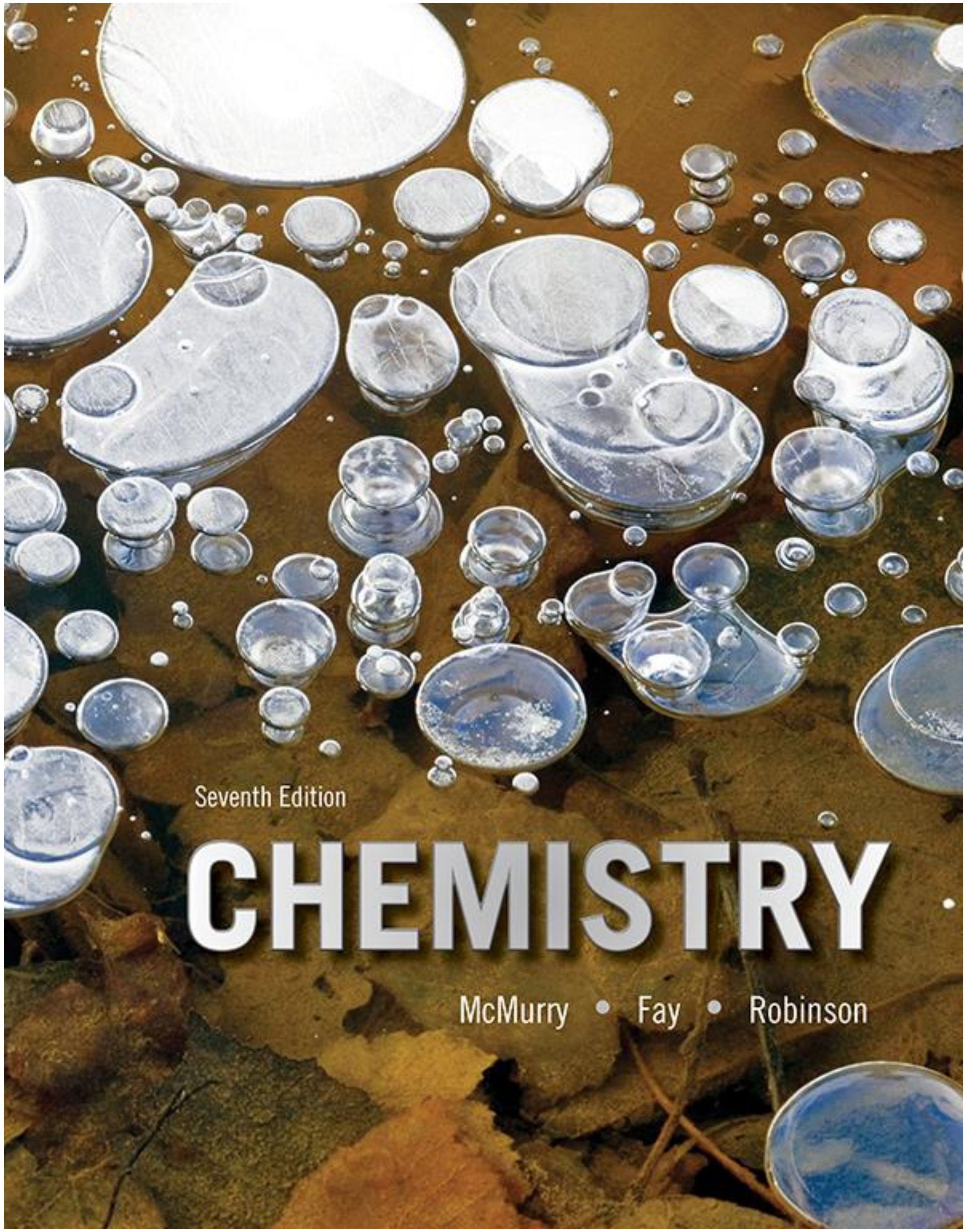


Lecture Presentation

Chapter 6

Ionic Compounds: Periodic Trends & Bonding Theory

HW: 6.1, 6.2, 6.3, 6.5, 6.6, 6.7, 6.9, 6.10, 6.11, 6.12, 6.13, 6.16, 6.20, 6.22, 6.24, 6.26, 6.38, 6.40, 6.46, 6.52, 6.62, 6.68

John E. McMurry
Robert C. Fay

Electron Configurations of Ions (for main group elements)

stable ion = closest noble gas electron configuration (to octet)

metal – lose electrons (group 1A to group 3A)
(+1 lose one electron to +3 lose 3 electrons)

non metal – gain electrons (group 5A to group 7A)
(-3 gain 3 electrons to -1 gain 1 electron)

+ charge = remove electron until get electron configuration of closest prior noble gas

- charge = add electron until get electron configuration of closest next noble gas

Electron Configurations of Ions

Ne

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

Ar

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

Ne

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

Ar

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.

HW: Electron Configurations of Ions (for main group elements)

metal – lose electrons to nearest noble gas (+ charge)
non metal – gain electrons to nearest noble gas (- charge)

What is the electron configuration of F ?

What is the electron configuration of F^{-1} ?

What is the electron configuration of Mg ?

What is the electron configuration of Mg^{+2} ?

End
10/30 D
section

Electron Configurations of Ions (transition metals)

stable ion = lose electrons to half filled d or lose only s electrons (for early TM), usually lose s electron before d electron

Many transition metals have multiple charges possible

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] $3d^5$

. End 10/31 Thursday D section

HW: Electron Configurations of Ions (for transition metal elements)

Usually lose s electrons first. Lose d electrons to half full or empty d subshell,

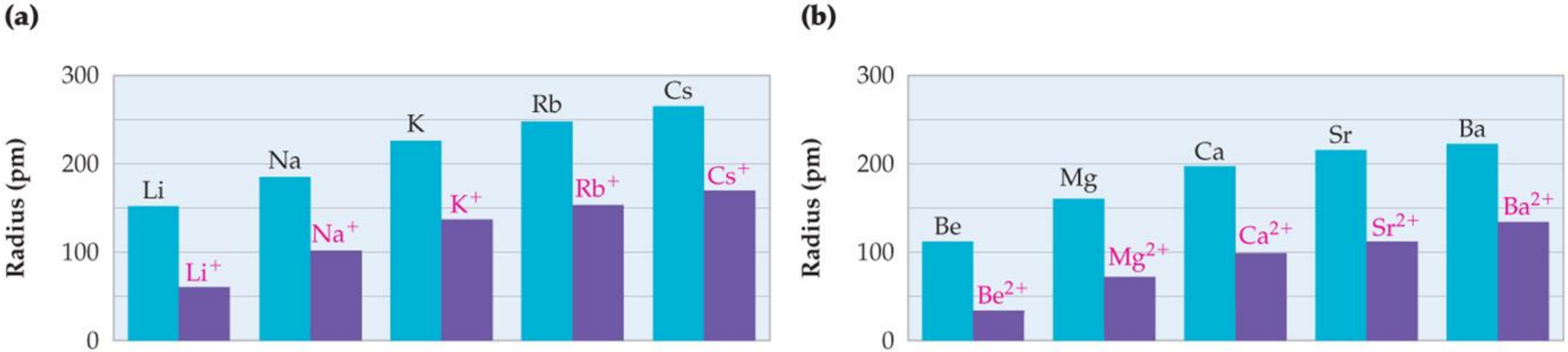
What is the electron configuration of Mn ?

What is the electron configuration of Mn^{+2} ?

What is the electron configuration of Zn ?

What is the electron configuration of Zn^{+2} ?

Ionic Radii



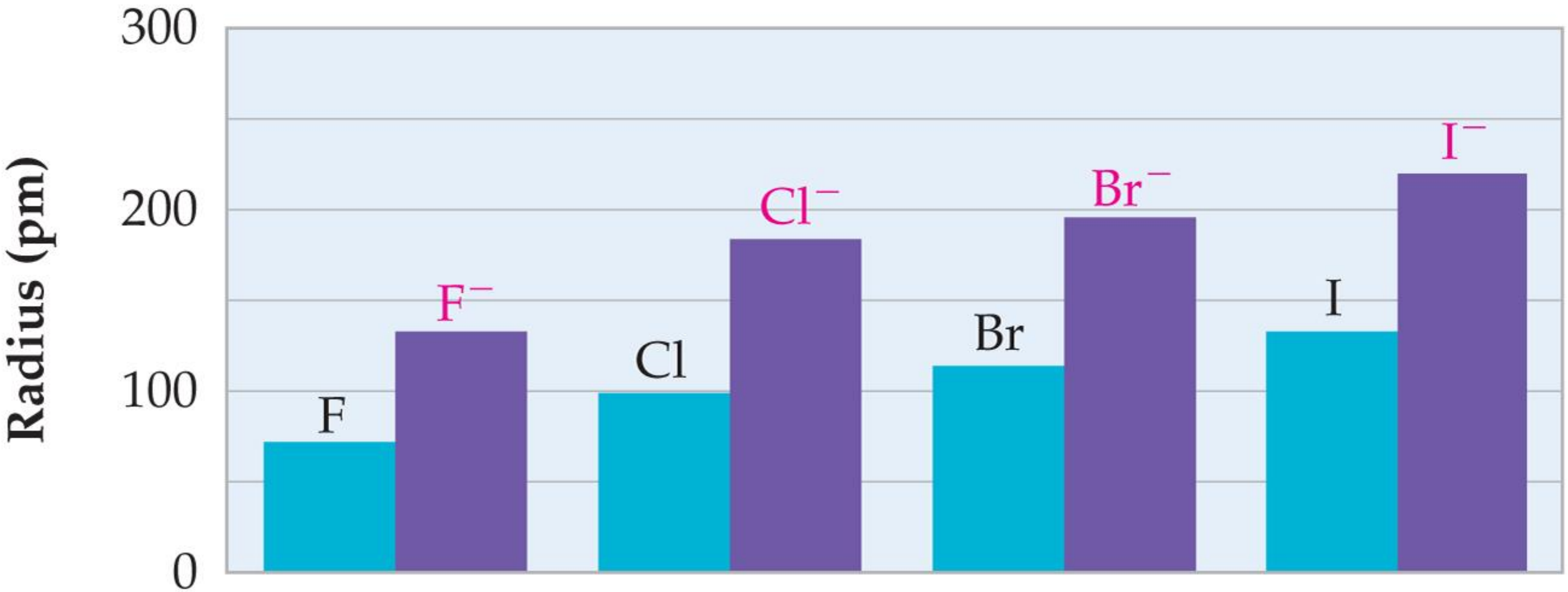
Cations are smaller than the corresponding neutral atoms, both because the principal quantum number of the valence-shell electrons is smaller for the cations than it is for the neutral atoms and because Z_{eff} is larger.

cations are smaller than neutral atoms (keep same nucleus but fewer electrons)

because fewer electrons to lower n (principal quantum number) & larger Z_{eff}

End 10/28 Monday F section

Ionic Radii



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

Anions – keep same nucleus but larger number of electrons

HW: Ionic Radii

Cations (+ charged) are smaller than neutral atoms
Anions (- charged) are larger than neutral atoms

Which is larger ?

Ca vs Ca^{+2}

S vs S^{-2}

Fe vs Fe^{+2} vs Fe^{+3}

HW: Ionic Radii

Cations (+ charged) are smaller than neutral atoms

Anions (- charged) are larger than neutral atoms

Which is larger ?

Ca vs Ca^{+2} — Ca

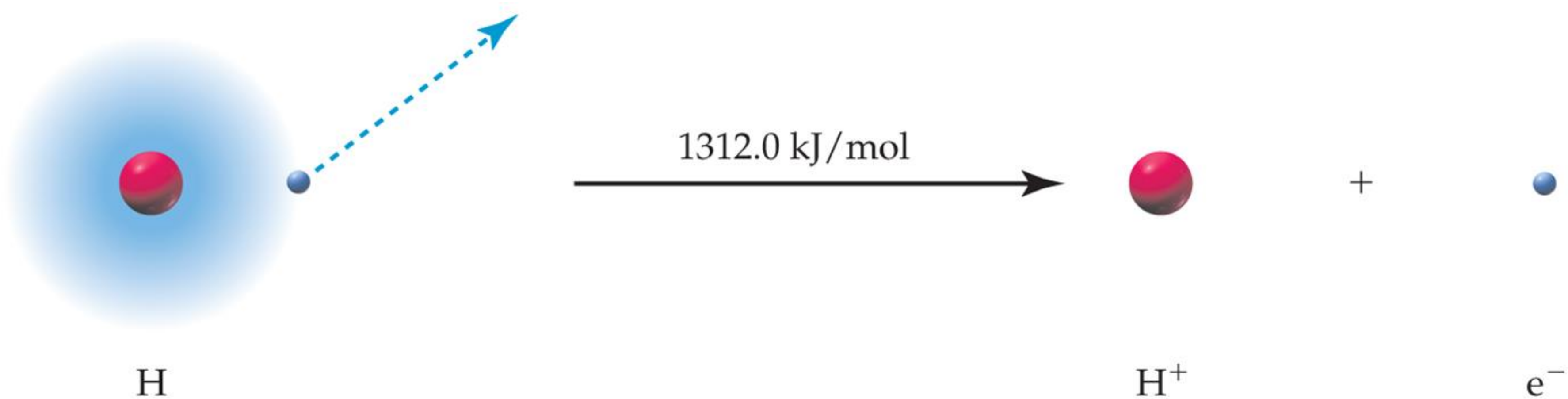
S vs S^{-2} — S^{-2}

Fe vs Fe^{+2} vs Fe^{+3} — $Fe > Fe^{+2} > Fe^{+3}$

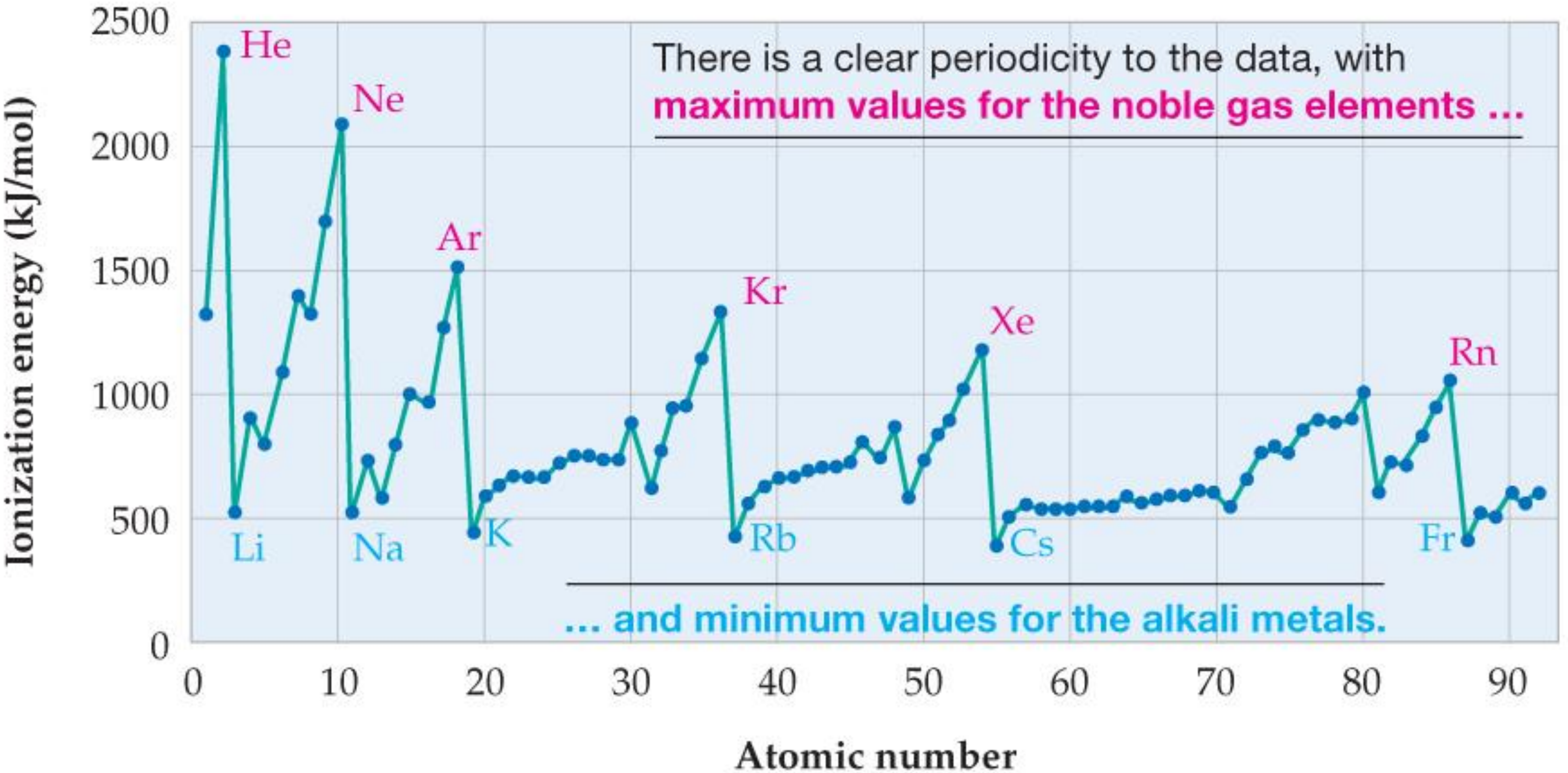
Ionization Energy

element → element$^+$ + electron

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 – largest (E_i) is noble gases



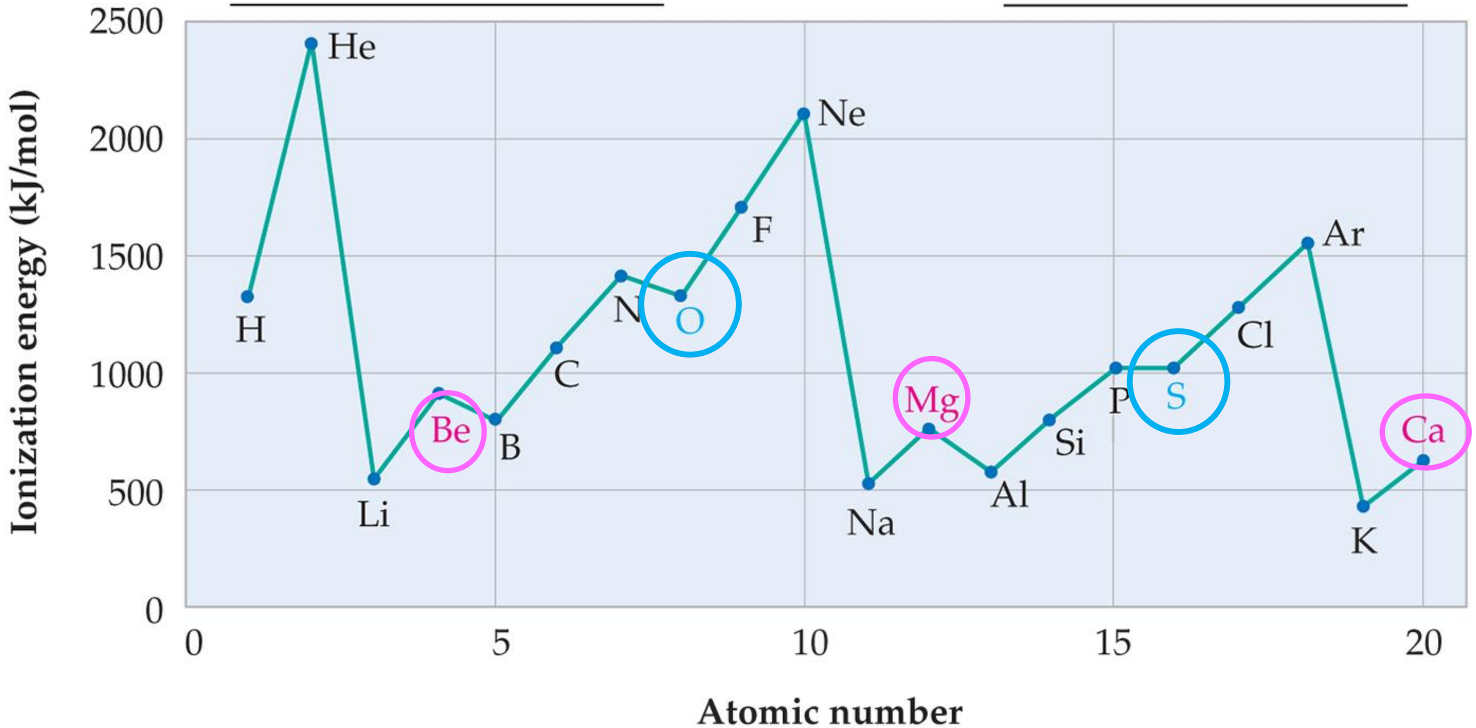
Ionization Energy

bc of stability of
½ p subshell

bc of stability of
full s subshell

The **group 2A elements (Be, Mg, Ca)** have slightly larger E_i values than might be expected.

The **group 6A elements (O, S)** have slightly smaller E_i values than might be expected.



Boron has a lower E_i due to a smaller Z_{eff}
(shielding by the 2*s* electrons or remove only e in p subshell).

Higher Ionization Energies (zig zag line - big jump in ionization energy if have to remove core electrons – to lower shell)

End 10/28 M G section

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

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

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

First ionization Ei1

second ionization Ei2

Third ionization Ei3

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.

HW: Higher Ionization Energies (zig zag line - big jump in ionization energy if have to remove core electrons – to lower shell)

M + energy → M^{+} + e^{-} First ionization Ei1

M^{+} + energy → M^{2+} + e^{-} second ionization Ei2

M^{2+} + energy → M^{3+} + e^{-} Third ionization Ei3

For the following element show the step for first ionization energy, second ionization energy and third ionization energy in equation & electron configuration orbital diagram

Reaction e configuration orbital diagram

Mg

Mg^{+1}

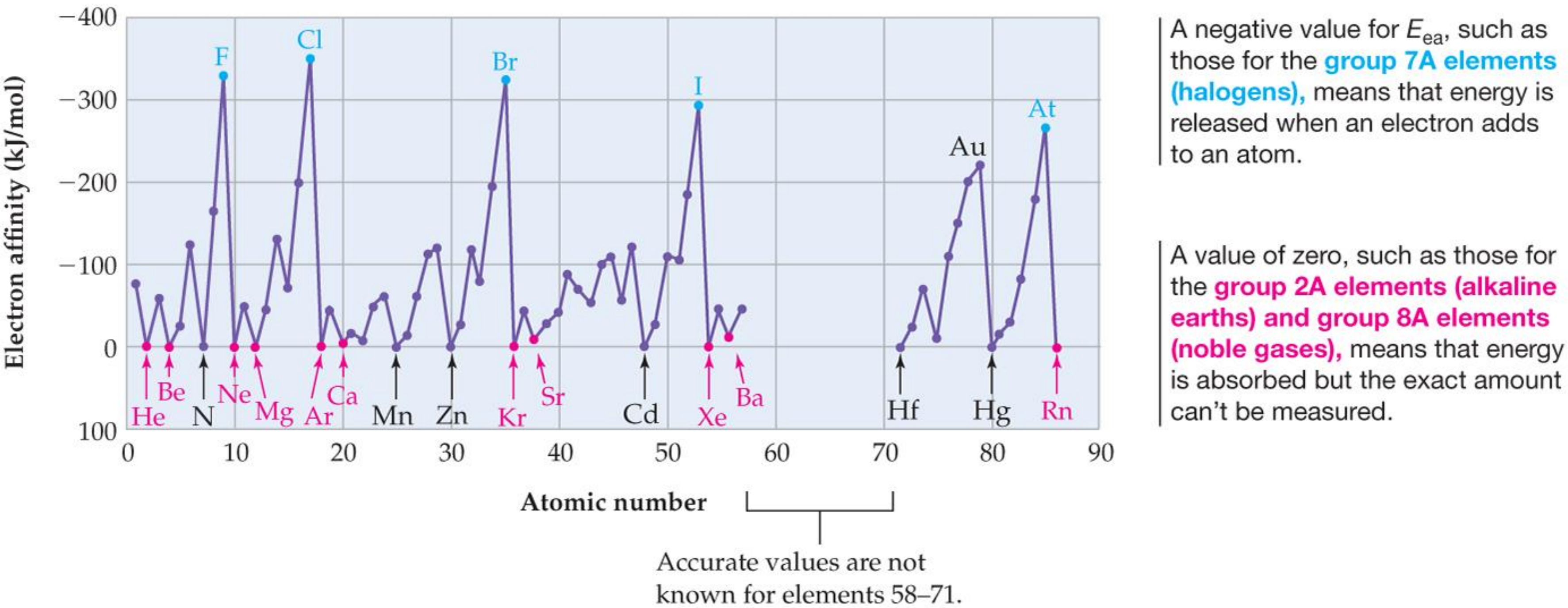
Mg^{+2}

Electron Affinity

End 10/30 F section

element + electron → element $^-$

Electron Affinity (*E*$_{ea}$): The energy change that occurs when an electron is added to an isolated atom in the gaseous state (usually negative energy, energy released)



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. (extra stability if get 8 electron outer shell)

Octet Rule – stable with 8 electrons outer shell

10/30/19 end G section

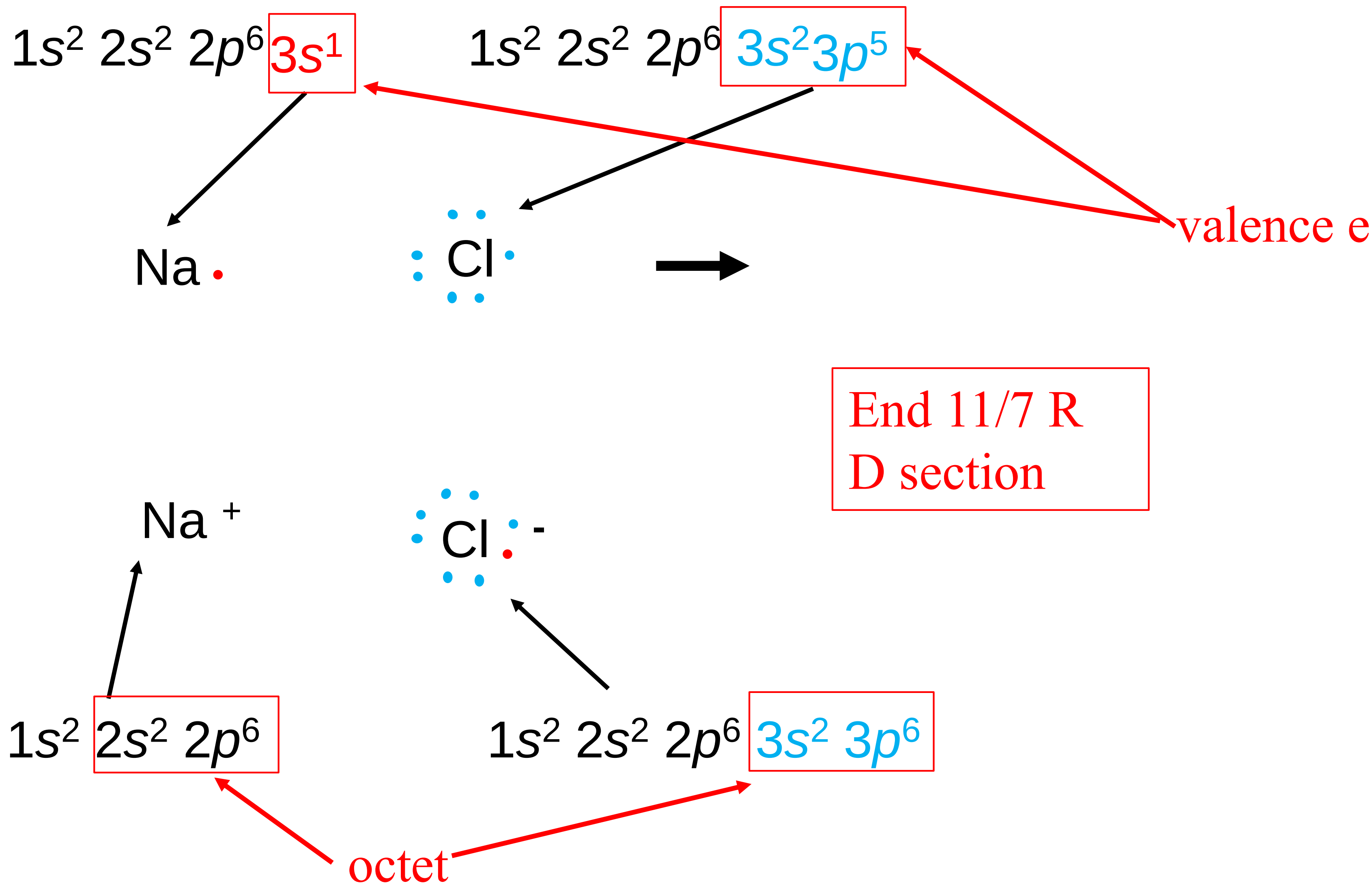
Metals tend to have **low** E_i (have to add little energy to lose e)
and **low** E_{ea}.(get out little energy when gain e)

They tend to *lose* one or more electrons.

Nonmetals tend to have **high** E_i (have to add lot of energy to lose e)
and **high** E_{ea}. (get out lots energy when gain e)

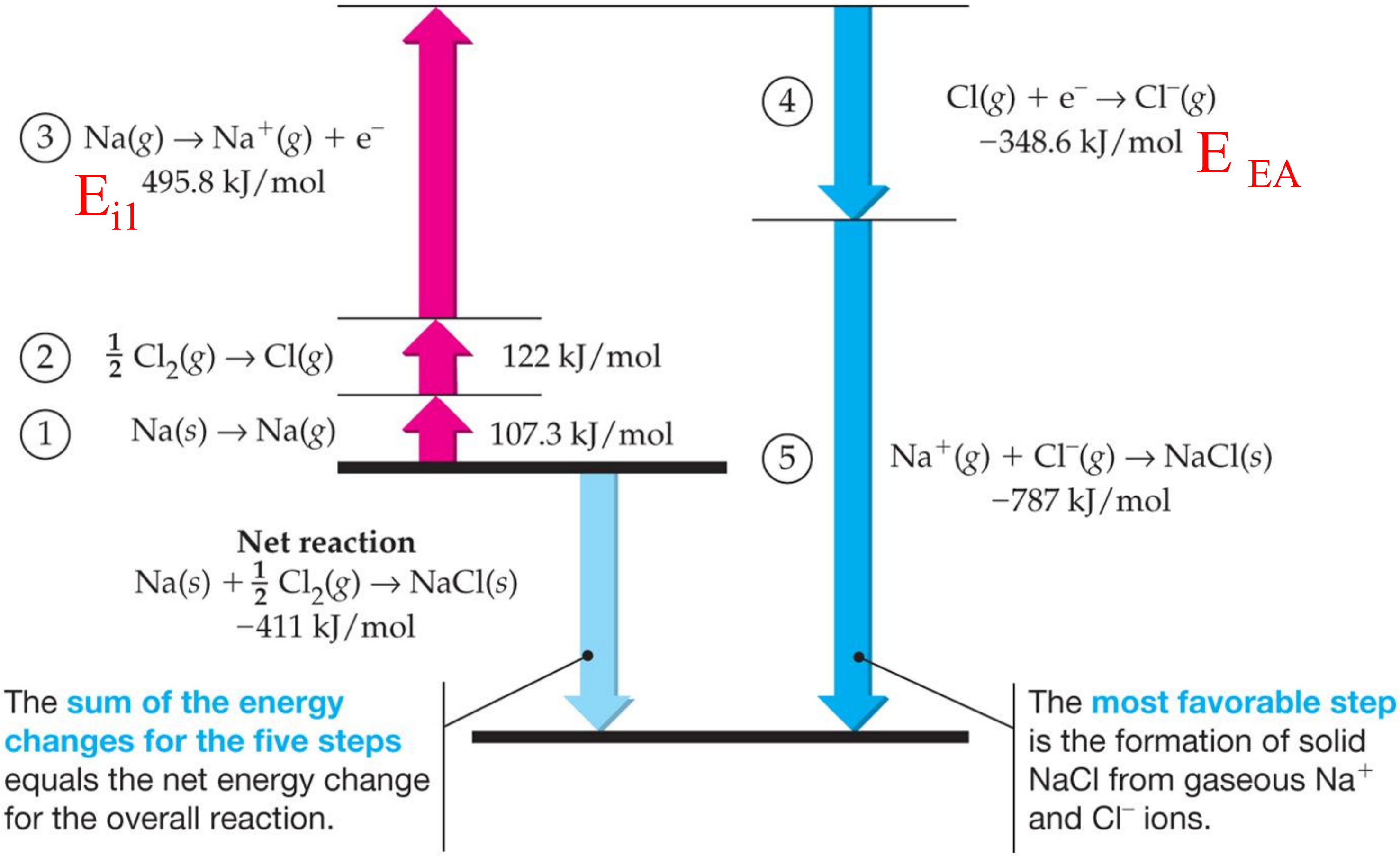
They tend to *gain* one or more electrons.

Ionic Bonds and the Formation of Ionic Solids



Ionic Bonds and the Formation of Ionic Solids (steps to form/break up ionic compound)

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

sublimination – solid to gas

Step 2: $\frac{1}{2}Cl_2(g) \longrightarrow Cl(g)$ +122 kJ/mol

bond dissociation

Step 3: $Na(g) \longrightarrow Na^+(g) + e^-$ +495.8 kJ/mol

ionization energy

Step 4: $Cl(g) + e^- \longrightarrow Cl^-(g)$ –348.6 kJ/mol

electron affinity

Step 5: $Na^+(g) + Cl^-(g) \longrightarrow NaCl(s)$ –787 kJ/mol

reverse of lattice energy

$Na(s) + Cl_2(g) \longrightarrow NaCl(s)$ –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 + (to form ionic solid – sign)

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

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 + (to form ionic solid – sign)

$$U = k * \frac{z_1 z_2}{d}$$

[k = constant, $z_1 z_2$ are charges, d=distance between ions – related to atomic radius]

U larger for smaller atomic radius

U larger for larger charges on ions

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 + sign (to form ionic solid – sign)

U larger for smaller atomic radius

ex: U(LiF) > U(NaF) > U(KF)

(Li small size < Na < K big size)

U larger for larger ion charge

ex: U(AlI_3) > U(MgI_2) > U(NaI)

charges (Al +3 > Mg +2 > Na +1)

Lattice Energies in Ionic Solids

Lattice Energy (*U*): break up an ionic solid into individual gaseous ions + sign (to form ionic solid – sign)

end 11/1 F, G section

U larger for smaller atomic radius
U larger for larger ion charge

Which has larger U ? $AlCl_3$

$AlCl_3$ vs $MgCl_2$?

+3	+2
Al smaller	Mg bigger