Chapter 7

Electron Configuration and the Periodic Table

7.1 Development of the Periodic Table

- 1864 - John Newlands - Law of Octaves- every 8th element had similar properties when arranged by atomic masses (not true past Ca)
- 1869 - Dmitri Mendeleev & Lothar Meyer - independently proposed idea of periodicity (recurrence of properties)

- Mendeleev
  - Grouped elements (66) according to properties
  - Predicted properties for elements not yet discovered
  - Though a good model, Mendeleev could not explain inconsistencies, for instance, all elements were not in order according to atomic mass

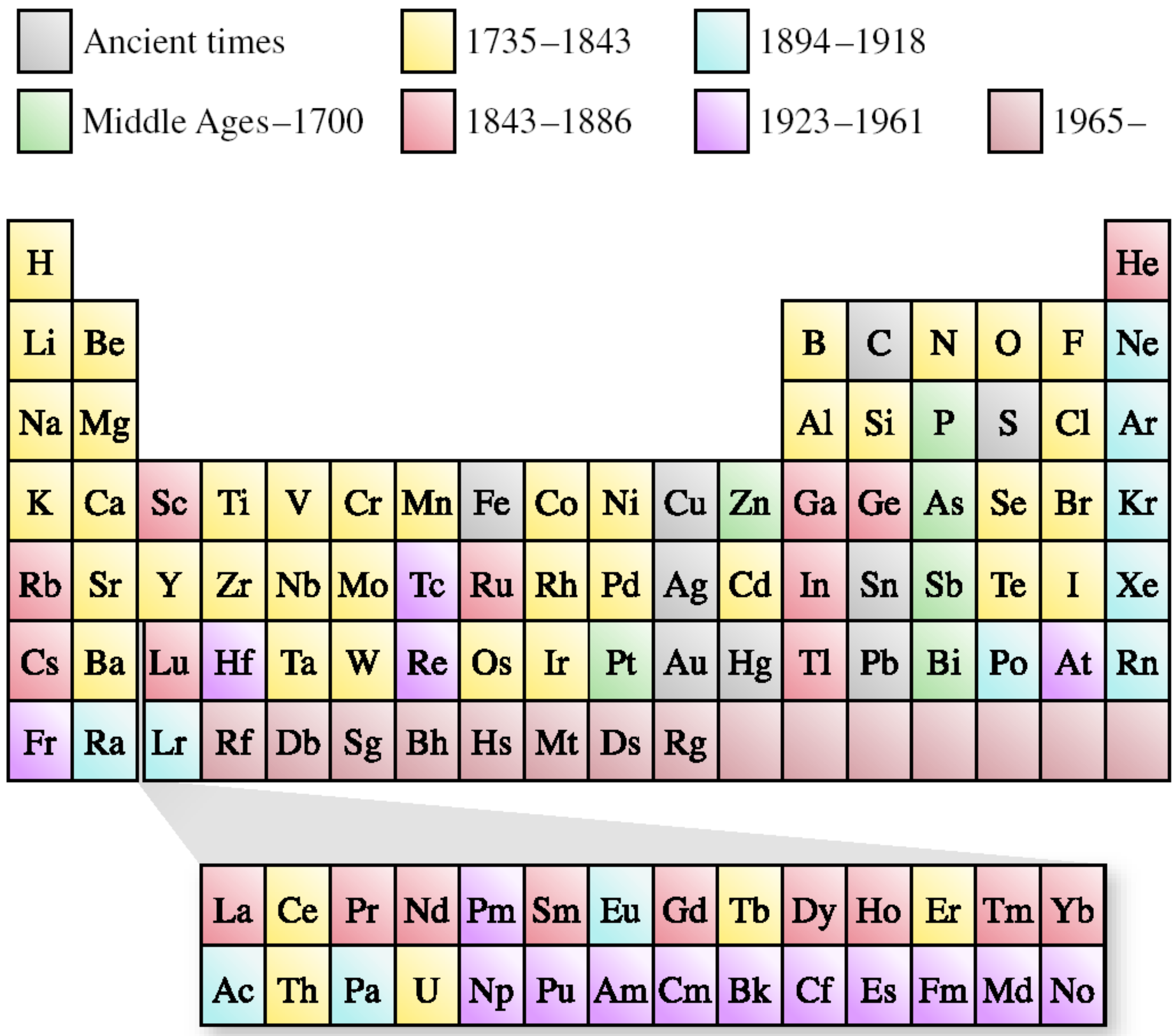
Mendeleev Early Periodic Table

TABELLE II

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	—	— — — —

- 1913 - Henry Moseley explained the discrepancy
  - Discovered correlation between number of protons (atomic number) and frequency of X rays generated
  - Today, elements are arranged in order of increasing atomic number

Periodic Table by Dates of Discovery



The Modern Periodic Table

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	8B 9	8B 10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	
1	1 **H** $1s^1$																	2 **He** $1s^2$	1
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$	2
3	11 **Na** $3s^1$	12 **Mg** $3s^2$											13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$	3
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $3d^{10}4s^2$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$	4
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$	5
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	71 **Lu** $6s^25d^14f^{14}$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^15d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$	6
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	103 **Lr** $7s^25f^{14}6d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 **Ds** $7s^26d^8$	111 **Rg** $7s^26d^9$	112 **—** $7s^26d^{10}$	113 **—** $7s^27p^1$	114 **—** $7s^27p^2$	115 **—** $7s^27p^3$	116 **—** $7s^27p^4$	(117)	118 **—** $7s^27p^6$	7

57 **La** $6s^25d^1$	58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$
89 **Ac** $7s^26d^1$	90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$

7.2 The Modern Periodic Table

- Classification of Elements
  - Main group elements - “representative elements” Group 1A-7A
  - Noble gases - Group 8A all have ns^2np^6 configuration(exception-He)
  - Transition elements - 1B, 3B - 8B “*d*-block”
  - Lanthanides/actinides - “*f*-block”

Periodic Table Colored Coded By Main Classifications

	1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18	
1	H																	He	1
2	Li	Be											B	C	N	O	F	Ne	2
3	Na	Mg											Al	Si	P	S	Cl	Ar	3
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	4
5	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	5
6	Cs	Ba	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	6
7	Fr	Ra	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg								7

6	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	6
7	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	7

TABLE 7.1 Electron Configurations of Group 1A and Group 2A Elements

Group 1A		Group 2A	
Li	$[He]2s^1$	Be	$[He]2s^2$
Na	$[Ne]3s^1$	Mg	$[Ne]3s^2$
K	$[Ar]4s^1$	Ca	$[Ar]4s^2$
Rb	$[Kr]5s^1$	Sr	$[Kr]5s^2$
Cs	$[Xe]6s^1$	Ba	$[Xe]6s^2$
Fr	$[Rn]7s^1$	Ra	$[Rn]7s^2$

- Predicting properties
  - Valence electrons are the outermost electrons and are involved in bonding
  - Similarity of valence electron configurations help predict chemical properties
  - Group 1A, 2A and 8A all have similar properties to other members of their respective group

- Groups 3A - 7A show considerable variation among properties from metallic to nonmetallic

- Transition metals do not always exhibit regular patterns in their electron configurations but have some similarities as a whole such as colored compounds and multiple oxidation states.

- Representing Free Elements in Chemical Equations
  - **Metals** are always represented by their empirical formulas (same as symbol for element)
  - **Nonmetals** may be written as empirical formula (C) or as polyatomic molecules (H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, and P_4).
  - Sulfur usually S instead of S_8

- **Noble Gases** all exist as isolated atoms, so use symbols (Xe, He, etc.)

- **Metalloids** are represented with empirical formulas (B, Si, Ge, etc.)

7.3 Effective Nuclear Charge

- ***Z* (*nuclear charge*) =** the number of protons in the nucleus of an atom

- ***Z*$_{\text{eff}}$ (*effective nuclear charge*)** = the magnitude of positive charge “experienced” by an electron in the atom

- ***Z*$_{\text{eff}}$** increases from left to right across a period; changes very little down a column

- **Shielding** occurs when an electron in a many-electron atom is partially shielded from the positive charge of the nucleus by other electrons in the atom.

- However, **core electrons (inner electrons)** shield the most and are constant across a period.

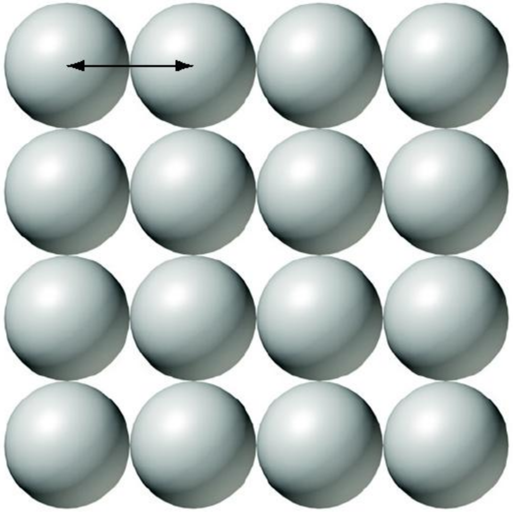
- $Z_{eff} = Z - \sigma$
  - **σ** represents the shielding constant (greater than 0 but less than *Z*) ***equals the number of inner electrons***
  - Example:

	Li	**Be**	**B**	**C**	**N**
Z	3	4	5	6	7
Z_{eff}	1	2	3	4	5

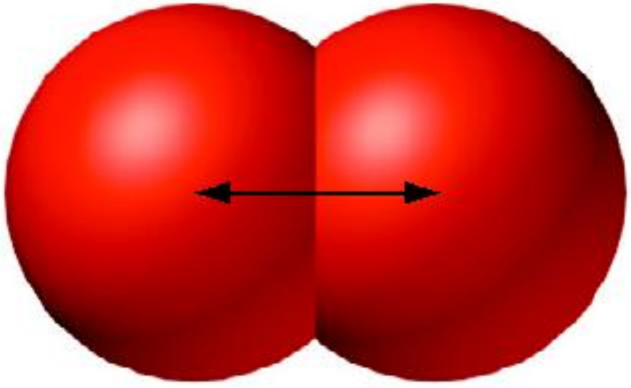
7.4 Periodic Trends in Properties of Elements

- ***Atomic radius***: distance between nucleus of an atom and its valence shell

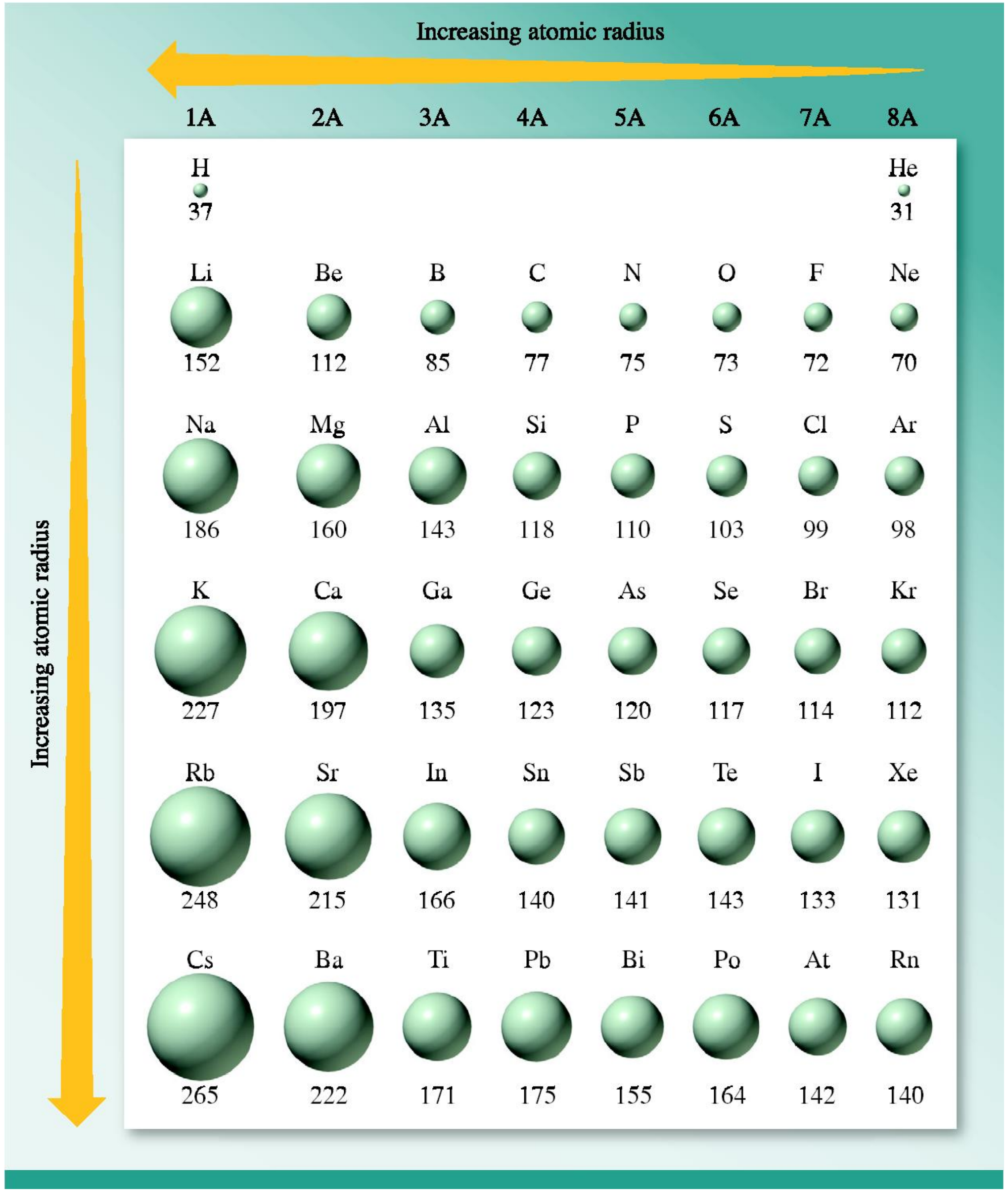
- ***Metallic radius***: half the distance between nuclei of two adjacent, identical metal atoms



- **Covalent radius:** half the distance between adjacent, identical nuclei in a molecule



Atomic Radii (pm) of the Elements



Explain

- What do you notice about the atomic radius across a period? Why? (hint: Z_{eff})

- What do you notice about the atomic radius down a column? Why? (hint: n)

- What do you notice about the atomic radius across a period? Why? (hint: Z_{eff})

Atomic radius decreases from left to right across a period due to increasing Z_{eff}.(*as $Z_{eff.}$ Increases the attraction forces between the nucleus and V.S.E increases so the size decreases.*

- What do you notice about the atomic radius down a column? Why? (hint: *n*)

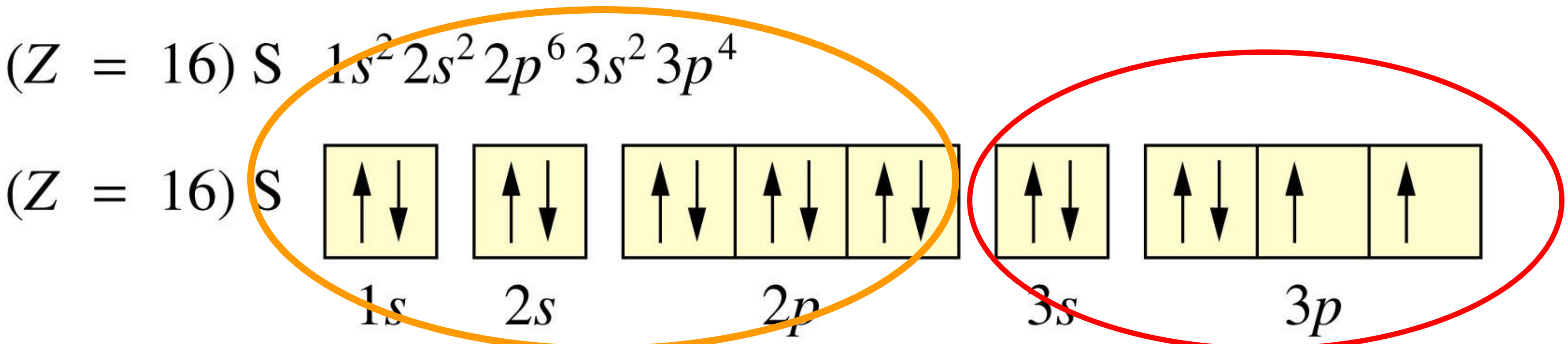
Atomic radius increases down a column of the periodic table because the distance of the electron from the nucleus increases as *n* increases.

Valence and Core Electrons

Valence electrons are those with the highest principal quantum number

Electrons in inner shells are called core electrons

Sulfur has 10 core electrons and 6 valence electrons



Noble Gases

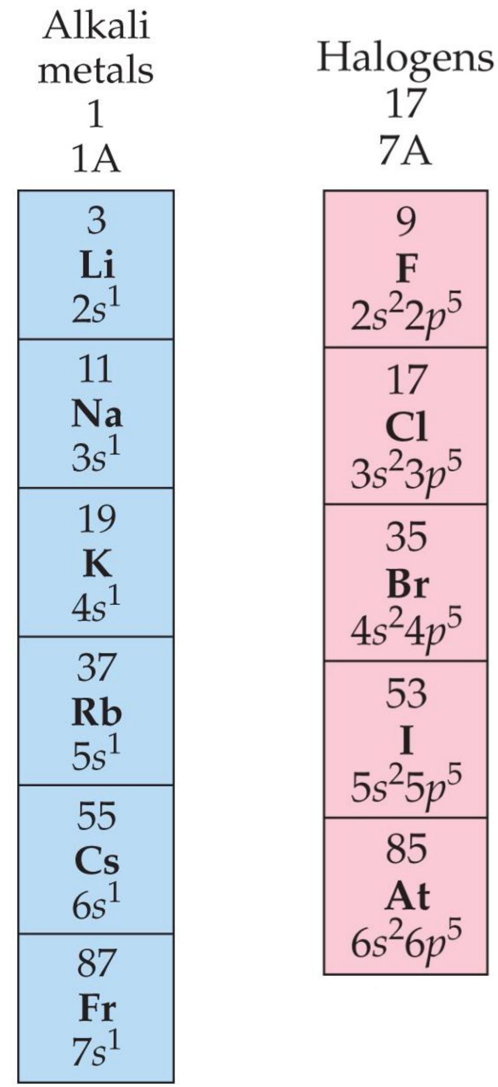
- Noble gases have **filled** energy levels
- The noble gases have **8 valence electrons**.
  - Except for **He,** which has only **2 electrons**.
- We know the noble gases are especially non-reactive.
  - He and Ne are practically inert.
- The reason the noble gases are so non-reactive is that the electron configuration of the noble gases is especially stable.

Noble gases 18 8A
2 **He** $1s^2$
10 **Ne** $2s^22p^6$
18 **Ar** $3s^23p^6$
36 **Kr** $4s^24p^6$
54 **Xe** $5s^25p^6$
86 **Rn** $6s^26p^6$

Everyone Wants to Be Like a Noble Gas!

- **The alkali metals** have one more electron than the previous noble gas.
- Forming a cation with a **1+ charge**

- **The halogens** all have one electron less than the next noble gas.
- Forming an anion with **charge -1**



Stable Electron Configuration

Atom	Atom's electron config	Ion	Ion's electron config
Na	$[Ne]3s^1$	Na^+	[Ne]
Mg	$[Ne]3s^2$	Mg^{2+}	[Ne]
Al	$[Ne]3s^2 3p^1$	Al^{3+}	[Ne]
O	$[He]2s2p^4$	O^{2-}	[Ne]
F	$[He]2s^2 2p^5$	F^-	[Ne]

- **Ionization energy (*IE*):** minimum energy needed to remove an electron from an atom in the gas phase
  - Representation:

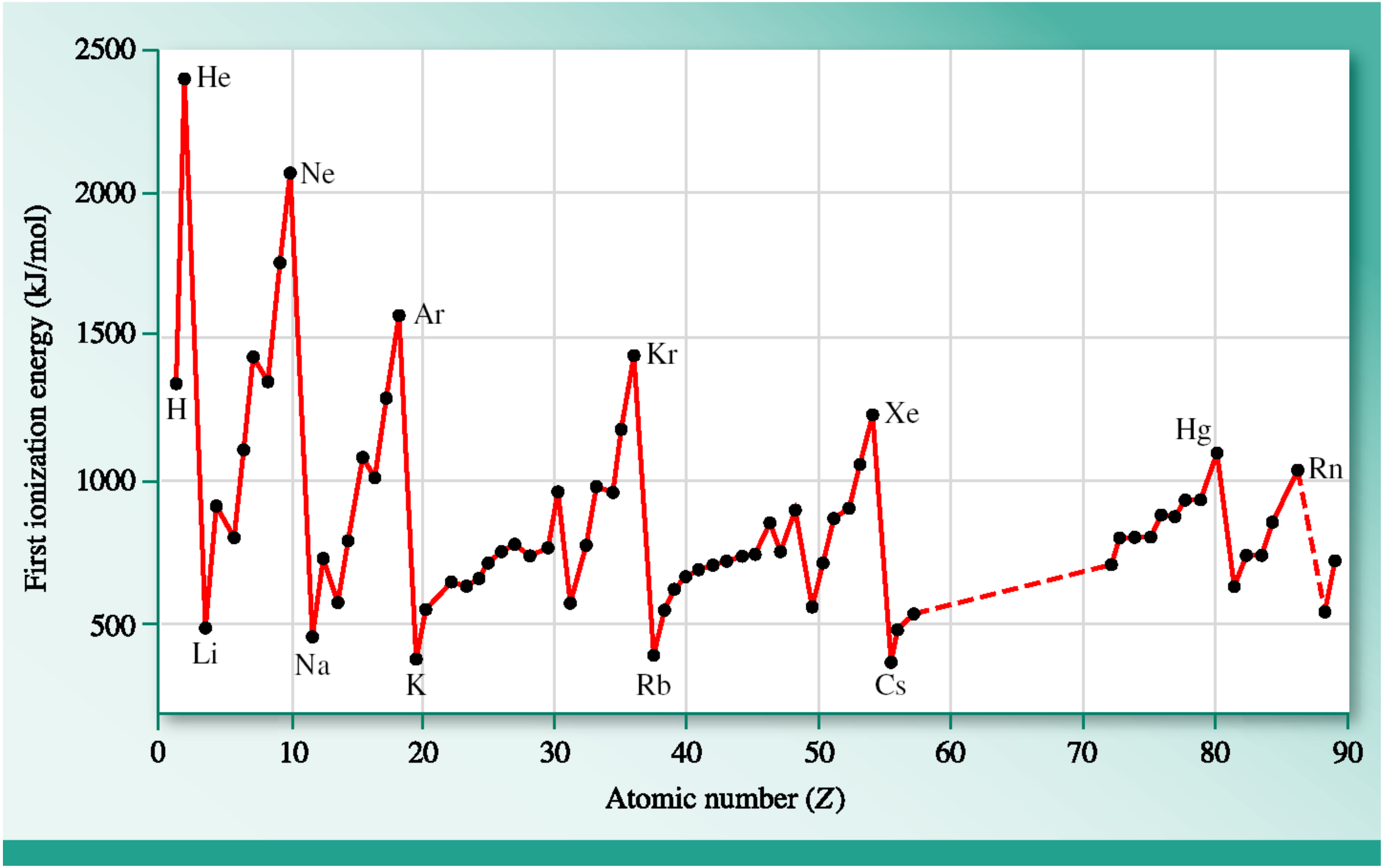
$$Na_{(g)} \rightarrow Na^+_{(g)} + e^-$$

  - *IE* for this 1st ionization = 495.8 kJ/mol
- **In general,** ionization energy increases as Z_{eff} increases
  - Exceptions occur due to the stability of specific electron configurations

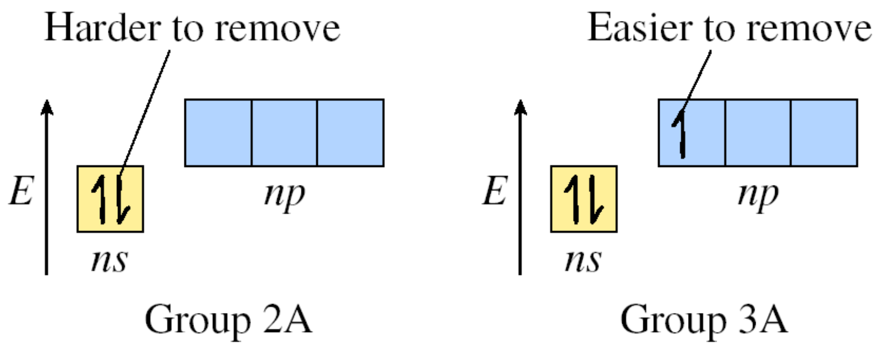
IE_1 (kJ/mol) Values for Main Group Elements

1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H 1312							**He** 2372
Li 520	**Be** 899	**B** 800	**C** 1086	**N** 1402	**O** 1314	**F** 1681	**Ne** 2080
Na 496	**Mg** 738	**Al** 577	**Si** 786	**P** 1012	**S** 999	**Cl** 1256	**Ar** 1520
K 419	**Ca** 590	**Ga** 579	**Ge** 761	**As** 947	**Se** 941	**Br** 1143	**Kr** 1351
Rb 403	**Sr** 549	**In** 558	**Sn** 708	**Sb** 834	**Te** 869	**I** 1009	**Xe** 1170
Cs 376	**Ba** 503	**Tl** 589	**Pb** 715	**Bi** 703	**Po** 813	**At** (926)	**Rn** 1037

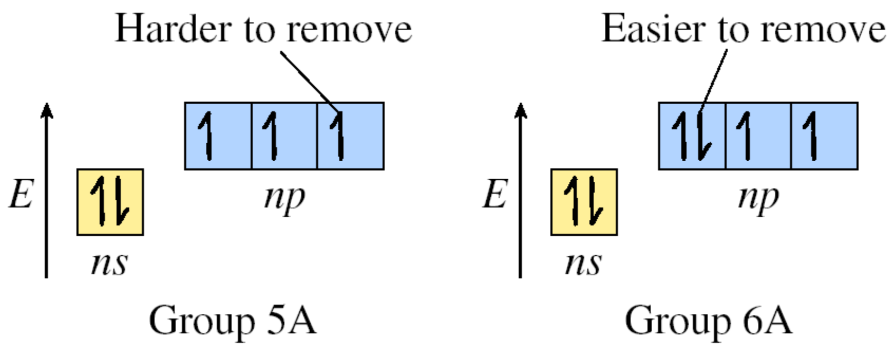
Periodic Trends in IE_1



- What do you notice about the 1st *IE* between 2A and 3A? Why? (hint: draw the electron configuration)



- What do you notice about the 1st *IE* between 5A and 6A? Why? (hint: draw the electron configuration)



- **Multiple Ionizations**: it takes more energy to remove the 2nd, 3rd, 4th, etc. electron and much more energy to remove core electrons
- Why?
  - Core electrons are closer to nucleus
  - Core electrons experience greater Z_{eff}

Ionization Energy

- **for Mg**

  $Mg = 1s^2\,2s^2\,2p^6\,\mathbf{3s^2}$

  – $I_1 = 735$ kJ/mole

  $Mg^+ = 1s^2\,2s^2\,2p^6\,\mathbf{3s^1}$

  – $I_2 = 1445$ kJ/mole

  $Mg^{+2} = 1s^2\,2s^2\,2p^6\,\mathbf{3s^0}$

  – $I_3 = 7730$ kJ/mole
- The effective nuclear charge increases as electrons are removed
- It takes much more energy to remove a core electron than a valence electron

TABLE 7.3 Ionization Energies (in kJ/mol) for Elements 3 through 11*

	Z	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
Li	3	520	7,298	11,815							
Be	4	899	1,757	14,848	21,007	21,007					
B	5	800	2,427	3,660	25,026	32,827					
C	6	1,086	2,353	4,621	6,223	37,831	47,277				
N	7	1,402	2,856	4,578	7,475	9,445	53,267	64,360			
O	8	1,314	3,388	5,301	7,469	10,990	13,327	71,330	84,078		
F	9	1,681	3,374	6,050	8,408	11,023	15,164	17,868	92,038	106,434	
Ne	10	2,080	3,952	6,122	9,371	12,177	15,238	19,999	23,069	115,380	131,432
Na	11	496	4,562	6,910	9,543	13,354	16,613	20,117	25,496	28,932	141,362

*Cells shaded with blue represent the removal of core electrons.

- **Electron Affinity (*EA*):** energy released when an atom in the gas phase accepts an electron
  - Representation:

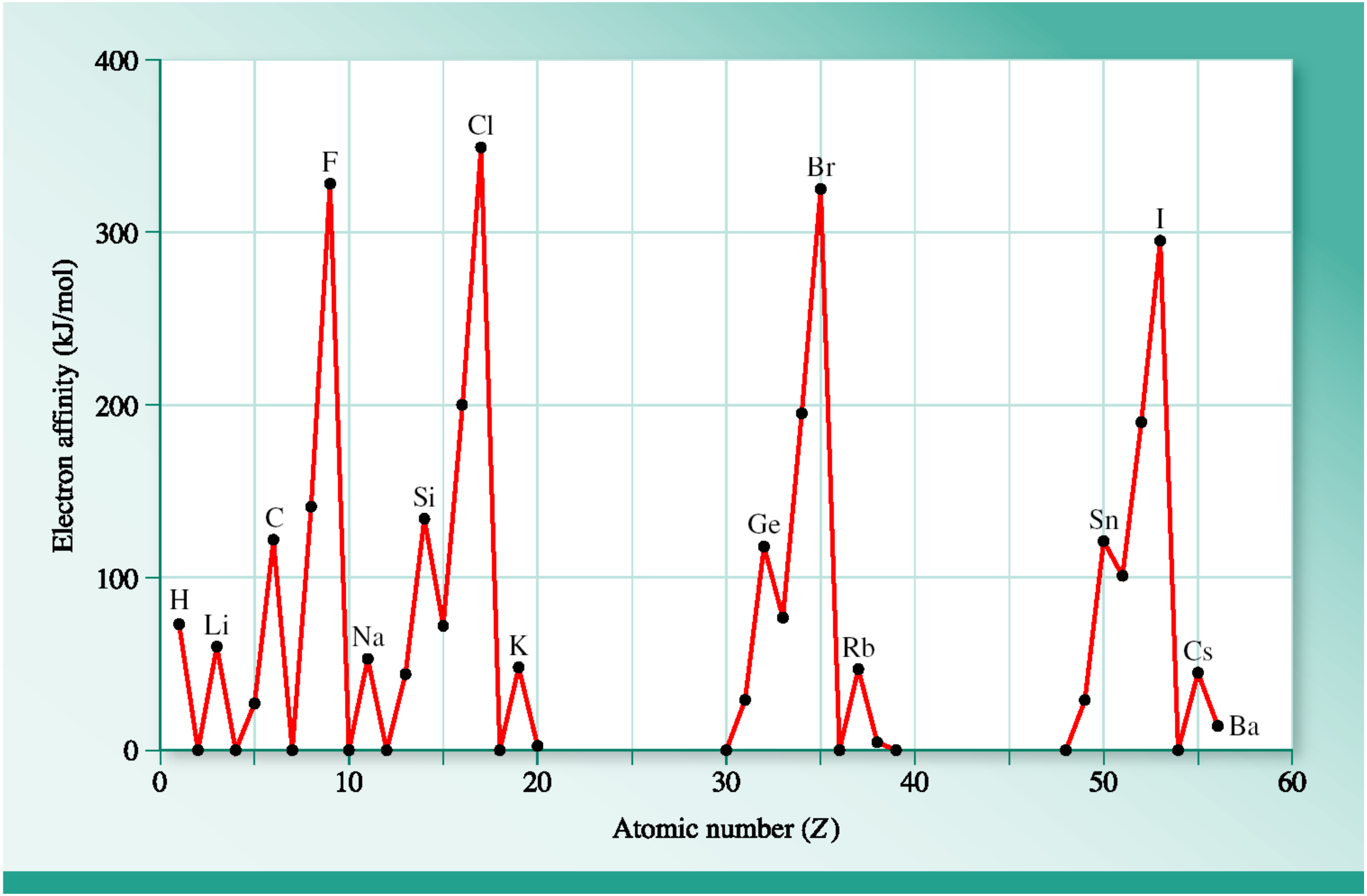
$$Cl_{(g)} + e^{-} \rightarrow Cl^{-}_{(g)}$$

  - *EA* for this equation 349.0 kJ/mol energy released (ΔH = negative)

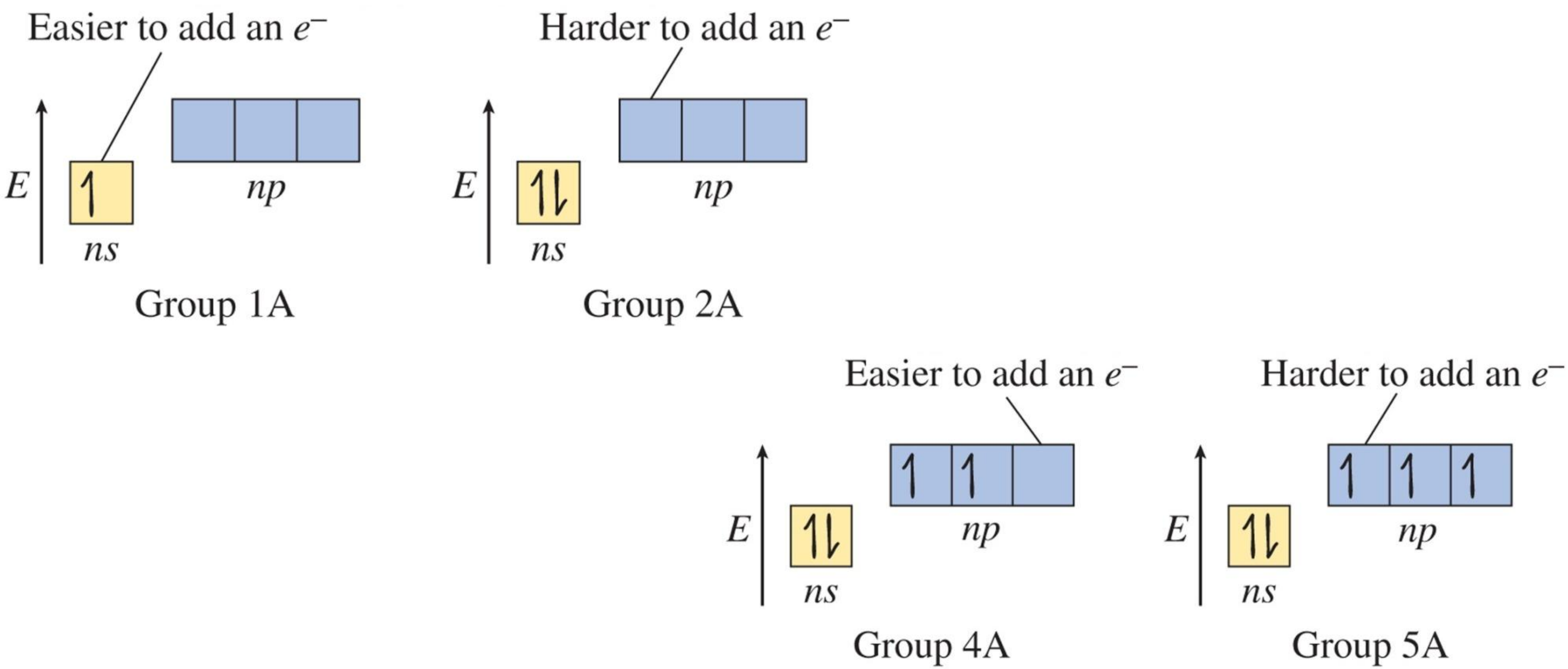
EA (kJ/mol) Values for Main Group Elements

1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H +72.8							**He** (0.0)
Li +59.6	**Be** ≤0	**B** +26.7	**C** +122	**N** −7	**O** +141	**F** +328	**Ne** (−29)
Na +52.9	**Mg** ≤0	**Al** +42.5	**Si** +134	**P** +72.0	**S** +200	**Cl** +349	**Ar** (−35)
K +48.4	**Ca** +2.37	**Ga** +28.9	**Ge** +119	**As** +78.2	**Se** +195	**Br** +325	**Kr** (−39)
Rb +46.9	**Sr** +5.03	**In** +28.9	**Sn** +107	**Sb** +103	**Te** +190	**I** +295	**Xe** (−41)
Cs +45.5	**Ba** +13.95	**Tl** +19.3	**Pb** +35.1	**Bi** +91.3	**Po** +183	**At** +270	**Rn** (−41)

Periodic Trends in *EA*



- Periodic Interruptions in *EA*
  - Explained in much the same way as *IE* except not the same elements!



- Metallic Character
  - Metals
    - Shiny, lustrous, malleable
    - Good conductors
    - Low *IE* (form cations)
    - Form ionic compounds with chlorine
    - Form basic, ionic compounds with oxygen
    - Metallic character increases top to bottom in group and decreases left to right across a period

- Nonmetals
  - Vary in color, not shiny
  - Brittle
  - Poor conductors
  - Form acidic, molecular compounds with oxygen
  - High *EA* (form anions)
- Metalloids
  - Properties between the metals and nonmetals

7.5 Electron Configuration of Ions

- Follow Hund’s rule and Pauli exclusion principle as for atoms
- Writing electron configurations helps explain charges memorized earlier

- Ions of main group elements
  - Noble gases (8A) almost completely unreactive due to electron configuration
    - ns^2np^6 (except He $1s^2$)
  - Main group elements tend to gain or lose electrons to become **isoelectronic** (same valence electron configuration as nearest noble gas)

Na: $1s^2 2s^2 2p^6 3s^1 \rightarrow$ Na^+ $1s^2 2s^2 2p^6$

Na: $[Ne]3s^1 \rightarrow$ Na^+ $[Ne]$

(Na^+ 10 electrons - isoelectronic with Ne)

Cl: $1s^2 2s^2 2p^6 3s^2 3p^5 \rightarrow$ Cl^- $1s^2 2s^2 2p^6 3s^2 3p^6$

Cl: $[Ne]3s^2 3p^5 \rightarrow$ Cl^- $[Ar]$

(Cl^- 18 electrons - isoelectronic with Ar)

- Ions of *d*-Block Elements
  - Recall that the 4*s* orbital fills before the 3*d* orbital in the first row of transition metals
  - Electrons are always lost from the highest "*n*" value (then from *d*)

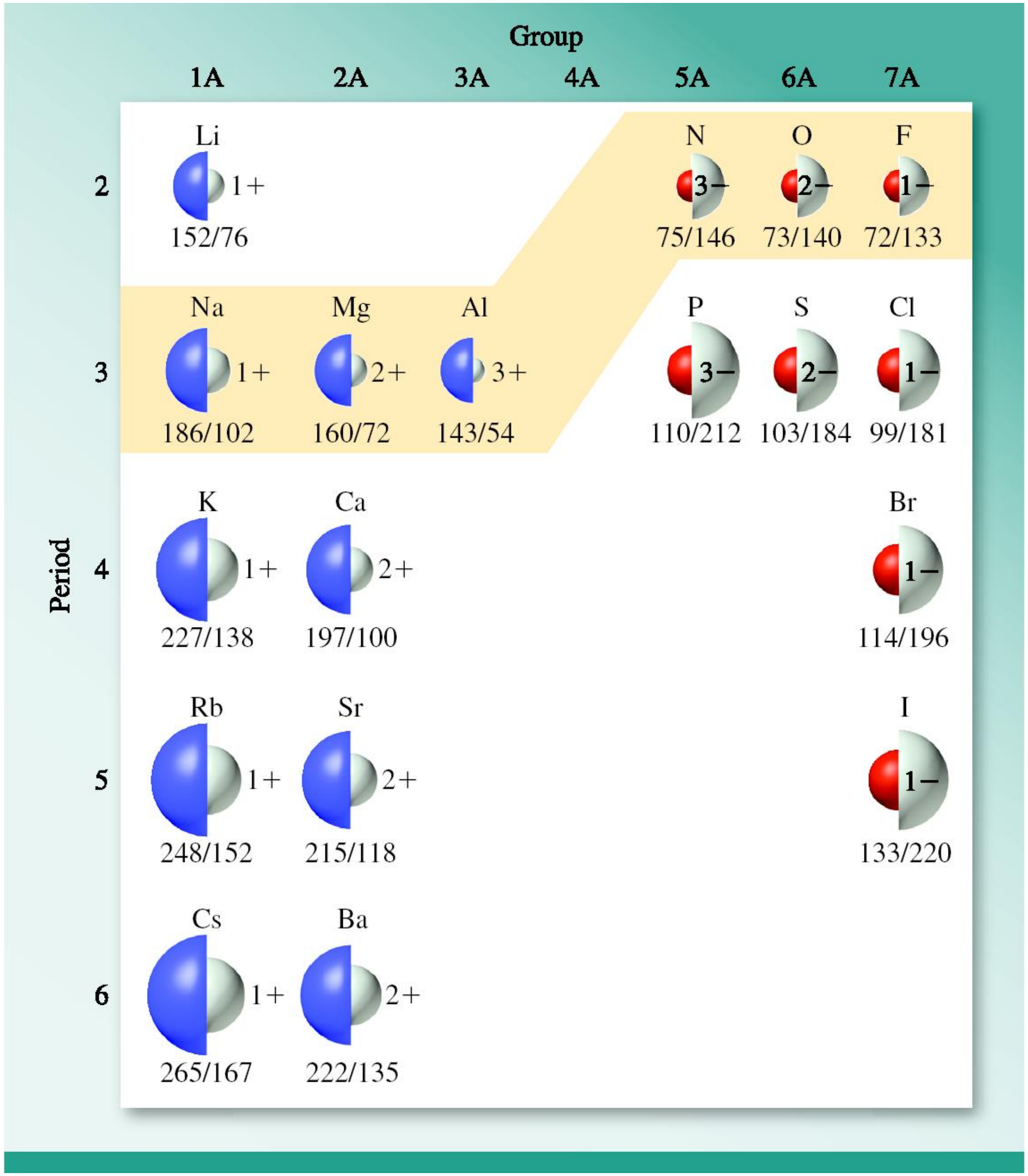
Fe: $[Ar]4s^2 3d^6 \rightarrow Fe^{2+}: [Ar]3d^6$

Fe: $[Ar]4s^2 3d^6 \rightarrow Fe^{3+}: [Ar]3d^5$

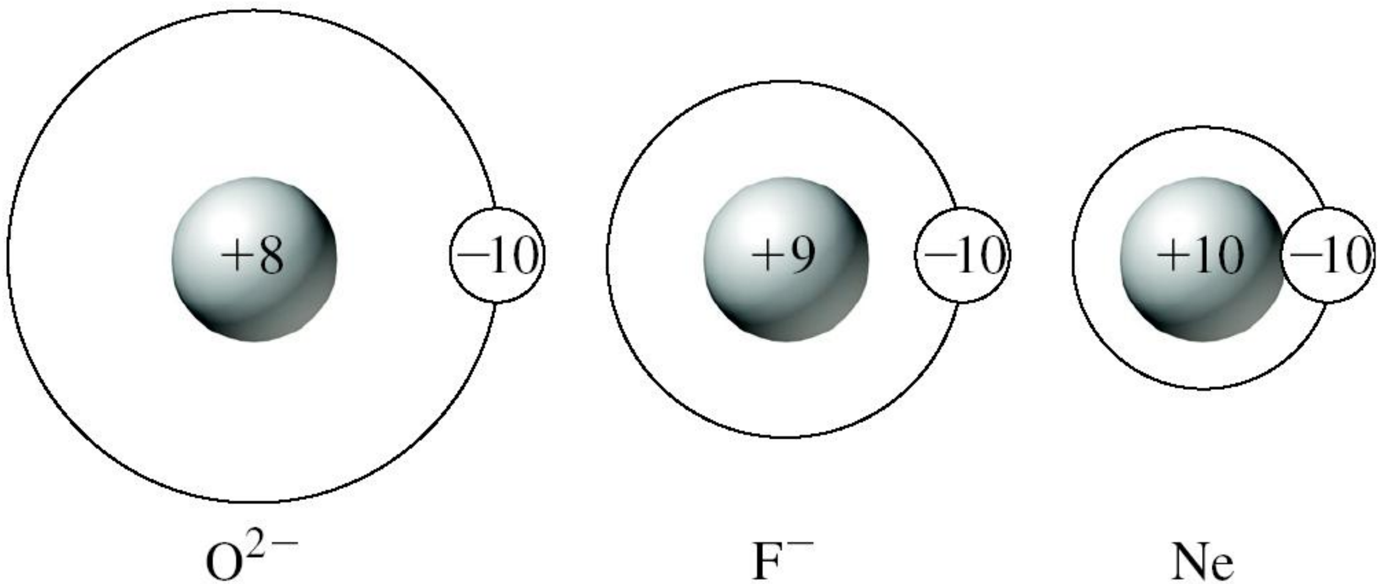
7.6 Ionic Radius

- When an atom gains or loses electrons, the radius changes
- Cations are always smaller than their parent atoms (often losing an energy level)
- Anions are always larger than their parent atoms (increased e^- repulsions)

Comparison of Atomic and Ionic Radii

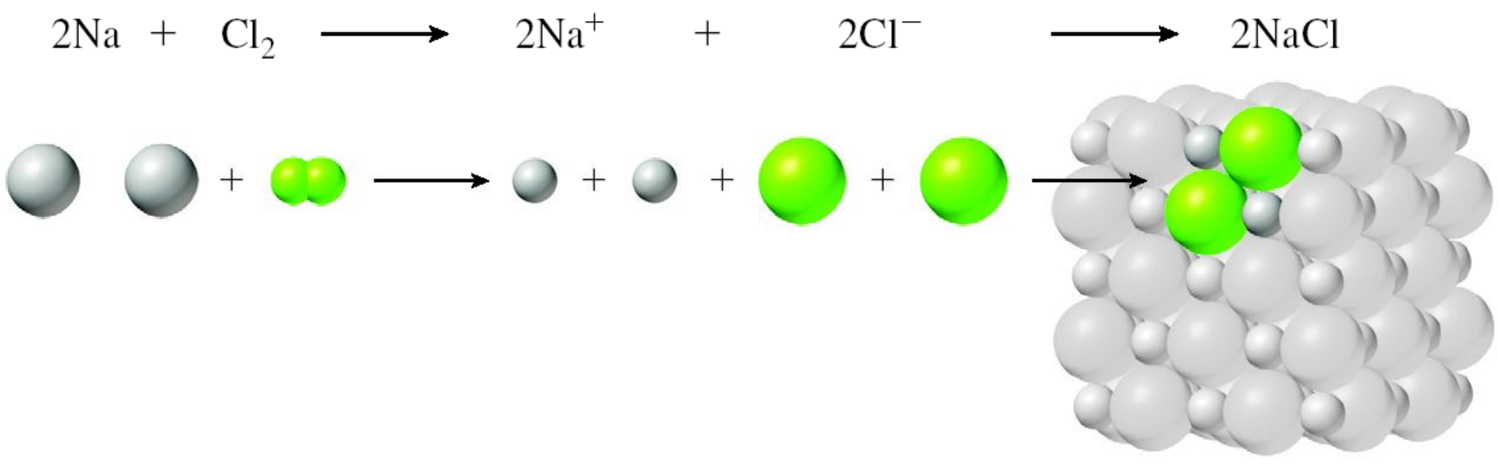


- Isoelectronic Series
  - Two or more species having the same electron configuration but different nuclear charges
  - Size varies significantly



7.7 Periodic Trends in Chemical Properties of Main Group Elements

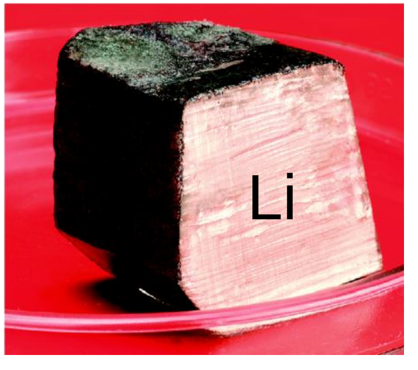
- *IE* and *EA* enable us to understand types of reactions that elements undergo and the types of compounds formed



- **General Trends in Chemical Properties**
  - Elements in same group have same valence electron configuration; similar properties
  - Same group comparison most valid if elements have same metallic or nonmetallic character
  - Group 1A and 2A; Group 7A and 8A
  - Careful with Group 3A - 6A

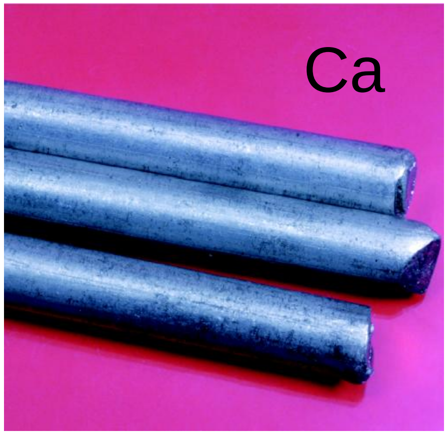
– Hydrogen ($1s^1$)
  - Group by itself
  - Forms +1 (H^+)
    – Most important compound is water
  - Forms –1 (H^-), the hydride ion, with metals
    – Hydrides react with water to produce hydrogen gas and a base
    – $CaH_2(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

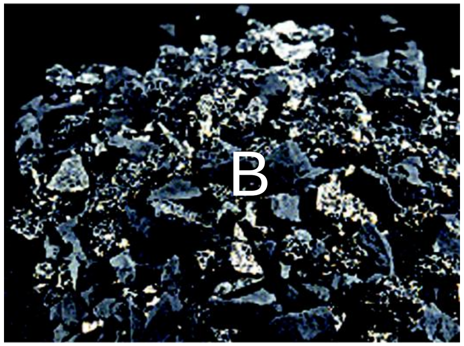
- Properties of the active metals
  - **Group 1A (*ns*^{1})**
    - Low *IE*
    - Never found in nature in elemental state
    - React with oxygen to form metal oxides
    - Peroxides and superoxides with some



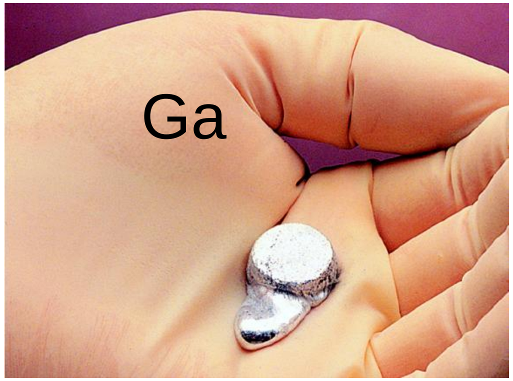
– **Group 2A (*ns*^{2})**

- Less reactive than 1A
- Some react with water to produce H_2
- Some react with acid to produce H_2



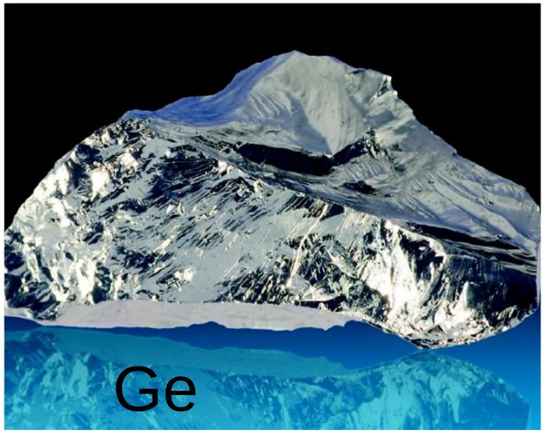


– **Group 3A (*ns^2np^1*)**
- Metalloid (B) and metals (all others)
- Al forms Al_2O_3 with oxygen
- Al forms +3 ions in acid
- Other metals form +1 and +3



– **Group 4A (*ns*²*np*²)**

- Nonmetal (C) metalloids (Si, Ge) and other metals
- Form +2 and +4 oxidation states
- Sn, Pb react with acid to produce H_2

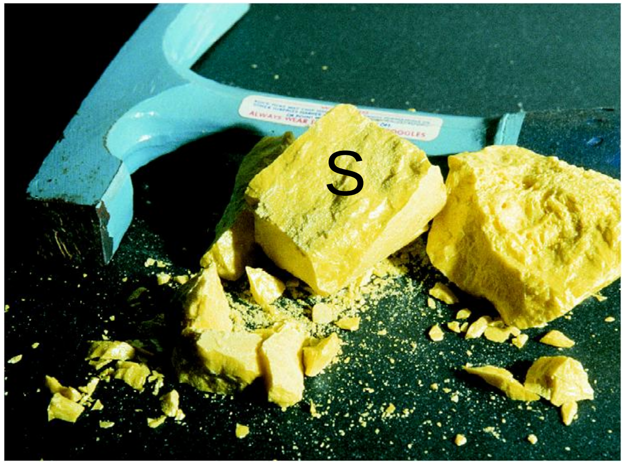


– **Group 5A (*ns*^{2}*np*^{3})**

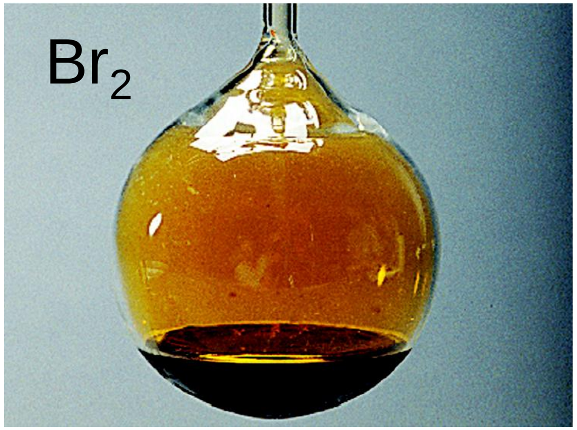
- Nonmetal (N_2, P) metalloid (As,Sb) and metal (Bi)
- Nitrogen, N_2 forms variety of oxides
- Phosphorus, P_4
- As, Sb, Bi (crystalline)
- HNO_3 and H_3PO_4 important industrially

– **Group 6A (*ns^2np^4*)**

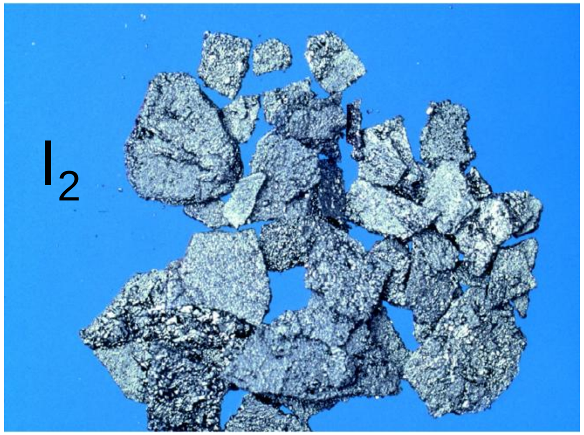
- Nonmetals (O, S, Se)
- Metalloids (Te, Po)
- Oxygen, O_2
- Sulfur,S_8
- Selenium, Se_8
- Te, Po (crystalline)
- SO_2, SO_3, H_2S, H_2SO_4

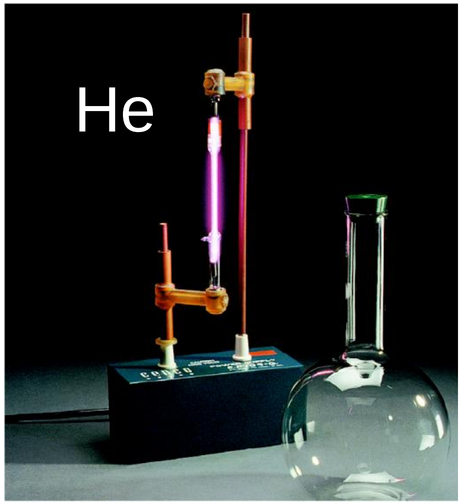


– **Group 7A (*ns^2np^5*)**

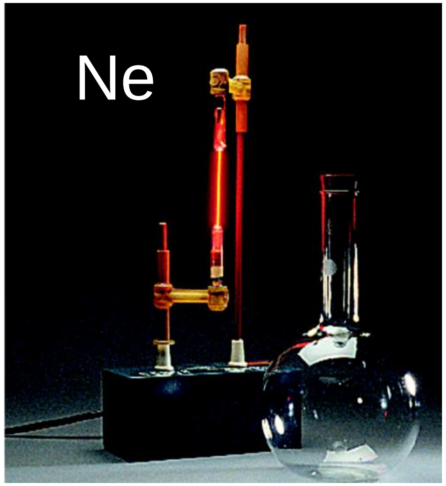


- All diatomic
- Do not exist in elemental form in nature
- Form ionic “salts”
- Form molecular compounds with each other





– **Group 8A (*ns^2np^6*)**
- All monatomic
- Filled valence shells
- Considered “inert” until 1963 when Xe and Kr were used to form compounds
- No major commercial use



- Comparison of 1A and 1B
  - Have single valence electron
  - Properties differ
  - Group 1B much less reactive than 1A
  - High *IE* of 1B - incomplete shielding of nucleus by inner "*d*" and outer "*s*" electrons of 1B strongly attracted to nucleus
  - 1B metals often found elemental in nature (coinage metals)

- Properties of oxides within a period
  - Metal oxides are usually basic

$$Na_2O(s) + H_2O(l) \longrightarrow 2NaOH(aq)$$

  - Nonmetal oxides are usually acidic

$$SO_3(g) + H_2O(l) \longrightarrow 2H_2SO_4(aq)$$

  - Amphoteric oxides are located at intermediate positions on the periodic table

$$Al_2O_3(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2O(l)$$

$$Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \longrightarrow 2NaAl(OH)_4(aq)$$

TABLE 7.4 Some Properties of Oxides of the Third-Period Elements

	Na_2O	MgO	Al_2O_3	SiO_2	P_4O_{10}	SO_3	Cl_2O_7
Type of compound	←	Ionic	→	←	Molecular		→
Structure	←	Extensive three-dimensional		→	← Discrete molecular units →		
Melting point (°C)	1275	2800	2045	1610	580	16.8	−91.5
Boiling point (°C)	?	3600	2980	2230	?	44.8	82
Acid-base nature	Basic	Basic	Amphoteric	←	Acidic		→

Key Points

- Development of the periodic table
- Modern table and its arrangement
- Main group elements
- Valence electrons
- Effective nuclear charge and relationship to periodic trends
- Atomic radius (ionic radii, covalent radii, metallic radii)

Key Points

- Ionization energy (*IE*) - trends of 1st and multiple *IE’*s
- Electron affinity (*EA*) - trends
- Properties of metals, metalloids and nonmetals
- Isoelectronic - predict charges of ions and electron configurations of ions

Key Points

- Write and/or recognize an isoelectronic series
- Characteristics of main group elements
- Know the most reactive metal and nonmetal groups and why
- Variability among groups
- Acidic, basic and amphoteric substances