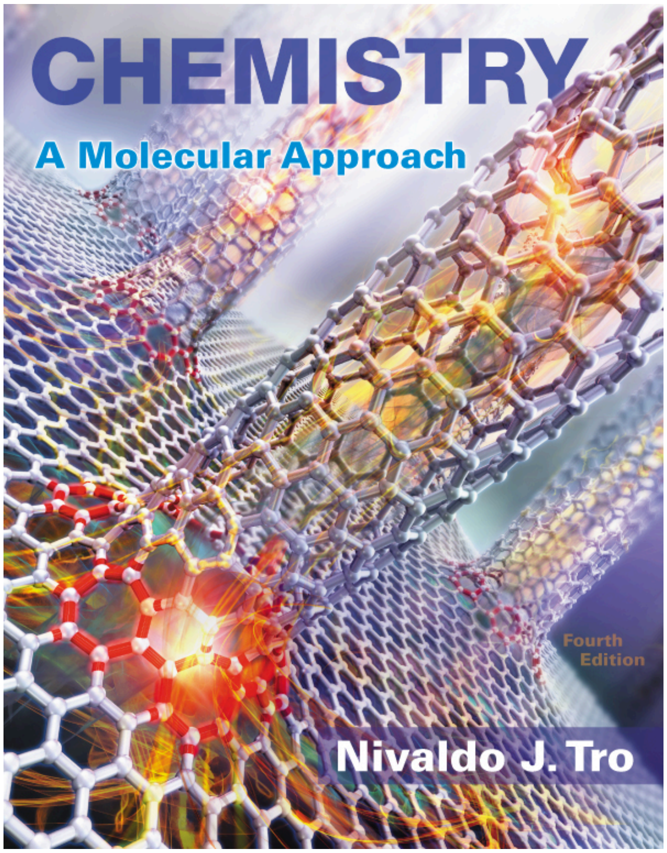


Lecture Presentation

Chapter 8

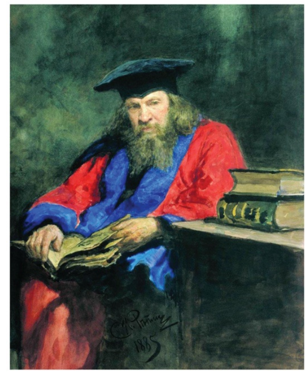
Periodic Properties of the Elements

Nerve Signal Transmission

- Movement of ions across cell membranes is the basis for the transmission of nerve signals.
- Na^+ and K^+ ions are pumped across membranes in opposite directions through ion channels.
  - Na^+ out and K^+ in
- The ion channels can differentiate Na^+ from K^+ by their difference in size.
- Ion size and other properties of atoms are **periodic properties—**properties whose values can be predicted based on the element's position on the periodic table.

3	Li	6.94
11	Na	22.99
19	K	39.10
37	Rb	85.47
55	Cs	132.91
87	Fr	(223.02)

Mendeleev (1834–1907)



- Ordered elements by atomic mass
- Saw a repeating pattern of properties
- **Periodic law**: When the elements are arranged in order of increasing atomic mass, certain sets of properties recur periodically.
- Put elements with similar properties in the same column
- Used pattern to predict properties of undiscovered elements
- Where atomic mass order did not fit other properties, he reordered by other properties.
  - Te and I

Mendeleev's Predictions

Gallium (eka-aluminum)

	Mendeleev's predicted properties	Actual properties
Atomic mass	About 68 amu	69.72 amu
Melting point	Low	29.8 °C
Density	5.9 g/cm^3	5.90 g/cm^3
Formula of oxide	X_2O_3	Ga_2O_3
Formula of chloride	XCl_3	$GaCl_3$

Germanium (eka-silicon)

	Mendeleev's predicted properties	Actual properties
Atomic mass	About 72 amu	72.64 amu
Density	5.5 g/cm^3	5.35 g/cm^3
Formula of oxide	XO_2	GeO_2
Formula of chloride	XCl_4	$GeCl_4$

What versus Why

- Mendeleev's periodic law allows us to predict *what* the properties of an element will be based on its position on the table.
- It doesn't explain why the pattern exists.
- Quantum mechanics is a theory that explains *why* the periodic trends in the properties exist.
  – Knowing *why* allows us to predict *what*.

Electron Configurations

- Quantum-mechanical theory describes the behavior of electrons in atoms.
- The electrons in atoms exist in orbitals.
- A description of the orbitals occupied by electrons is called an **electron configuration**.

H $1s^1$ Number of electrons in orbital

Orbital

How Electrons Occupy Orbitals

- Calculations with Schrödinger’s equation show that hydrogen’s one electron occupies the lowest energy orbital in the atom.
- Schrödinger’s equation calculations for multielectron atoms cannot be exactly solved.
  - Due to additional terms added for electron–electron interactions
- Approximate solutions show the orbitals to be hydrogen-like.
- Two additional concepts affect multielectron atoms: electron spin and energy splitting of sublevels.

Electron Spin

- Experiments by Stern and Gerlach showed that a beam of silver atoms is split in two by a magnetic field.
- The experiment reveals that the electrons spin on their axis.
- As they spin, they generate a magnetic field.
  - Spinning charged particles generates a magnetic field.
- If there is an even number of electrons, about half the atoms will have a net magnetic field pointing “north” and the other half will have a net magnetic field pointing “south.”

The Property of Electron Spin

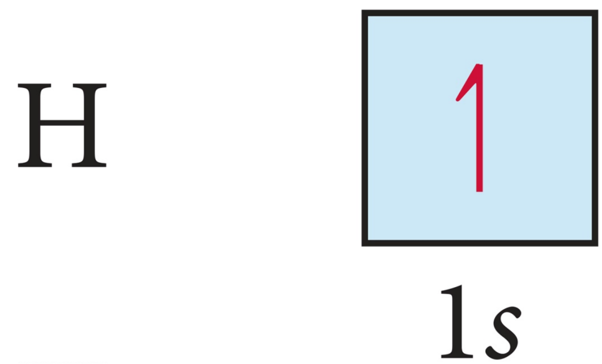
- Spin is a fundamental property of all electrons.
- All electrons have the same amount of spin.
- The orientation of the electron spin is quantized; it can be in only one direction or its opposite.
  - Spin up or spin down
- The electron's spin adds a fourth quantum number to the description of electrons in an atom, called the **spin quantum number, m_s**.
  - Not in the Schrödinger equation

Spin Quantum Number, m_s, and Orbital Diagrams

- m_s can have values of +½ or −½.
- Orbital diagrams use a square to represent each orbital and a half-arrow to represent each electron in the orbital.
- By convention, a half-arrow pointing up is used to represent an electron in an orbital with spin up.
- Spins must cancel in an orbital.
  - **Paired**

Orbital Diagrams

- We often represent an orbital as a square and the electrons in that orbital as arrows.
  - The direction of the arrow represents the spin of the electron.



Pauli Exclusion Principle

- No two electrons in an atom may have the same set of four quantum numbers.

- Therefore, no orbital may have more than two electrons, and they must have opposite spins.

Pauli Exclusion Principle

- Knowing the number of orbitals in a sublevel allows us to determine the maximum number of electrons in the sublevel:
  - *s* sublevel has 1 orbital; therefore, it can hold 2 electrons.
  - *p* sublevel has 3 orbitals; therefore, it can hold 6 electrons.
  - *d* sublevel has 5 orbitals; therefore, it can hold 10 electrons.
  - *f* sublevel has 7 orbitals; therefore, it can hold 14 electrons.

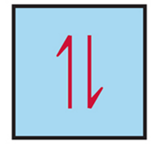
Quantum Numbers of Helium's Electrons

- Helium has two electrons.
- Both electrons are in the first energy level.
- Both electrons are in the *s* orbital of the first energy level.
- Because they are in the same orbital, they must have opposite spins.

Allowed Quantum Numbers

Electron configuration | Orbital diagram

He $1s^2$



$1s$

n	l	m_l	m_s
1	0	0	$+\frac{1}{2}$
1	0	0	$-\frac{1}{2}$

Sublevel Splitting in Multielectron Atoms

- The sublevels in each principal energy shell of hydrogen all have the same energy or other single electron systems.
- We call orbitals with the same energy **degenerate**.
- For multielectron atoms, the energies of the sublevels are split.
  - Caused by charge interaction, shielding, and penetration
- The lower the value of the *l* quantum number, the less energy the sublevel has.
  - s (l = 0) < p (l = 1) < d (l = 2) < f (l = 3)

Coulomb's Law

$$E = \frac{1}{4\pi\varepsilon_0}\frac{q_1 q_2}{r}$$

- Coulomb's law describes the attractions and repulsions between charged particles.
- For like charges, the potential energy (*E*) is positive and decreases as the particles get farther apart as *r* increases.
- For opposite charges, the potential energy is negative and becomes more negative as the particles get closer together.
- The strength of the interaction increases as the size of the charges increases.
  – Electrons are more strongly attracted to a nucleus with a 2+ charge than a nucleus with a 1+ charge.

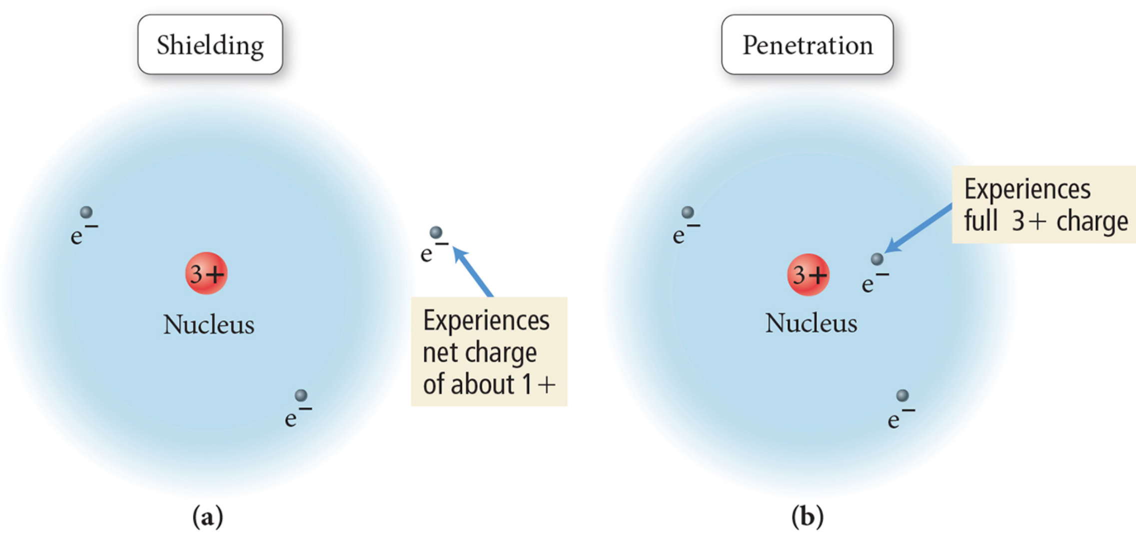
Shielding and Effective Nuclear Charge

- Each electron in a multielectron atom experiences both the attraction to the nucleus and the repulsion by other electrons in the atom.
- These repulsions cause the electron to have a net reduced attraction to the nucleus; it is **shielded** from the nucleus.
- The total amount of attraction that an electron feels for the nucleus is called the **effective nuclear charge** of the electron.

Penetration

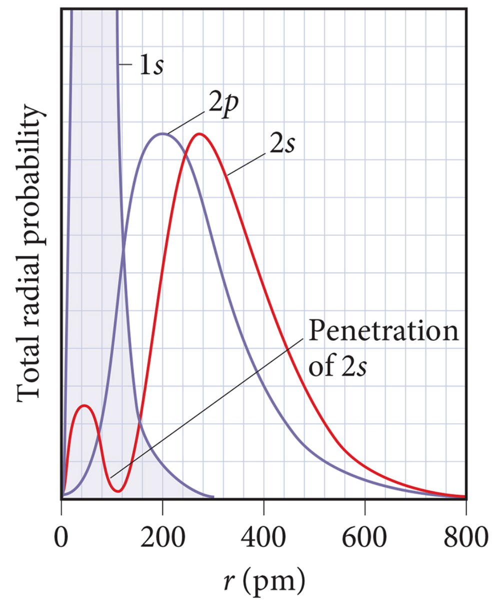
- The closer an electron is to the nucleus, the more attraction it experiences.
- The better an outer electron is at **penetrating** through the electron cloud of inner electrons, the more attraction it will have for the nucleus.
- The degree of penetration is related to the orbital's radial distribution function.
  - In particular, the distance the maxima of the function are from the nucleus

Shielding and Penetration



Penetration and Shielding

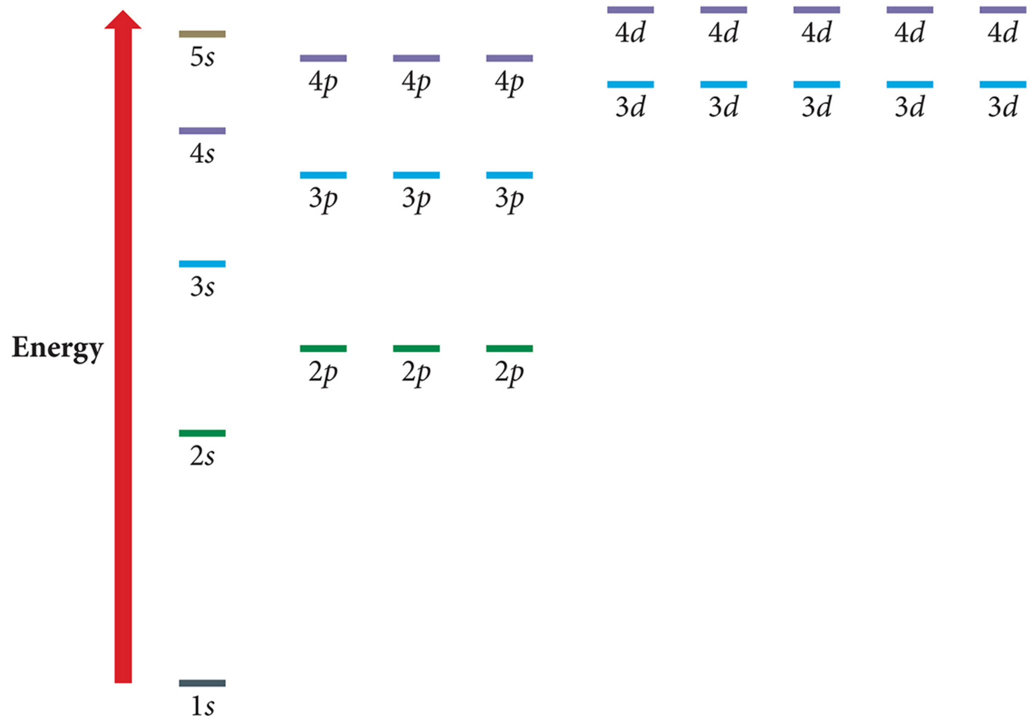
- The radial distribution function shows that the 2*s* orbital penetrates more deeply into the 1*s* orbital than does the 2*p.*
- The weaker penetration of the 2*p* sublevel means that electrons in the 2*p* sublevel experience more repulsive force; they are more shielded from the attractive force of the nucleus.
- The deeper penetration of the 2*s* electrons means electrons in the 2*s* sublevel experience a greater attractive force to the nucleus and are not shielded as effectively.



Effect of Penetration and Shielding

- Penetration causes the energies of sublevels in the same principal level to not be degenerate.
- In the fourth and fifth principal levels, the effects of penetration become so important that the *s* orbital lies lower in energy than the *d* orbitals of the previous principal level.
- The energy separations between one set of orbitals and the next become smaller beyond the 4*s.*
  - The ordering can therefore vary among elements, causing variations in the electron configurations of the transition metals and their ions.

General Energy Ordering of Orbitals for Multielectron Atoms



Filling the Orbitals with Electrons

- Energy levels and sublevels fill from lowest energy to highest:
  - $s \rightarrow p \rightarrow d \rightarrow f$
  - Aufbau principle
- Orbitals that are in the same sublevel have the same energy.
- No more than two electrons per orbital
  - Pauli exclusion principle
- When filling orbitals that have the same energy, place one electron in each before completing pairs.
  - Hund's rule

Electron Configuration of Atoms in Their Ground State

- The electron configuration is a listing of the sublevels in order of filling with the number of electrons in that sublevel written as a superscript.

  Kr = 36 electrons = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
- A short-hand way of writing an electron configuration is to use the symbol of the previous noble gas in brackets [] to represent all the inner electrons and then just write the last set.

Rb = 37 electrons = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$ = [Kr]$5s^1$

Electron Configurations

Symbol	Number of electrons	Electron configuration	Orbital diagram
Li	3	$1s^2 2s^1$	1s: [↑↓] 2s: [↑]
Be	4	$1s^2 2s^2$	1s: [↑↓] 2s: [↑↓]
B	5	$1s^2 2s^2 2p^1$	1s: [↑↓] 2s: [↑↓] 2p: [↑][ ][ ]
C	6	$1s^2 2s^2 2p^2$	1s: [↑↓] 2s: [↑↓] 2p: [↑][↑][ ]

Valence Electrons

- The electrons in all the sublevels with the highest principal energy shell are called the **valence electrons**.
- Electrons in lower energy shells are called **core electrons**.
- One of the most important factors in the way an atom behaves, both chemically and physically, is the number of valence electrons.

Outer Electron Configurations of Elements 1–18

1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
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** $2s^22p^6$
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** $3s^23p^6$

Electron Configuration and the Periodic Table

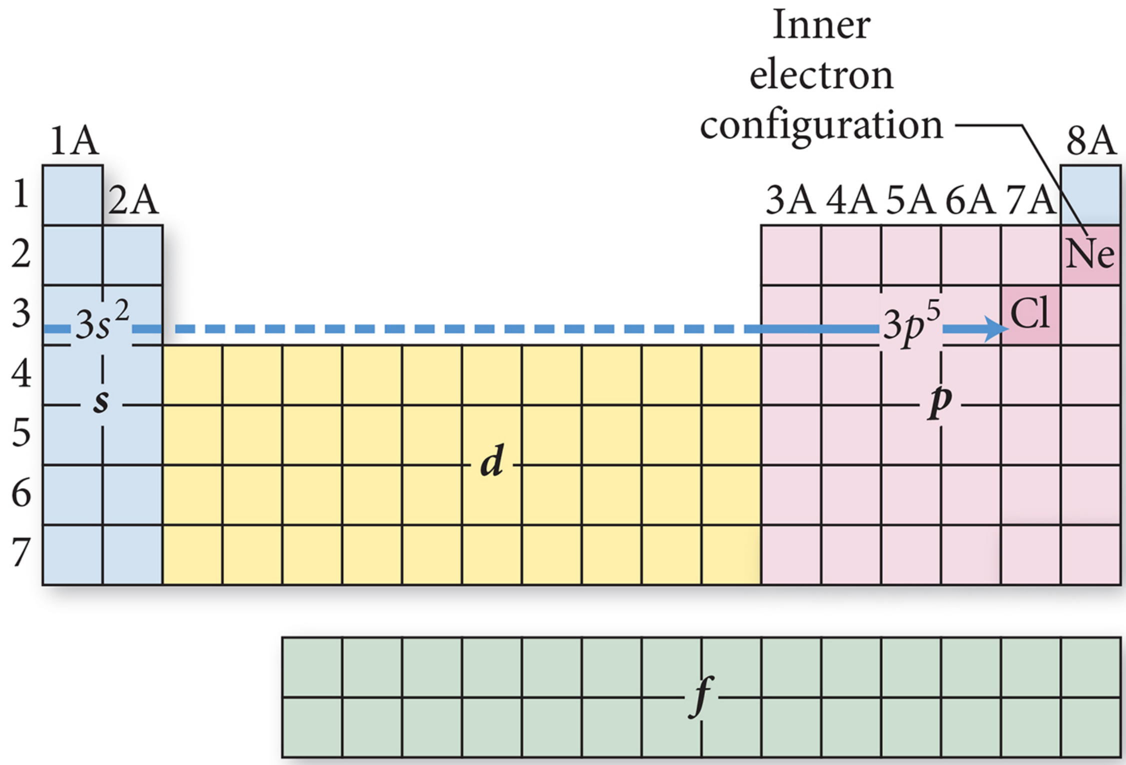
- The group number corresponds to the number of valence electrons.
- The length of each “block” is the maximum number of electrons the sublevel can hold.
- The period number corresponds to the principal energy level of the valence electrons.

Orbital Blocks of the Periodic Table

s-block elements; p-block elements; d-block elements; f-block elements

Periods	Groups 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	1 H $1s^1$																	2 He $1s^2$
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 $2s^22p^6$
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 $3s^23p^6$
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 $4s^23d^{10}$	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 $4s^24p^6$
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 $5s^25p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La $6s^25d^1$	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 $6s^25d^{10}$	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 $6s^26p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac $7s^26d^1$	104 Rf $7s^26d^2$	105 Db $7s^26d^3$	106 Sg $7s^26d^4$	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 **	114 Fl	115 **	116 Lv	117 **	118 **

Lanthanides	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 $6s^24f^{14}$	71 Lu $6s^24f^{14}5d^1$
Actinides	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 $7s^25f^{14}$	103 Lr $7s^25f^{14}6d^1$



Irregular Electron Configurations

- We know that, because of sublevel splitting, the 4*s* sublevel is lower in energy than the 3*d*; therefore, the 4*s* fills before the 3*d.*
- But the difference in energy is not large.
- Some of the transition metals have irregular electron configurations in which the *ns* only partially fills before the (n−1)*d* or doesn't fill at all.
- Therefore, their electron configuration must be found experimentally.

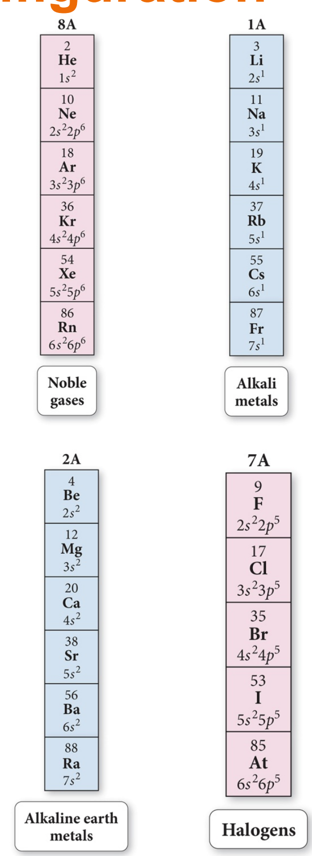
Irregular Electron Configurations

- Expected
- Cr = [Ar]$4s^2 3d^4$
- Cu = [Ar]$4s^2 3d^9$
- Mo = [Kr]$5s^2 4d^4$
- Ru = [Kr]$5s^2 4d^6$
- Pd = [Kr]$5s^2 4d^8$

- Found experimentally
- Cr = [Ar]$4s^1 3d^5$
- Cu = [Ar]$4s^1 3d^{10}$
- Mo = [Kr]$5s^1 4d^5$
- Ru = [Kr]$5s^1 4d^7$
- Pd = [Kr]$5s^0 4d^{10}$

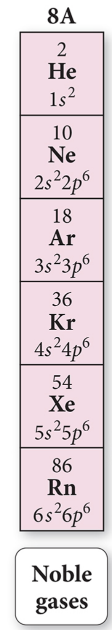
Properties and Electron Configuration

- The properties of the elements follow a periodic pattern.
  – Elements in the same column have similar properties.
  – The elements in a period show a pattern that repeats.
- The quantum-mechanical model explains this because the number of valence electrons and the types of orbitals they occupy are also periodic.



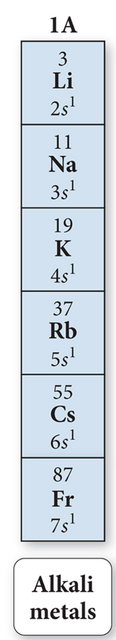
The Noble Gas Electron Configuration

- The noble gases have eight valence electrons.
  - Except for He, which has only two electrons
- They are especially nonreactive.
  - He and Ne are practically inert.
- The reason the noble gases are so nonreactive is that the electron configuration of the noble gases is especially stable.



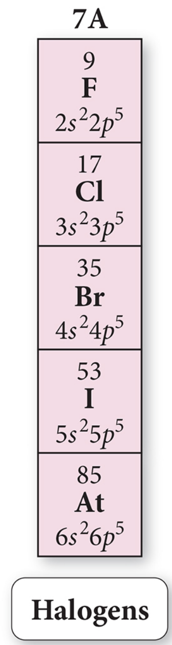
The Alkali Metals

- The **alkali** metals have one more electron than the previous noble gas.
- In their reactions, the alkali metals tend to lose one electron, resulting in the same electron configuration as a noble gas.
  - Forming a cation with a 1+ charge



The Halogens

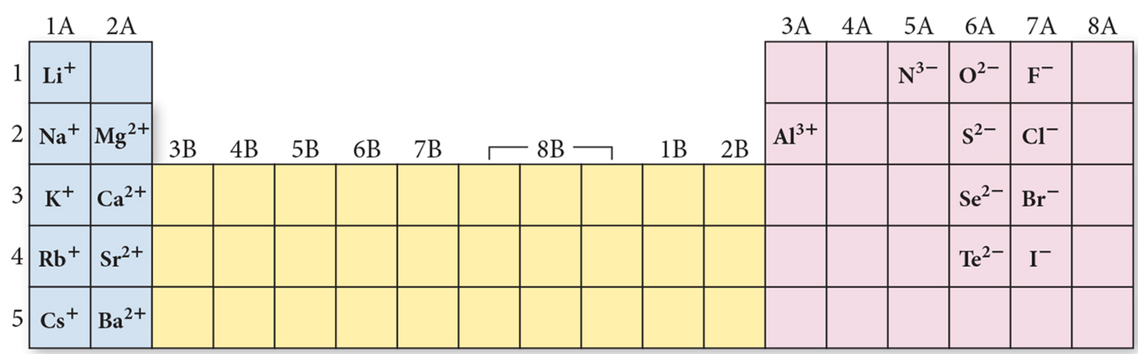
- Have one fewer electron than the next noble gas
- In their reactions with metals, the halogens tend to gain an electron and attain the electron configuration of the next noble gas, forming an anion with charge 1−.
- In their reactions with nonmetals, they tend to share electrons with the other nonmetal so that each attains the electron configuration of a noble gas.



Electron Configuration and Ion Charge

- We have seen that many metals and nonmetals form one ion and that the charge on that ion is predictable based on its position on the periodic table.
  - Group 1A = 1+, group 2A = 2+, group 7A = 1−, group 6A = 2−, etc.
- These atoms form ions that will result in an electron configuration that is the same as the nearest noble gas.

Elements That Form Ions with Predictable Charges



Electron Configuration of Anions in Their Ground State

- Anions are formed when nonmetal atoms gain enough electrons to have eight valence electrons.
  - Filling the *s* and *p* sublevels of the valence shell
- The sulfur atom has six valence electrons.

  S atom = $1s^22s^22p^63s^23p^4$
- To have eight valence electrons, sulfur must gain two more.

  S^{2-} anion = $1s^22s^22p^63s^23p^6$

Electron Configuration of Cations in Their Ground State

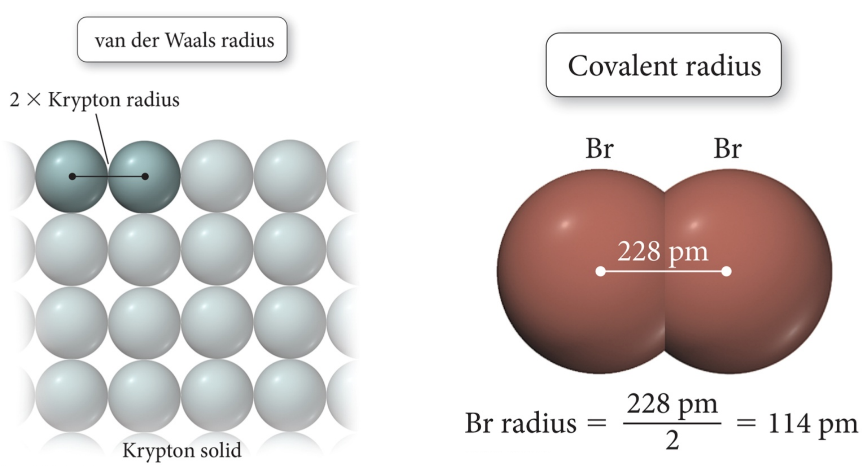
- Cations are formed when a metal atom loses all its valence electrons, resulting in a new lower energy level valence shell.
  - However, the process is always endothermic.

- The magnesium atom has two valence electrons.
  Mg atom = $1s^2 2s^2 2p^6 3s^2$

- When magnesium forms a cation, it loses its valence electrons.
  Mg^{2+} cation = $1s^2 2s^2 2p^6$

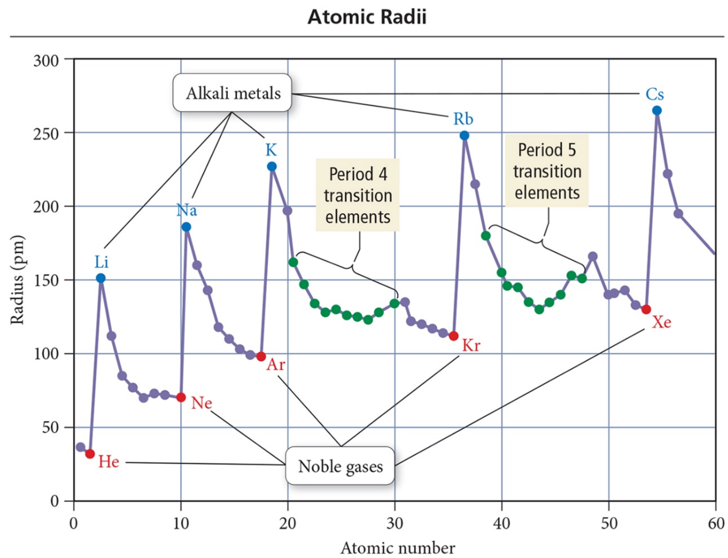
Trend in Atomic Radius: Main Group

- There are several methods for measuring the radius of an atom, and they give slightly different numbers.
  - Van der Waals radius = nonbonding
  - Covalent radius = bonding radius
  - Atomic radius is an average radius of an atom based on measuring large numbers of elements and compounds.



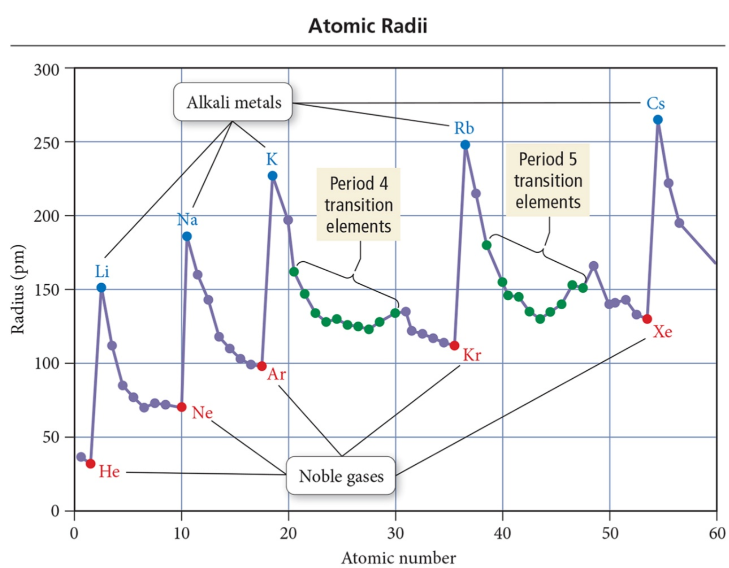
Trend in Atomic Radius: Main Group

- Atomic radius decreases across period (left to right).
  - Adding electrons to same valence shell
  - Effective nuclear charge increases
  - Valence shell held closer

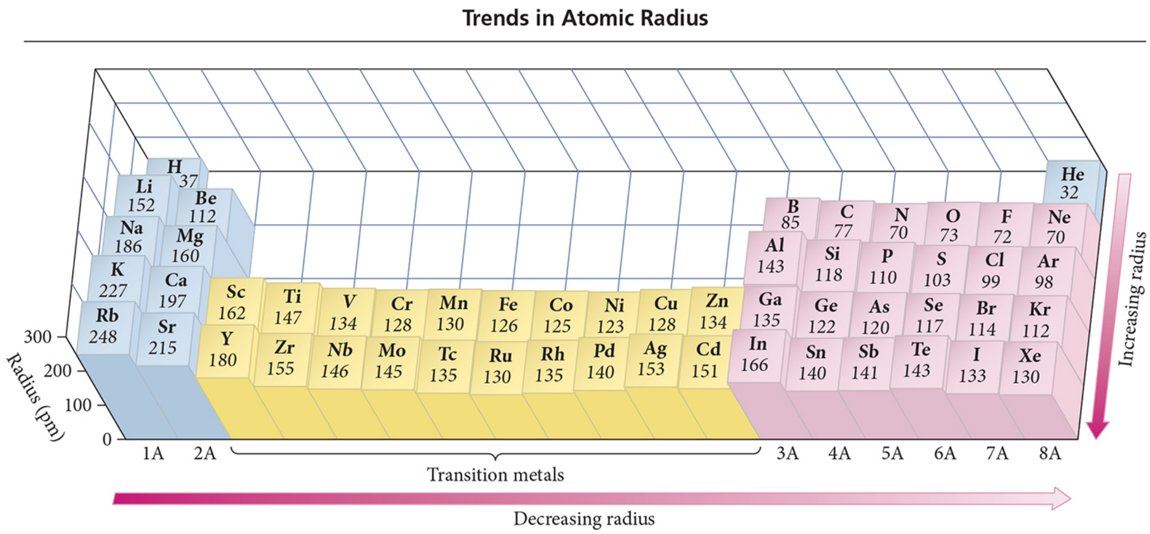


Trend in Atomic Radius: Main Group

- Atomic radius increases down group.
  - Valence shell farther from nucleus
  - Effective nuclear charge fairly close



Periodic Trends in Atomic Radius



Shielding

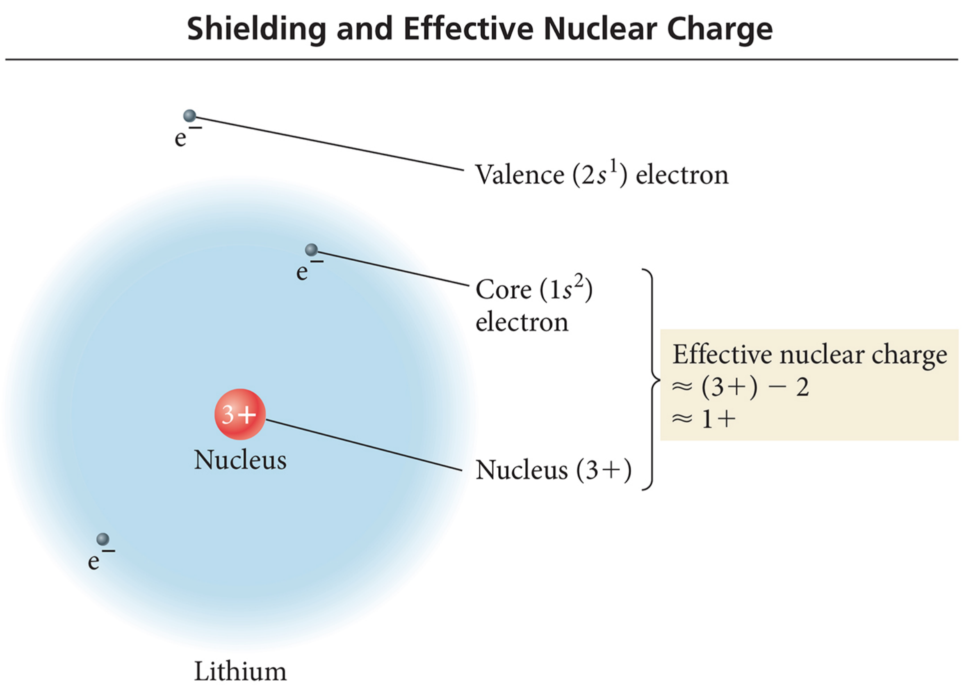
- In a multielectron system, electrons are simultaneously attracted to the nucleus and repelled by each other.
- Outer electrons are *shielded* from the nucleus by the core electrons.
  - Screening or shielding effect
  - Outer electrons do not effectively screen for each other.
- The shielding causes the outer electrons to *not* experience the full strength of the nuclear charge.

Effective Nuclear Charge

- The **effective nuclear charge** is a net positive charge that is attracting a particular electron.
- **Z** is the nuclear charge, and **S** is the number of electrons in lower energy levels.
  - Electrons in the same energy level contribute to screening, but since their contribution is so small they are not part of the calculation.
  - Trend is *s* > *p* > *d* > *f*.

$$\mathbf{Z_{effective}} = Z - S$$

Screening and Effective Nuclear Charge



Quantum-Mechanical Explanation for the Group Trend in Atomic Radius

- The size of an atom is related to the distance the valence electrons are from the nucleus.
- The larger the orbital an electron is in, the farther its most probable distance will be from the nucleus and the less attraction it will have for the nucleus.

Quantum-Mechanical Explanation for the Group Trend in Atomic Radius

- Traversing down a group adds a principal energy level.
- The larger the principal energy level an orbital is in, the larger its volume.
- Quantum-mechanics predicts the atoms should get larger down a column.

Quantum-Mechanical Explanation for the Period Trend in Atomic Radius

- The larger the effective nuclear charge an electron experiences, the stronger the attraction it will have for the nucleus.
- The stronger the attraction the valence electrons have for the nucleus, the closer their average distance will be to the nucleus.
- Traversing across a period increases the effective nuclear charge on the valence electrons.
- Quantum-mechanics predicts the atoms should get smaller across a period.

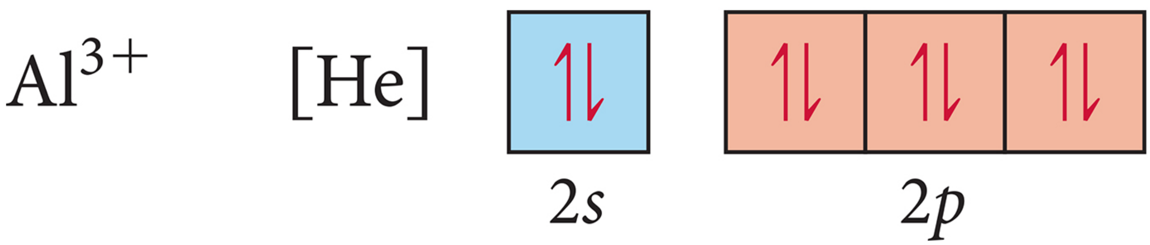
Trends in Atomic Radius: Transition Metals

- Atoms in the same group increase in size down the column.
- Atomic radii of transition metals are roughly the same size across the *d* block.
  - Much less difference than across main-group elements
  - Valence shell ns^2, not the $(n-1)d$ electrons
  - Effective nuclear charge on the ns^2 electrons approximately the same

Electron Configurations of Main Group Cations in Their Ground State

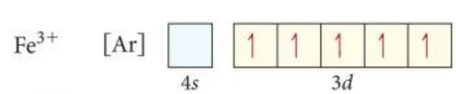
- Cations form when the atom loses electrons from the valence shell.

$$\text{Al atom} = 1s^2 2s^2 2p^6 3s^2 3p^1$$
$$\text{Al}^{3+}\ \text{ion} = 1s^2 2s^2 2p^6$$



Electron Configurations of Transition Metal Cations in Their Ground State

- When transition metals form cations, the first electrons removed are the valence electrons, even though other electrons were added after.
- Electrons may also be removed from the sublevel closest to the valence shell after the valence electrons.
- The iron atom has two valence electrons:
  Fe atom = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
- When iron forms a cation, it first loses its valence electrons:
  Fe^{2+} cation = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$
- It can then lose $3d$ electrons:
  Fe^{3+} cation = $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$



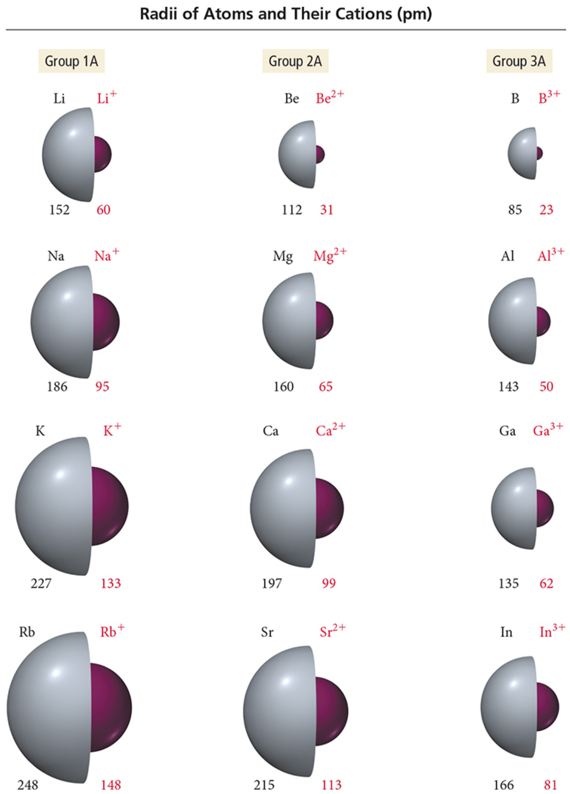
Magnetic Properties of Transition Metal Atoms and Ions

- Electron configurations that result in unpaired electrons mean that the atom or ion will have a net magnetic field; this is called **paramagnetism**.
  – Will be attracted to a magnetic field
- Electron configurations that result in all paired electrons mean that the atom or ion will have no magnetic field; this is called **diamagnetism**.
  – Slightly repelled by a magnetic field

Trends in Ionic Radius

- Ions in the same group have the same charge.
- Ion size increases down the column.
  – Higher valence shell, larger ion
- Cations are smaller than neutral atoms; anions are larger than neutral atoms.
- Cations are smaller than anions.
  – Except Rb^+ and Cs^+, bigger or same size as $F^−$ and $O^{2−}$
- Larger positive charge = smaller cation
  – For isoelectronic species
  – Isoelectronic = same electron configuration
- Larger negative charge = larger anion
  – For isoelectronic species

Periodic Trends in Ionic Radius



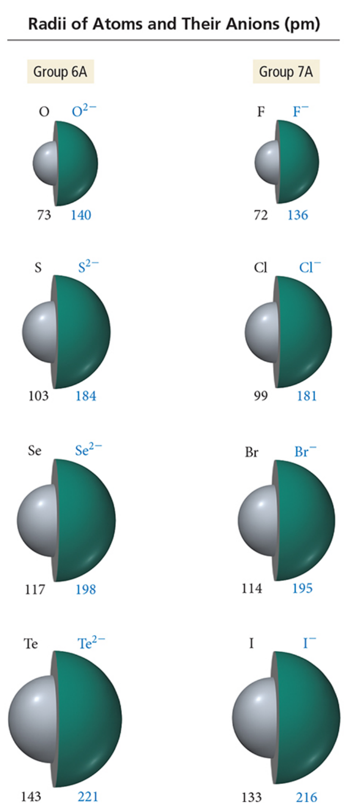
Explanation for the Trends in Cation Radius

- When atoms form cations, the valence electrons are removed.
- The farthest electrons from the nucleus are the *p* or *d* electrons in the (n − 1) energy level.
- This results in the cation being smaller than the atom.

Explanation for the Trends in Cation Radius

- These "new valence electrons" also experience a larger effective nuclear charge than the "old valence electrons," shrinking the ion even more.
- Traversing down a group increases the (n − 1) level, causing the cations to get larger.
- Traversing to the right across a period increases the effective nuclear charge for isoelectronic cations, causing the cations to get smaller.

Periodic Trends in Anionic Radius



Explanation for the Trends in Anion Radius

- When atoms form anions, electrons are added to the valence shell.
- These "new valence electrons" experience a smaller effective nuclear charge than the "old valence electrons," increasing the size.
- The result is that the anion is larger than the atom.

Explanation for the Trends in Anion Radius

- Traversing down a group increases the *n* level, causing the anions to get larger.
- Traversing to the right across a period decreases the effective nuclear charge for isoelectronic anions, causing the anions to get larger.

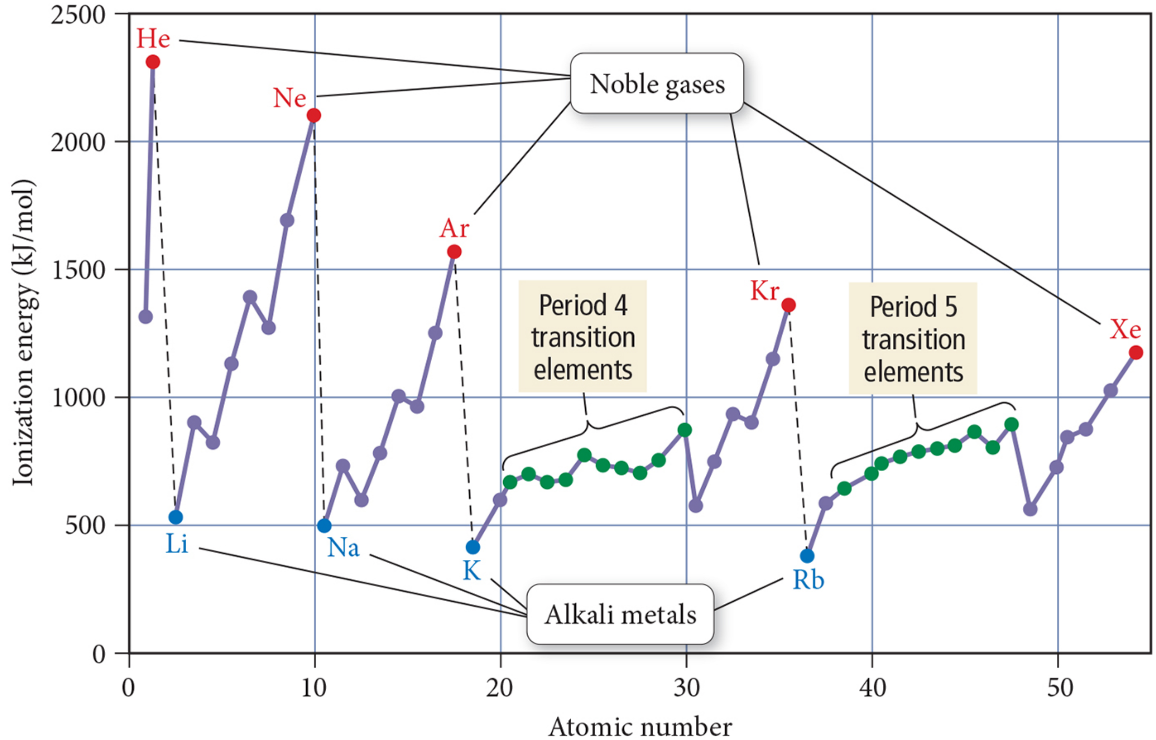
Ionization Energy (IE)

- Minimum energy needed to remove an electron from an atom or ion
  - Gas state
  - Endothermic process
  - Valence electron easiest to remove, lowest IE
  - $M(g) + IE_1 \rightarrow M^{1+}(g) + 1\ e^{-}$
  - $M^{+1}(g) + IE_2 \rightarrow M^{2+}(g) + 1\ e^{-}$
    - First ionization energy = energy to remove electron from neutral atom
    - Second IE = energy to remove electron from 1+ ion, etc.

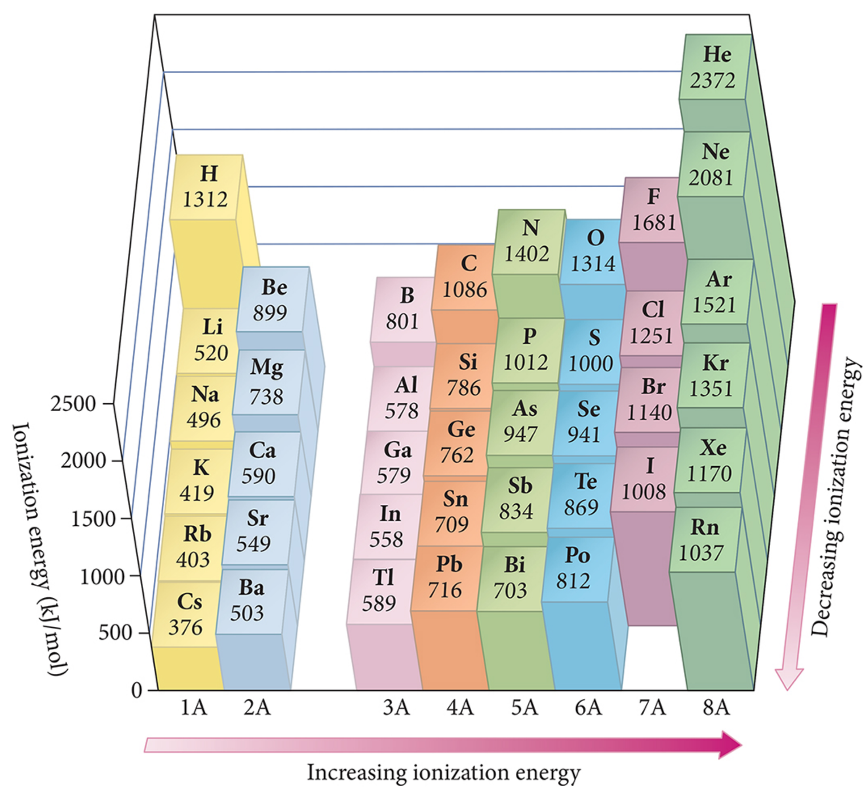
General Trends in First Ionization Energy

- The larger the effective nuclear charge on the electron, the more energy it takes to remove it.
- The farther the most probable distance the electron is from the nucleus, the less energy it takes to remove it.
- First IE *decreases* down the group.
  - Valence electron farther from nucleus
- First IE generally *increases* across the period.
  - Effective nuclear charge increases

First Ionization Energies



Trends in First Ionization Energy



Explanation for the Trends in First Ionization Energy

- The strength of attraction is related to the most probable distance the valence electrons are from the nucleus and the effective nuclear charge the valence electrons experience.
- The larger the orbital an electron is in, the farther its most probable distance will be from the nucleus and the less attraction it will have for the nucleus.
- Quantum-mechanics predicts the atom's first ionization energy should get lower down a column.

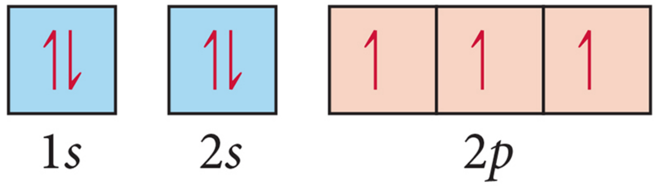
Explanation for the Trends in First Ionization Energy

- Traversing across a period increases the effective nuclear charge on the valence electrons.
- Quantum-mechanics predicts the atom's first ionization energy should get larger across a period.

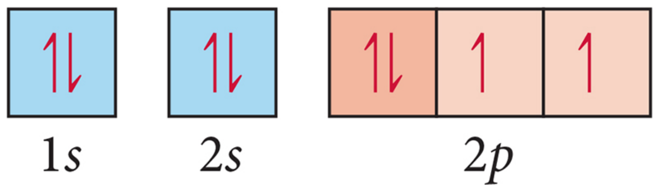
Exceptions in the First IE Trends

- First ionization energy generally increases from left to right across a period.
- Except from 2A to 3A and 5A to 6A

N $1s^2 2s^2 2p^3$

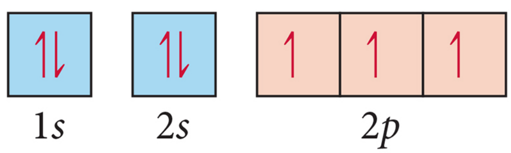


O $1s^2 2s^2 2p^4$



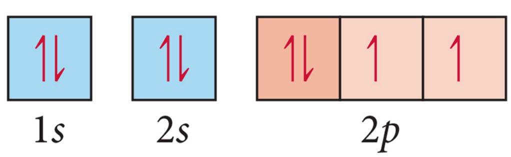
Exceptions in the First Ionization Energy Trends, N and O

N $1s^2 2s^2 2p^3$



To ionize N, you must break up a half-full sublevel, which costs extra energy.

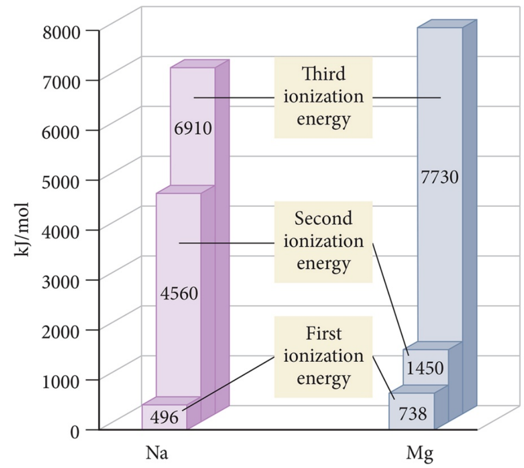
O $1s^2 2s^2 2p^4$



When you ionize O, you get a half-full sublevel, which costs less energy.

Trends in Successive Ionization Energies

- Removal of each successive electron costs more energy.
  - Shrinkage in size due to having more protons than electrons
  - Outer electrons closer to the nucleus; therefore harder to remove
- There's a regular increase in energy for each successive valence electron.
- There's a large increase in energy when core electrons are removed.



Trends in Second and Successive Ionization Energies

TABLE 8.1 Successive Ionization Energies for the Elements Sodium through Argon (kJ/mol)

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	496	4560					
Mg	738	1450	7730		**Core electrons**		
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 Affinity

- Energy is released when a neutral atom gains an electron.
  - Gas state
  - $M(g) + 1e^- \rightarrow M^{1-}(g) + EA$
- Electron affinity is defined as exothermic (−) but may actually be endothermic (+).
  - Some alkali earth metals and all noble gases are endothermic. Why?
- The more energy that is released, the larger the electron affinity.
  - The more negative the number, the larger the EA.

Trends in Electron Affinity

- Alkali metals decrease electron affinity down the column.
  - But not all groups do
  - Generally irregular increase in EA from second period to third period
- "Generally" increases across period
  - Becomes more negative from left to right
  - Not absolute
  - Group 5A generally lower EA than expected because extra electron must pair
  - Groups 2A and 8A generally very low EA because added electron goes into higher energy level or sublevel
- Highest EA in any period = halogen

Electron Affinities (kJ/mol)

1A	2A	3A	4A	5A	6A	7A	8A
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

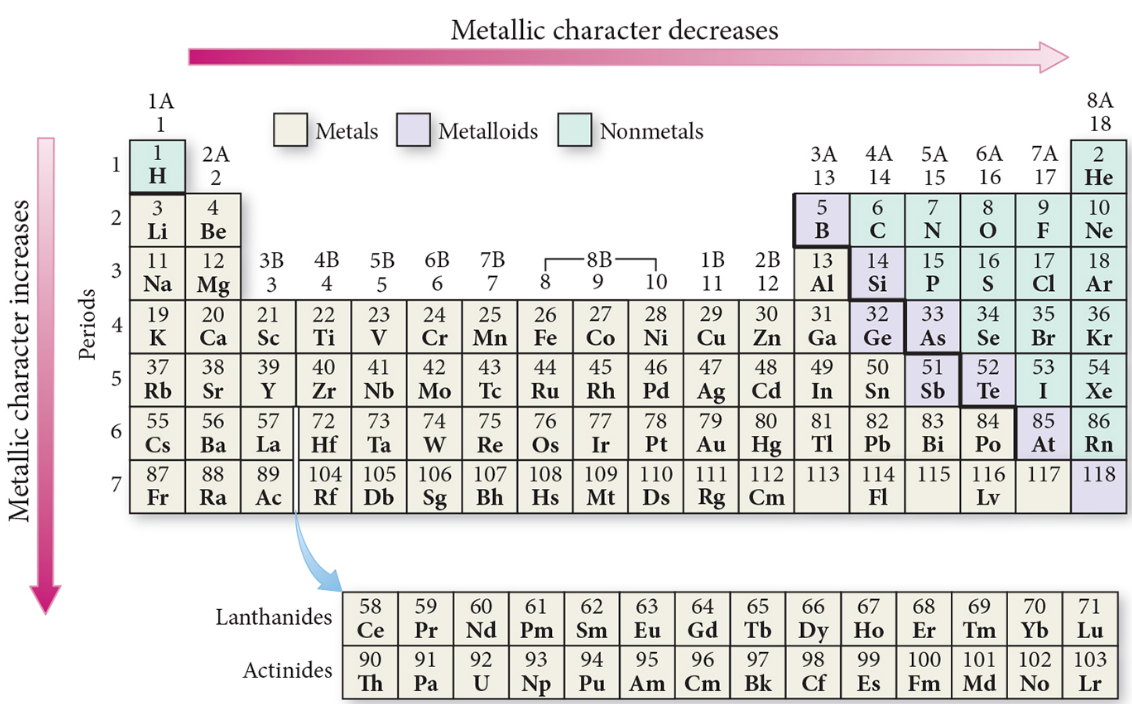
Properties of Metals and Nonmetals

- Metals
  - Malleable and ductile
  - Shiny, lustrous, reflect light
  - Conduct heat and electricity
  - Most oxides basic and ionic
  - Form cations in solution
  - Lose electrons in reactions—**oxidized**
- Nonmetals
  - Brittle in solid state
  - Dull, nonreflective solid surface
  - Electrical and thermal insulators
  - Most oxides are acidic and molecular
  - Form anions and polyatomic anions
  - Gain electrons in reactions—**reduced**

Metallic Character

- **Metallic character** is how closely an element’s properties match the ideal properties of a metal.
  - More malleable and ductile, better conductors, and easier to ionize
- Metallic character *decreases left to right* across a period.
  - Metals found at the left of the period, and nonmetals to the right
- Metallic character *increases down* the column.
  - Nonmetals found at the top of the middle main-group elements, and metals found at the bottom

Trends in Metallic Character



Explanation for the Trends in Metallic Character

- Metals generally have smaller first ionization energies, and nonmetals generally have larger electron affinities.
  - Except for the noble gases
- ∴ quantum mechanics predicts the atom's metallic character should increase down a column because the valence electrons are not held as strongly.
- ∴ quantum mechanics predicts the atom's metallic character should decrease across a period because the valence electrons are held more strongly and the electron affinity increases.

Trends in the Alkali Metals

- Atomic radius increases down the column.
- Ionization energy decreases down the column.
- Very low ionization energies
  - Good reducing agents; easy to oxidize
  - Very reactive; not found uncombined in nature
  - React with nonmetals to form salts
  - Compounds generally soluble in water ∴ found in seawater
- Electron affinity decreases down the column.
- Melting point decreases down the column.
  - All very low MP for metals
- Density increases down the column.
  - Except K
  - In general, the increase in mass is greater than the increase in volume.

Alkali Metals

TABLE 8.2 Properties of the Alkali Metals*

Element	Electron Configuration	Atomic Radius (pm)	IE_1 (kJ/mol)	Density at 25 °C (g/cm^3)	Melting Point (°C)
Li	$[He]\ 2s^1$	152	520	0.535	181
Na	$[Ne]\ 3s^1$	186	496	0.968	102
K	$[Ar]\ 4s^1$	227	419	0.856	98
Rb	$[Kr]\ 5s^1$	248	403	1.532	39
Cs	$[Xe]\ 6s^1$	265	376	1.879	29

*Francium is omitted because it has no stable isotopes.

Reactions of the Alkali Metals with Water

Lithium Sodium Potassium

Trends in the Halogens

- Atomic radius increases down the column.
- Ionization energy decreases down the column.
- Very high electron affinities
  - Good oxidizing agents; easy to reduce
  - Very reactive; not found uncombined in nature
  - React with metals to form salts
  - Compounds generally soluble in water ∴ found in seawater

Trends in the Halogens

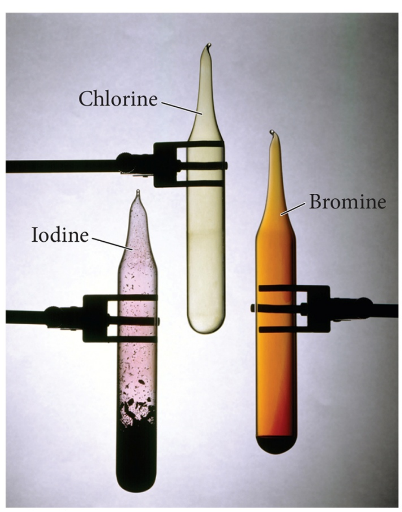
- Reactivity increases down the column.
- They react with hydrogen to form HX, acids.
- Melting point and boiling point increase down the column.
- Density increases down the column.
  - In general, the increase in mass is greater than the increase in volume.

Halogens

TABLE 8.3 Properties of the Halogens*

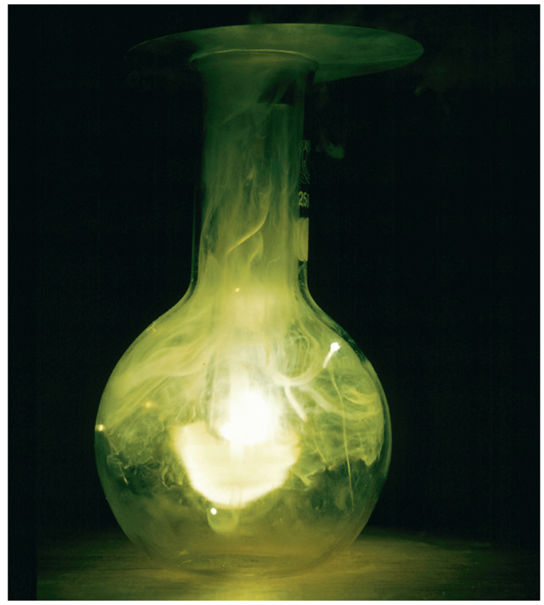
Element	Electron Configuration	Atomic Radius (pm)	EA (kJ/mol)	Melting Point (°C)	Boiling Point (°C)	Density of Liquid (g/cm^3)
F	[He] $2s^22p^5$	72	−328	−219	−188	1.51
Cl	[Ne] $3s^23p^5$	99	−349	−101	−34	2.03
Br	[Ar] $4s^24p^5$	114	−325	−7	59	3.19
I	[Kr] $5s^25p^5$	133	−295	114	184	3.96

*At is omitted because it is rare and radioactive.



Reactions of Alkali Metals with Halogens

- Alkali metals are oxidized to the 1+ ion.
- Halogens are reduced to the 1− ion.
- The ions then attach together by ionic bonds.
- The reaction is exothermic.



Reactions of Alkali Metals with Water

- Alkali metals are oxidized to the 1+ ion.
- H_2O is split into $H_2(g)$ and OH^- ion.
- The Li, Na, and K are less dense than the water, so they float on top.
- The ions then attach together by ionic bonds.
- The reaction is exothermic, and often the heat released ignites the $H_2(g)$.

Reactions of the Alkali Metals with Water

Lithium Sodium Potassium

Trends in the Noble Gases

- Atomic radius increases down the column.
- Ionization energy decreases down the column.
  – Very high IE
- Very unreactive
  – Only found uncombined in nature
  – Used as “inert” atmosphere when reactions with other gases would be undesirable

Trends in the Noble Gases

- Melting point and boiling point increase down the column.
  – All gases at room temperature
  – Very low boiling points
- Density increases down the column.
  – In general, the increase in mass is greater than the increase in volume.

Noble Gases

TABLE 8.4 Properties of the Noble Gases*

Element	Electron Configuration	Atomic Radius (pm)**	IE_1(kJ/mol)	Boiling Point (K)	Density of Gas (g/L at STP)
He	$1s^2$	32	2372	4.2	0.18
Ne	[He] $2s^22p^6$	70	2081	27.1	0.90
Ar	[Ne] $3s^23p^6$	98	1521	87.3	1.78
Kr	[Ar] $4s^24p^6$	112	1351	119.9	3.74
Xe	[Kr] $5s^25p^6$	130	1170	165.1	5.86

*Radon is omitted because it is radioactive.

**Since only the heavier noble gases form compounds, covalent radii for the smaller noble gases are estimated.