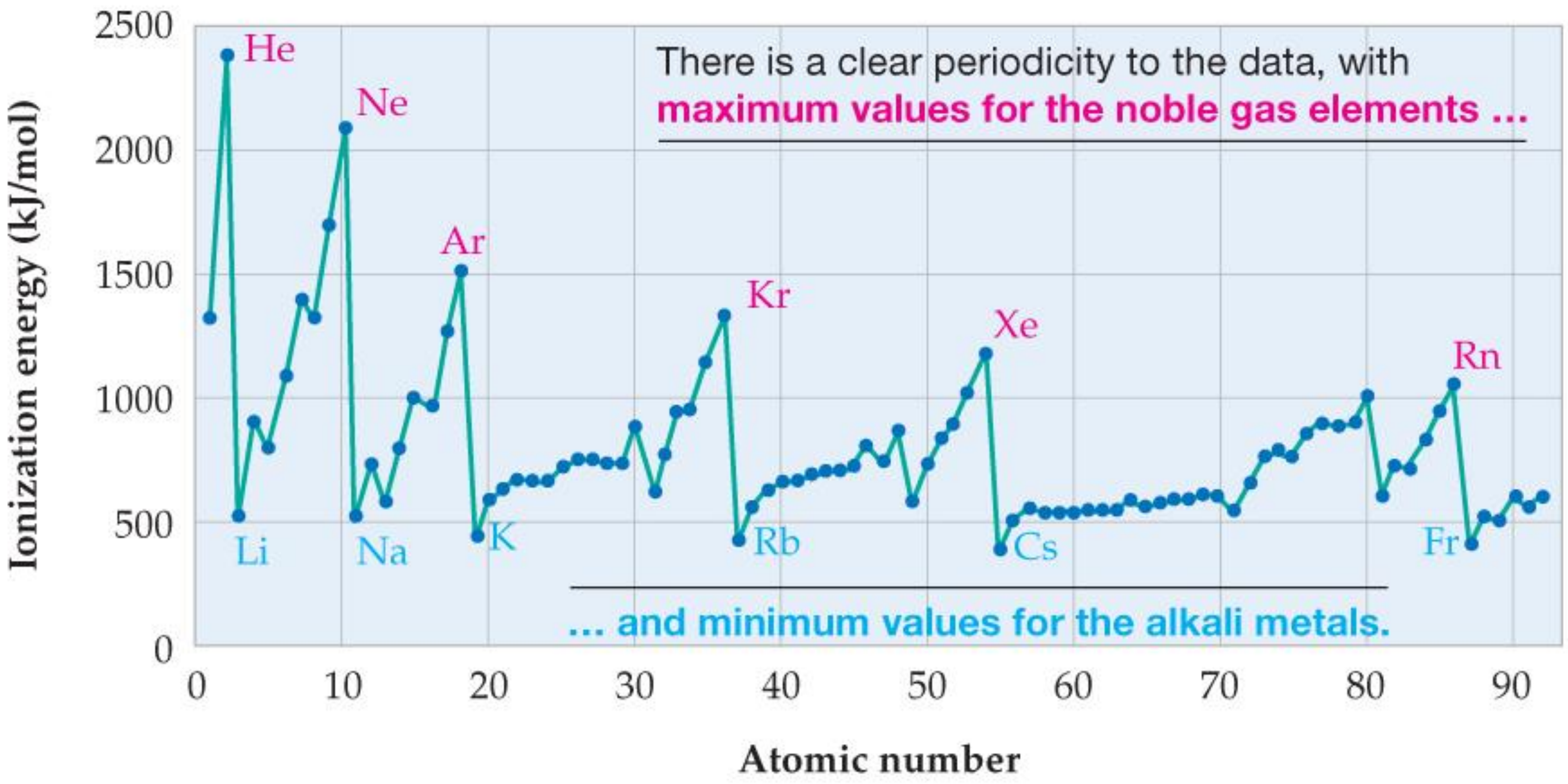
Anions are larger than their neutral atoms because of additional electron–electron repulsions and a decrease in Z_{eff}.

Ionization Energy

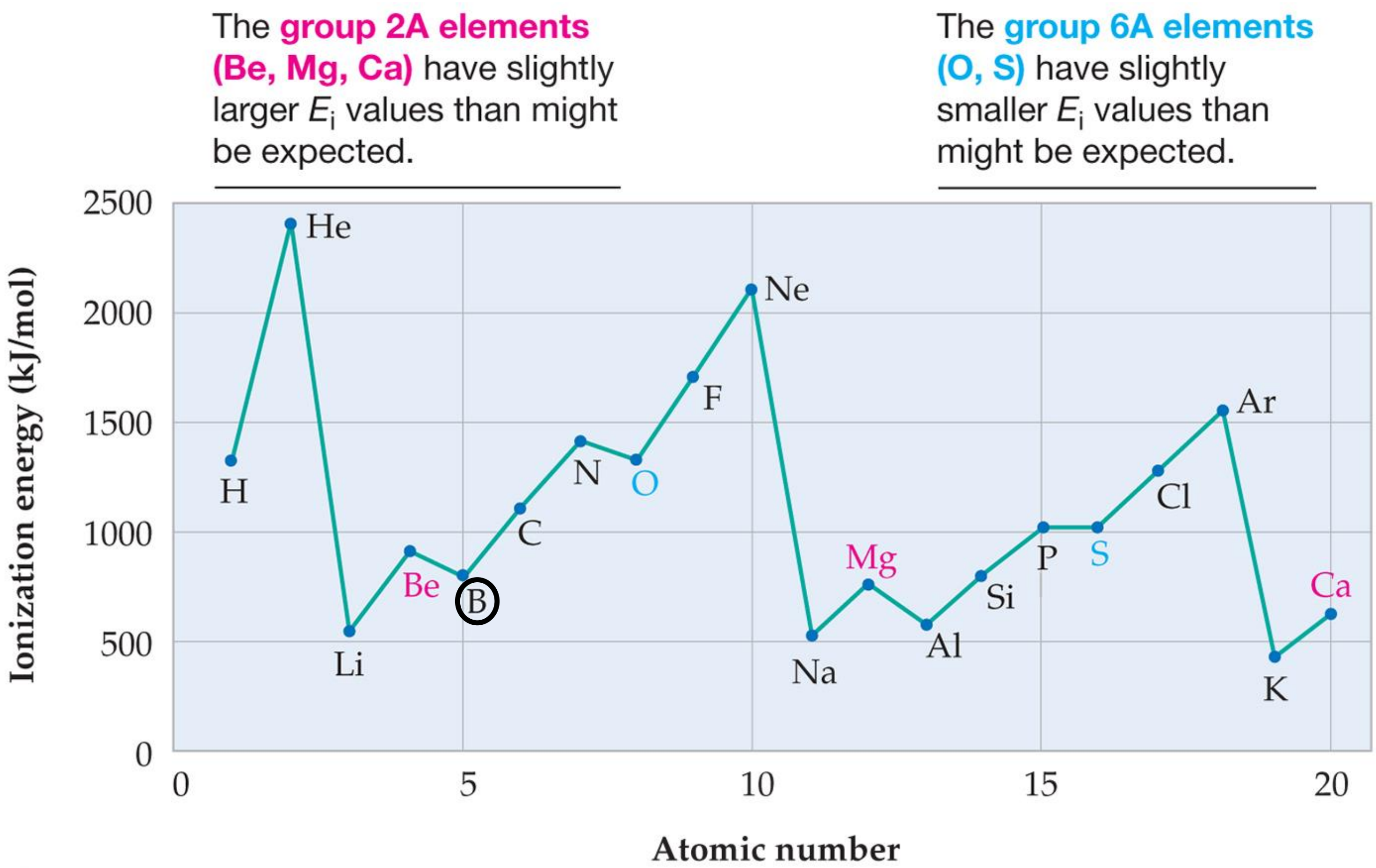
Ionization Energy (E_i): The amount of energy necessary to remove the highest-energy electron from an isolated neutral atom in the gaseous state



Ionization Energy

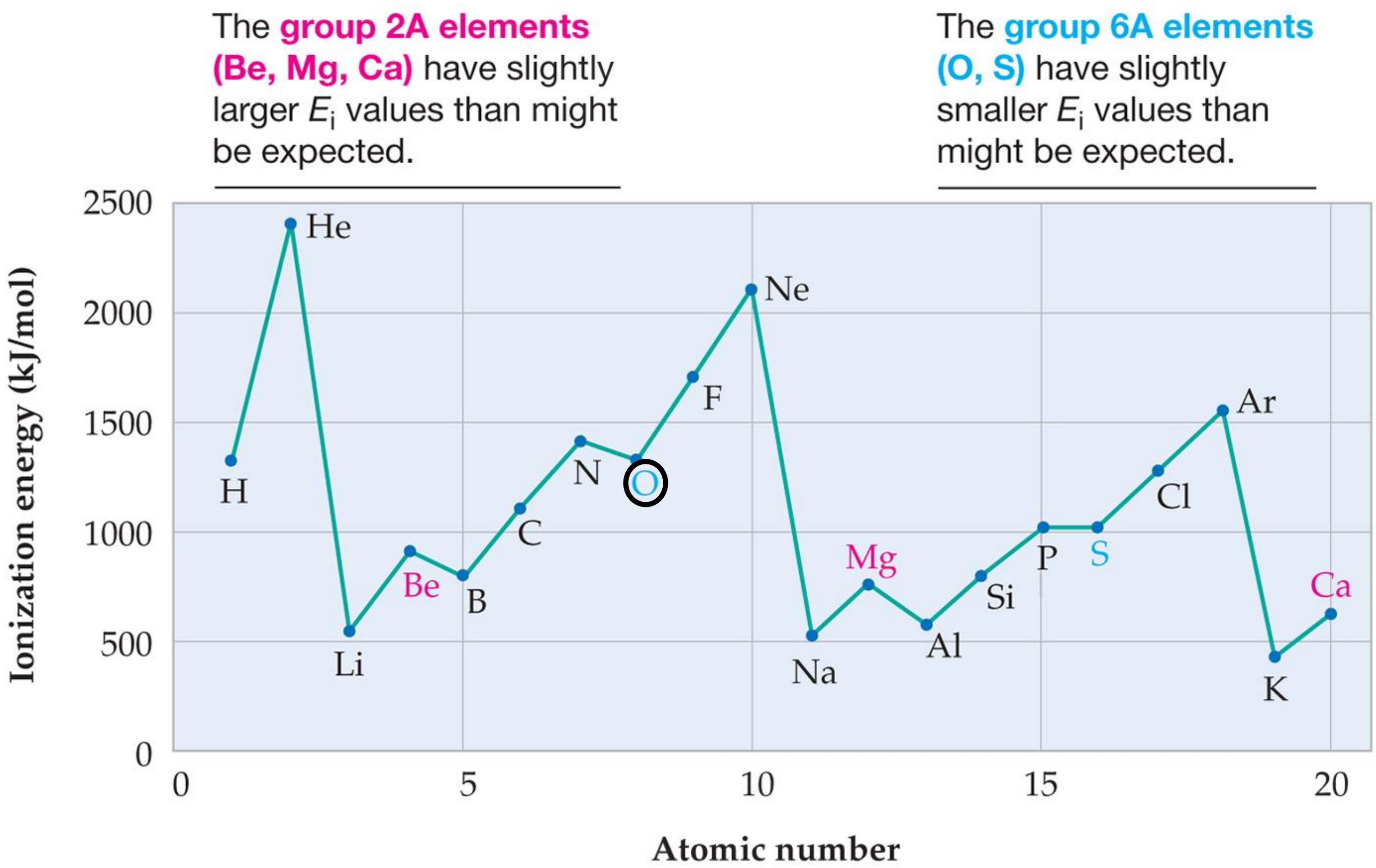


Ionization Energy



Boron has a lower E_i due to a smaller Z_{eff} (shielding by the 2*s* electrons).

Ionization Energy



Oxygen has a lower E_i, since the first electron is removed from a filled orbital.

Higher Ionization Energies

$$M + energy \longrightarrow M^{+} + e^{-}$$

$$M^{+} + energy \longrightarrow M^{2+} + e^{-}$$

$$M^{2+} + energy \longrightarrow M^{3+} + e^{-}$$

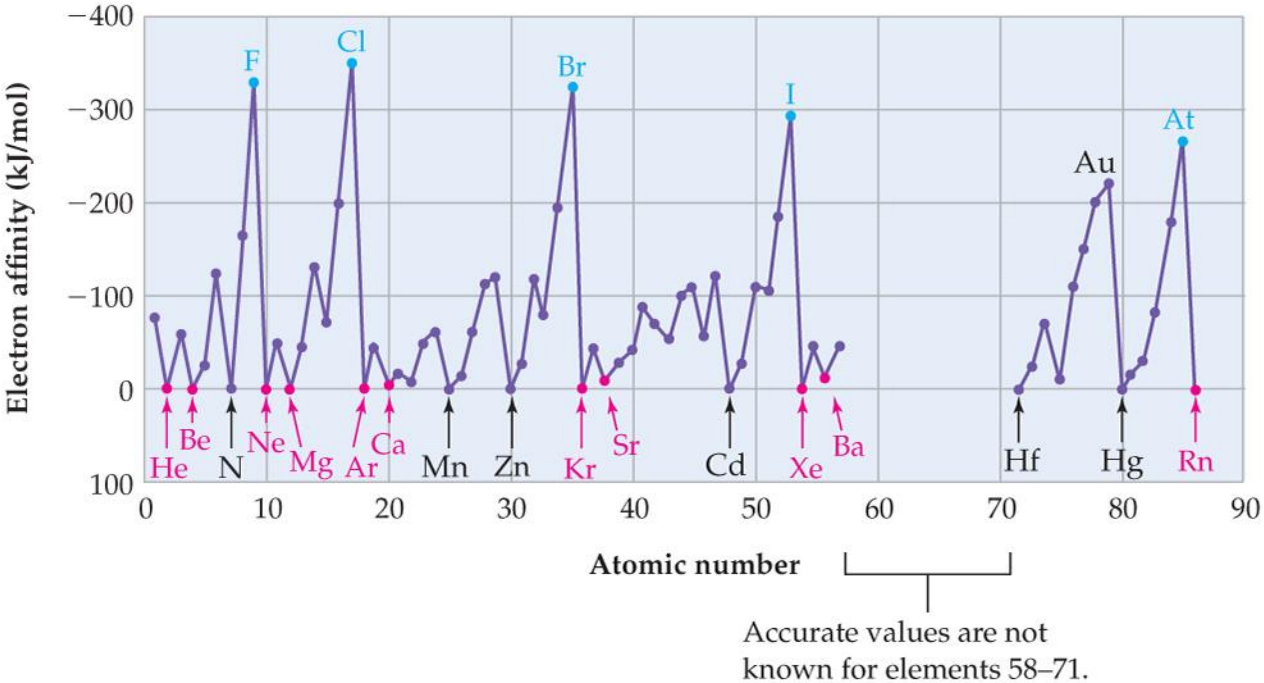
TABLE 6.2 Higher Ionization Energies (kJ/mol) for Main-Group Third-Row Elements

Group	1A	2A	3A	4A	5A	6A	7A	8A
E_i Number	Na	Mg	Al	Si	P	S	Cl	Ar
E_{i1}	496	738	578	787	1,012	1,000	1,251	1,520
E_{i2}	4,562	1,451	1,817	1,577	1,903	2,251	2,297	2,665
E_{i3}	6,912	7,733	2,745	3,231	2,912	3,361	3,822	3,931
E_{i4}	9,543	10,540	11,575	4,356	4,956	4,564	5,158	5,770
E_{i5}	13,353	13,630	14,830	16,091	6,273	7,013	6,540	7,238
E_{i6}	16,610	17,995	18,376	19,784	22,233	8,495	9,458	8,781
E_{i7}	20,114	21,703	23,293	23,783	25,397	27,106	11,020	11,995

The zigzag line marks the large jumps in ionization energies.

Electron Affinity

Electron Affinity (E_{ea}): The energy change that occurs when an electron is added to an isolated atom in the gaseous state



A negative value for E_{ea}, such as those for the **group 7A elements (halogens)**, means that energy is released when an electron adds to an atom.

A value of zero, such as those for the **group 2A elements (alkaline earths) and group 8A elements (noble gases)**, means that energy is absorbed but the exact amount can't be measured.

Octet Rule

Octet rule: Main-group elements tend to undergo reactions that leave them with eight outer-shell electrons. That is, main-group elements react so that they attain a noble-gas electron configuration with filled *s* and *p* sublevels in their valence electron shell.

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Metals tend to have **low** E_i and **low** E_{ea}.
They tend to *lose* one or more electrons.

Nonmetals tend to have **high** E_i and **high** E_{ea}.
They tend to *gain* one or more electrons.

Ionic Bonds and the Formation of Ionic Solids

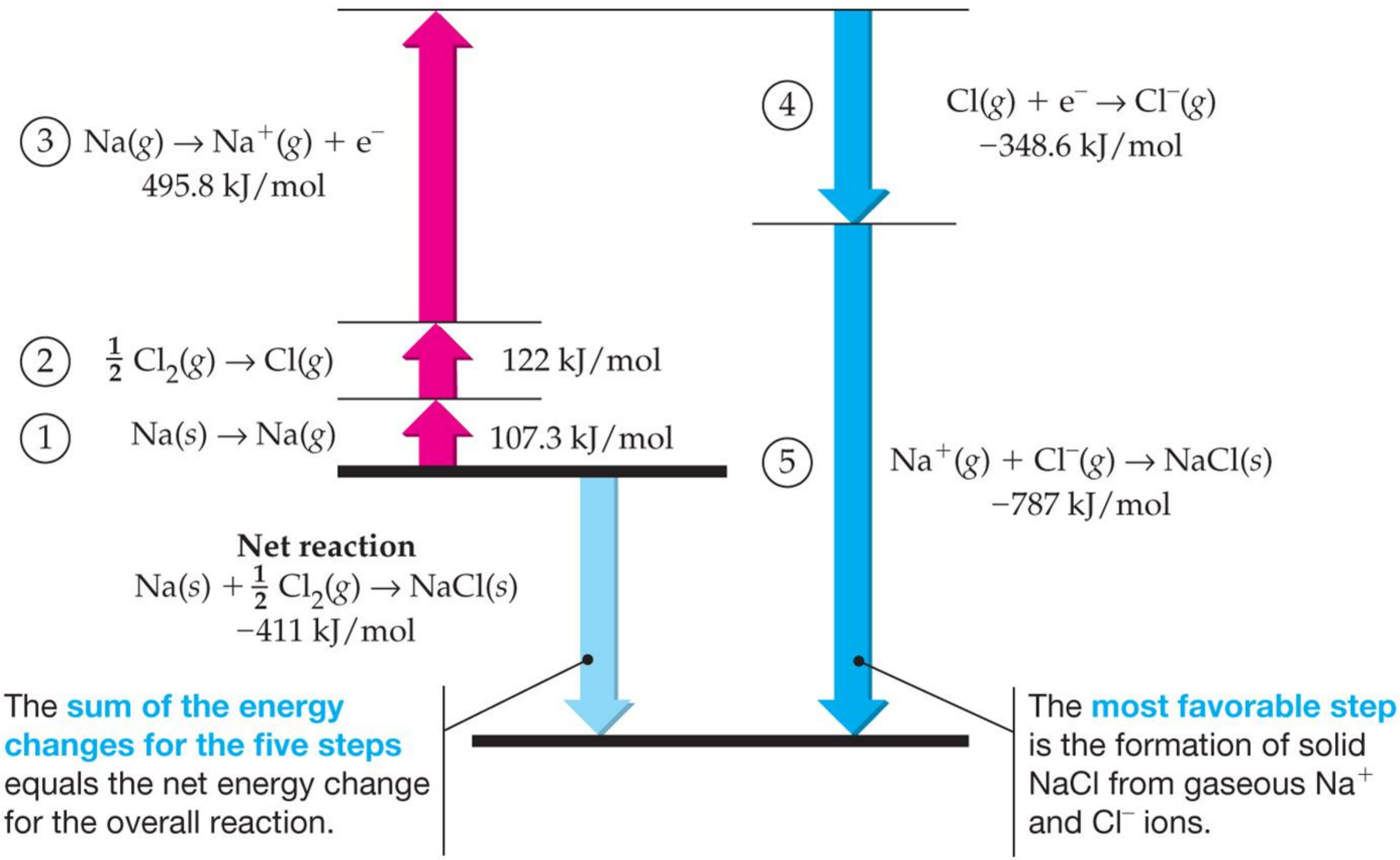
$1s^2\ 2s^2\ 2p^6\ 3s^1$ $\quad$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

$$\mathrm{Na} + \mathrm{Cl} \longrightarrow \mathrm{Na}^+ + \mathrm{Cl}^-$$

$1s^2\ 2s^2\ 2p^6$ $\quad$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Ionic Bonds and the Formation of Ionic Solids

Born-Haber Cycle



Ionic Bonds and the Formation of Ionic Solids

Born-Haber Cycle

Step 1:	$Na(s) \longrightarrow Na(g)$		+107.3 kJ/mol
Step 2:	$\frac{1}{2}Cl_2(g) \longrightarrow Cl(g)$		+122 kJ/mol
Step 3:	$Na(g) \longrightarrow Na^+(g) + e^-$		+495.8 kJ/mol
Step 4:	$Cl(g) + e^- \longrightarrow Cl^-(g)$		–348.6 kJ/mol
Step 5:	$Na^+(g) + Cl^-(g) \longrightarrow NaCl(s)$		–787 kJ/mol
	$Na(s) + Cl_2(g) \longrightarrow NaCl(s)$		–411 kJ/mol

Lattice Energies in Ionic Solids

Lattice Energy (*U*): The amount of energy that must be supplied to break up an ionic solid into individual gaseous ions

TABLE 6.3 Lattice Energies of Some Ionic Solids (kJ/mol)

	Anion				
Cation	**F^-**	**Cl^-**	**Br^-**	**I^-**	**O^{2-}**
Li^+	1036	853	807	757	2925
Na^+	923	787	747	704	2695
K^+	821	715	682	649	2360
Be^{2+}	3505	3020	2914	2800	4443
Mg^{2+}	2957	2524	2440	2327	3791
Ca^{2+}	2630	2258	2176	2074	3401