

WILEY

Chapter 7
The Quantum
Mechanical Atom

Chemistry, 7th Edition
International Student Version
Brady/Jespersen/Hyslop

Chapter in Context

- Define and use quantum numbers
- Write the ground state electron configuration of atoms
- Connect the organization of the periodic table to electron configurations
- Draw the shapes of s , p , and d orbitals
- Use electron configurations and the organization of the periodic table to predict chemical properties of elements

Wave Functions

Schrödinger's equation

- Solutions give wave functions and energy levels of electrons

Wave function

- Wave that corresponds to electron
- Called **orbitals** for electrons in atoms

Amplitude of wave function squared

- Can be related to **probability** of finding electron at that given point

Nodes

- Regions where electrons will not be found

Orbitals Characterized by Three Quantum Numbers:

Quantum Numbers:

- Shorthand
- Describes characteristics of electron's position
- Predicts its behavior

n = principal quantum number

- All orbitals with same n are in same shell

ℓ = secondary quantum number

- Divides shells into smaller groups called subshells

m_ℓ = magnetic quantum number

- Divides subshells into individual orbitals

$n =$ Principal Quantum Number

- Allowed values: positive integers from 1 to ∞
 - $n = 1, 2, 3, 4, 5, \dots \infty$
- Determines:
 - Size of orbital
 - Total energy of orbital
- $R_H hc = 2.18 \times 10^{-18}$ J/atom
- For given atom,
 - Lower $n =$ Lower (more negative) E
= More stable

$$E = -\frac{Z^2 R_H hc}{n^2}$$

ℓ = Orbital Angular Momentum Quantum Number

- Allowed values: 0, 1, 2, 3, 4, 5...($n - 1$)
- Letters: s, p, d, f, g, h

Orbital designation

number $\longrightarrow n\ell \longleftarrow$ letter

- Possible values of ℓ depend on n
- n different values of ℓ for given n
- Determines
 - Shape of orbital

m_ℓ = Magnetic Quantum Number

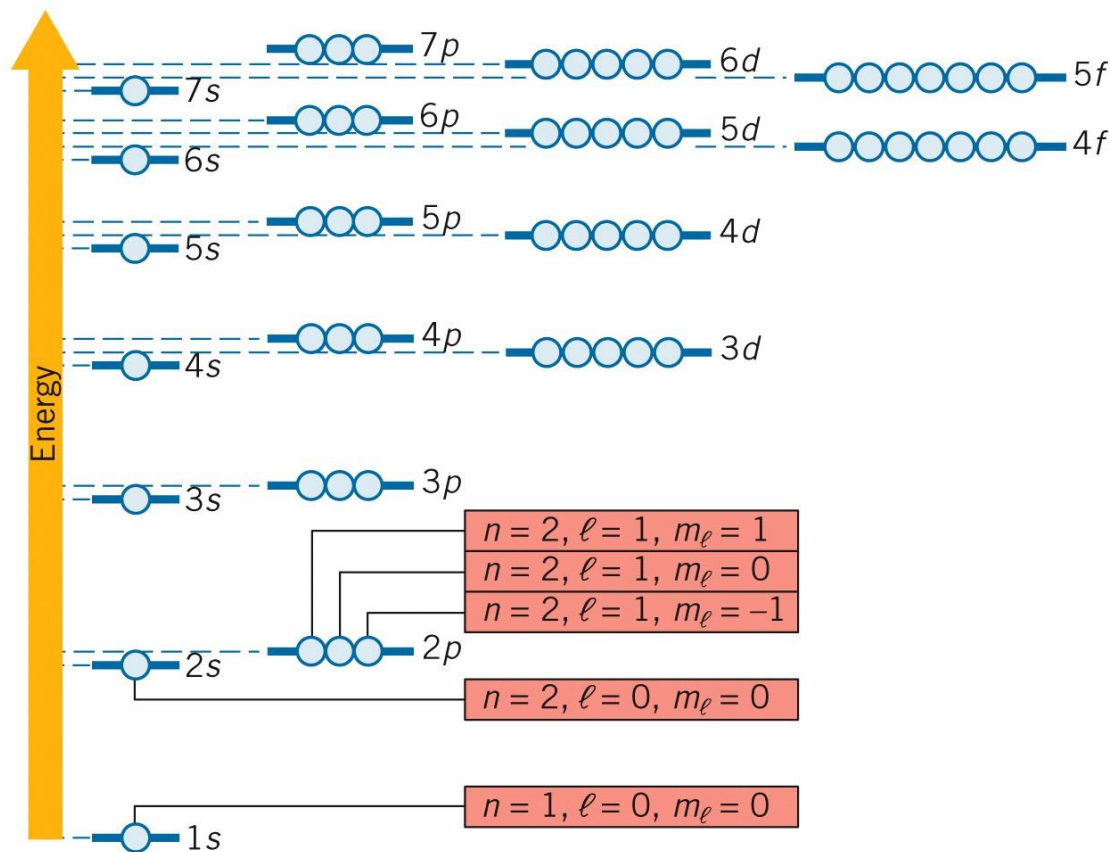
- Allowed values: from $-\ell$ to 0 to $+\ell$
 - **Example:** when $\ell = 2$ then m_ℓ can be
 - $-2, -1, 0, +1, +2$
- Possible values of m_ℓ depend on ℓ
 - There are $2\ell+1$ different values of m_ℓ for given ℓ
- Determines orientation of orbital in space
- To designate specific orbital, you need three quantum numbers
 - n, ℓ, m_ℓ

Summary of Relationships Among the Quantum Numbers n , ℓ , and m_ℓ

TABLE 7.1 Summary of Relationships among the Quantum Numbers n , ℓ , m_ℓ

Value of n	Value of ℓ	Value of m_ℓ	Subshell	Number of Orbitals
1	0	0	1s	1
2	0	0	2s	1
	1	-1, 0, 1	2p	3
3	0	0	3s	1
	1	-1, 0, 1	3p	3
	2	-2, -1, 0, 1, 2	3d	5
4	0	0	4s	1
	1	-1, 0, 1	4p	3
	2	-2, -1, 0, 1, 2	4d	5
	3	-3, -2, -1, 0, 1, 2, 3,	4f	7

Orbitals of Many Electrons

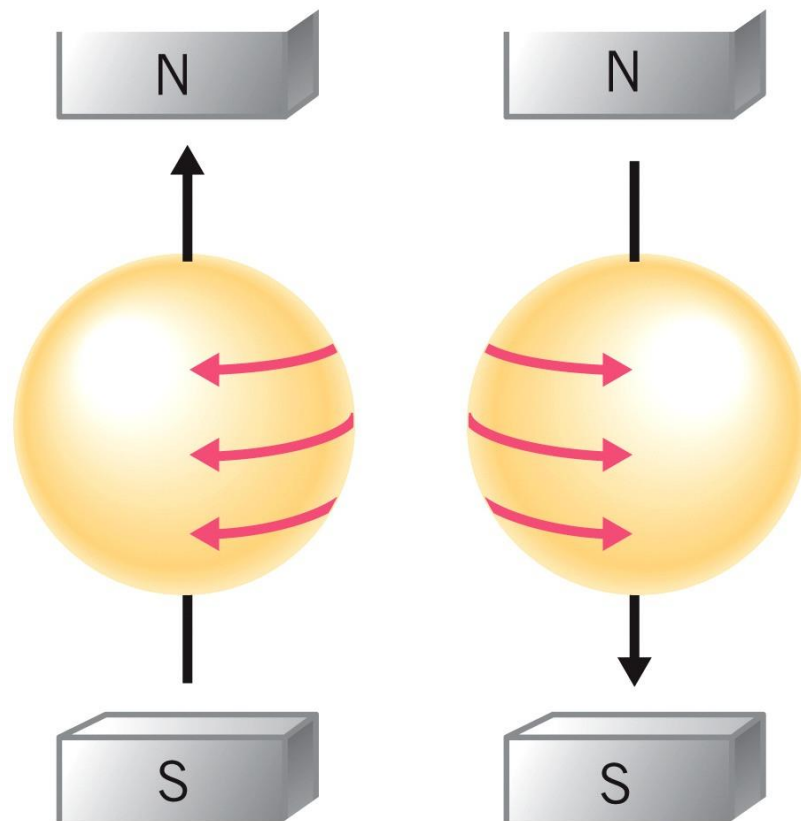


Orbital Designation

- Based on first two quantum numbers
- Number for n and letter for ℓ
- How many electrons can go in each orbital?
 - Two electrons
 - Need another quantum number

Spin Quantum Number, m_s

- Arises out of behavior of electron in magnetic field
- electron acts like a top
- Spinning charge is like a magnet
 - Electron behave like tiny magnets
- Leads to two possible directions of electron spin
 - Up and down
 - North and south



Possible Values:

$+1/2$

$-1/2$

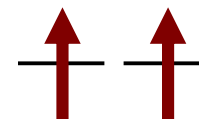
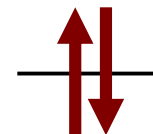


Pauli Exclusion Principle

- No two electrons in same atom can have same set of all four quantum numbers (n, ℓ, m_ℓ, m_s)
- **Can only have two electrons per orbital**
- Two electrons in same orbital must have opposite spin
 - Electrons are said to be paired

Know from Magnetic Properties

- Two electrons in same orbital have different spins
 - Spins paired—**diamagnetic**
 - Sample not attracted to magnetic field
 - Magnetic effects tend to cancel each other
- Two electrons in different orbital with same spin
 - Spins unpaired—**paramagnetic**
 - Sample attracted to a magnetic field
 - More unpaired electrons lead to a stronger attraction to a magnet
- Measure extent of attraction
 - Gives number of unpaired spins



Number of Orbitals and Electrons in the Orbitals

Subshell	Number of Orbitals	Maximum Number of Electrons
<i>s</i>	1	2
<i>p</i>	3	6
<i>d</i>	5	10
<i>f</i>	7	14

Shell	Subshells	Maximum Shell Population
1	1 <i>s</i>	2
2	1 <i>s</i> 2 <i>p</i>	8 (2 + 6)
3	3 <i>s</i> 3 <i>p</i> 3 <i>d</i>	18 (2 + 6 + 10)
4	4 <i>s</i> 4 <i>p</i> 4 <i>d</i> 4 <i>f</i>	32 (2 + 6 + 10 + 14)

Your Turn!

Which of the following is a valid set of four quantum numbers (n, ℓ, m_ℓ, m_s) ?

- A. 3, 2, 3, $+1/2$
- B. 3, 2, 1, 0
- C. 3, 0, 0, $-1/2$
- D. 3, 3, 0, $+1/2$
- E. 0, -1, 0, $-1/2$

Your Turn!

What is the maximum number of electrons allowed in a set of **4p** orbitals?

- A. 14
- B. 6
- C. 0
- D. 2
- E. 10

Your Turn!

Determine the number of electrons in an atom that can have a principal quantum number of 3, and list the possible subshells they occupy.

- A. 18 ($3s$, $3p$, $3d$)
- B. 8 ($3s$, $3p$)
- C. 2 ($3s$)
- D. 32 ($3s$, $3p$, $3d$, $3f$)

Ground State Electron Arrangements

Electron Configurations

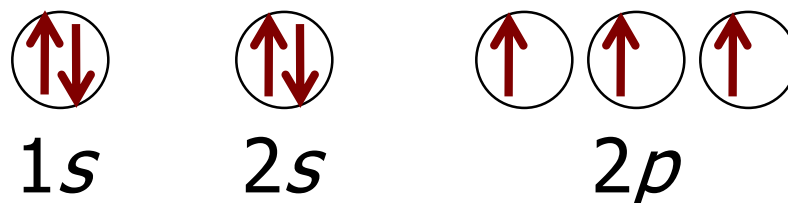
- Distribution of electrons among orbitals of atom
 1. List subshells that contain electrons
 2. Indicate their electron population with superscript

e.g., N is $1s^2 2s^2 2p^3$

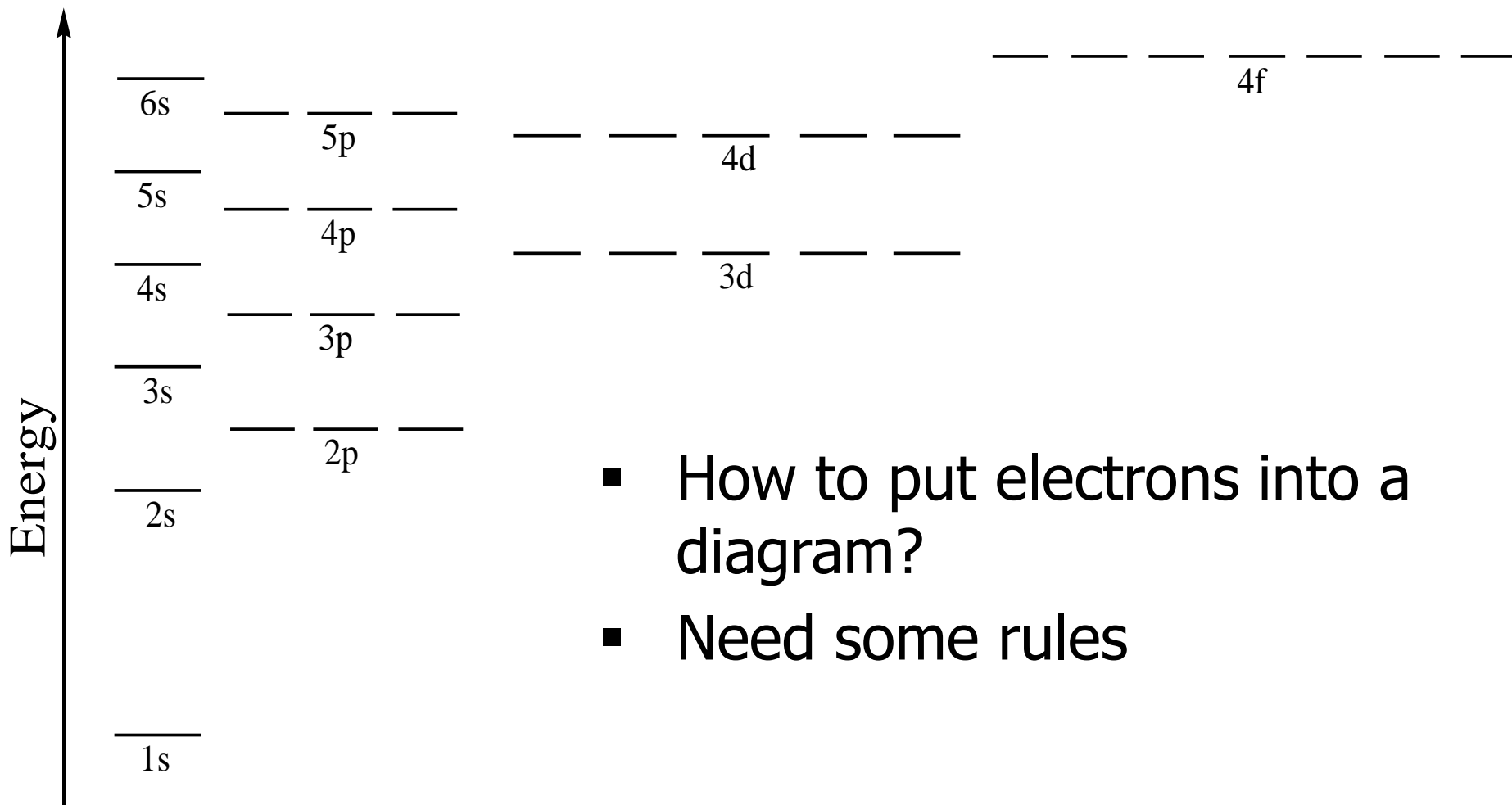
Orbital Diagrams

- Way to represent electrons in orbitals
 1. Represent each orbital with circle (or line)
 2. Use arrows to indicate spin of each electron

e.g., N is



Energy Level Diagram for Multi Electron Atom/Ion



- How to put electrons into a diagram?
- Need some rules

Aufbau Principle

- Building-up principle

Pauli Exclusion Principle

- Two electrons per orbital
- Fill following the order suggested by the periodic table
- Spins must be paired

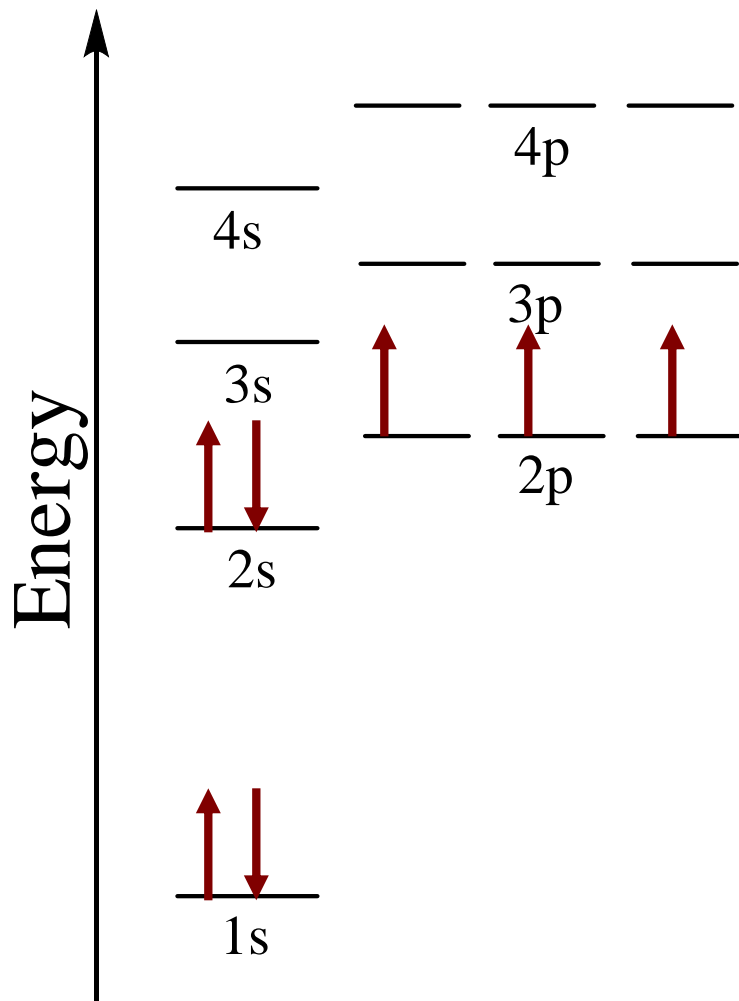
Hund's Rule

- If you have more than one orbital all at the same energy
 - Put one electron into each orbital with spins parallel (all up) until all are half filled
 - After orbitals are half full, pair up electrons

Why?

- Repulsion of electrons in same region of space
- Empirical observation based on magnetic properties

Orbital Diagram and Electron Configurations: e.g., N, Z = 7

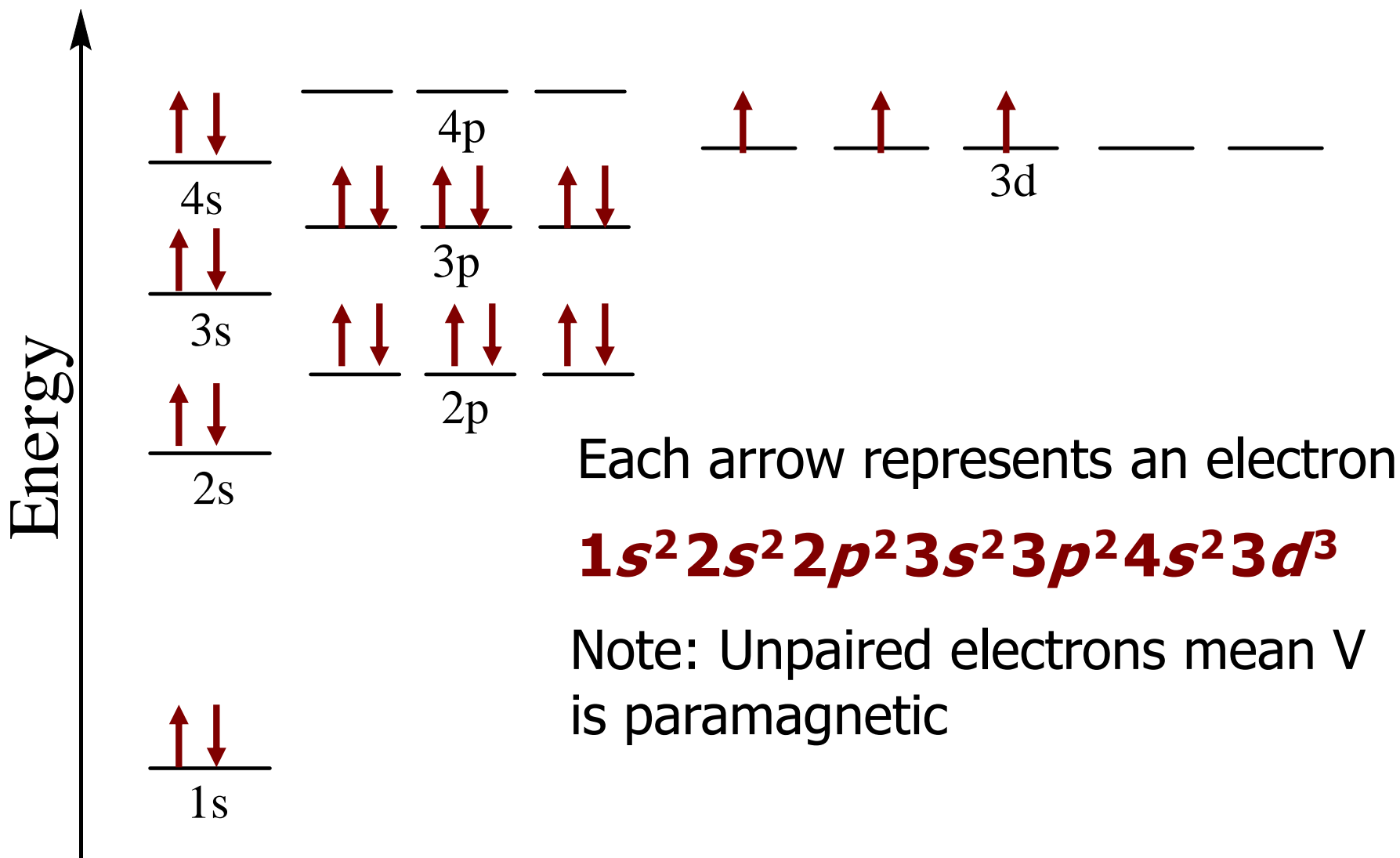


Each arrow represents electron



Note: Unpaired electrons mean N is paramagnetic

Orbital Diagram and Electron Configurations: e.g., V, Z = 23

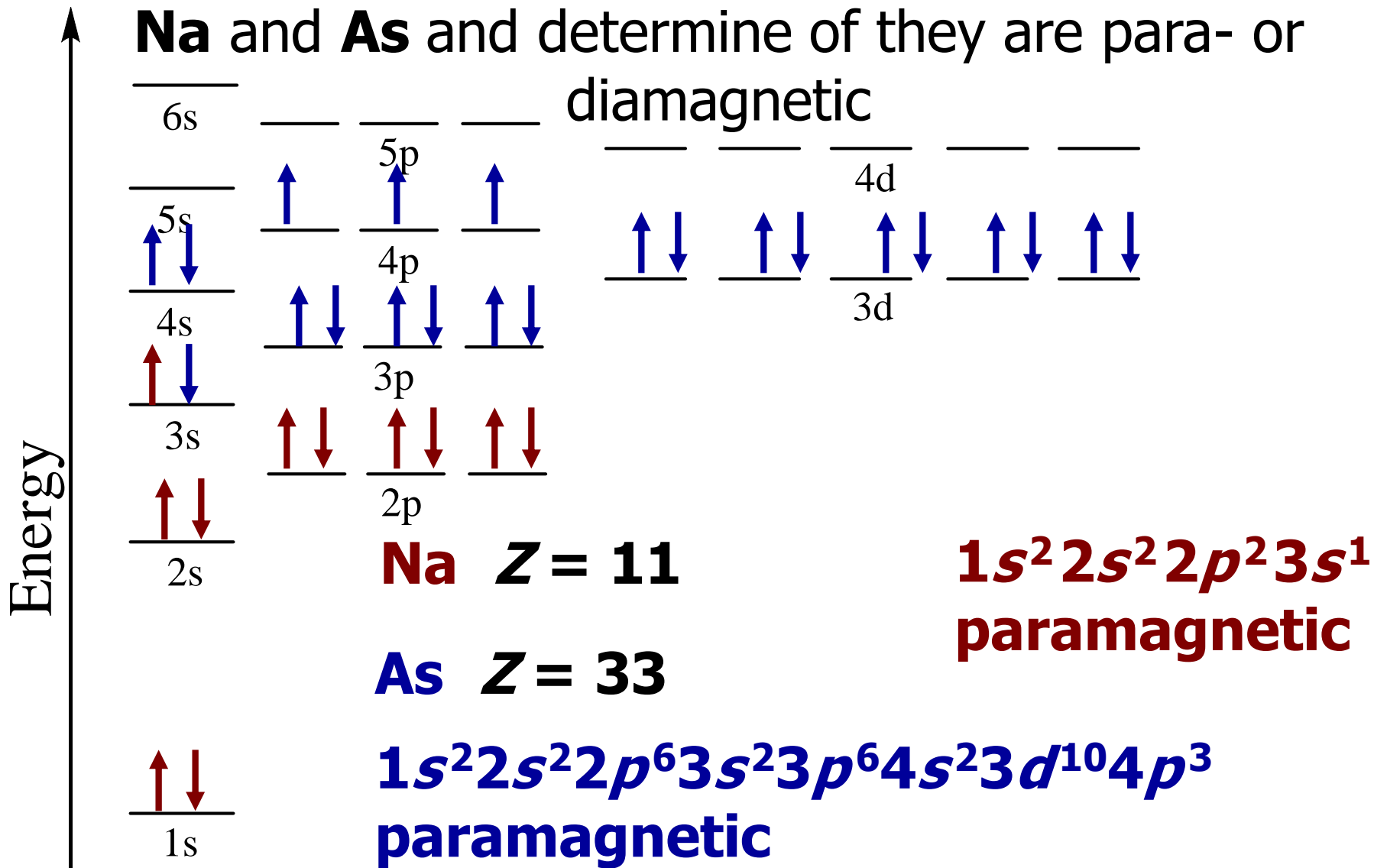


Learning Check

Give electron configurations and orbital diagrams for

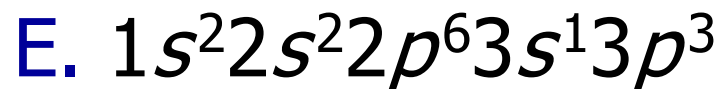
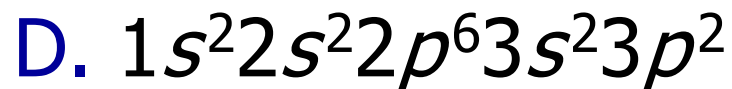
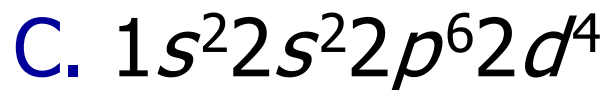
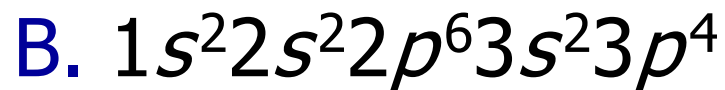
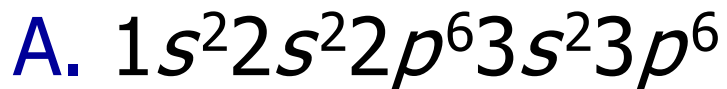
Na and **As** and determine if they are para- or

diamagnetic



Your Turn!

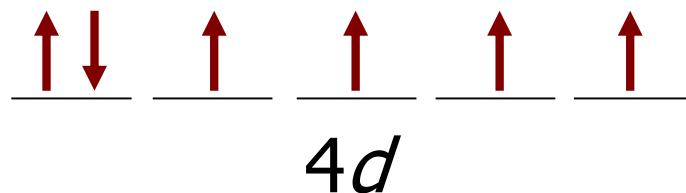
What is the correct ground state electron configuration for Si?



Your Turn!

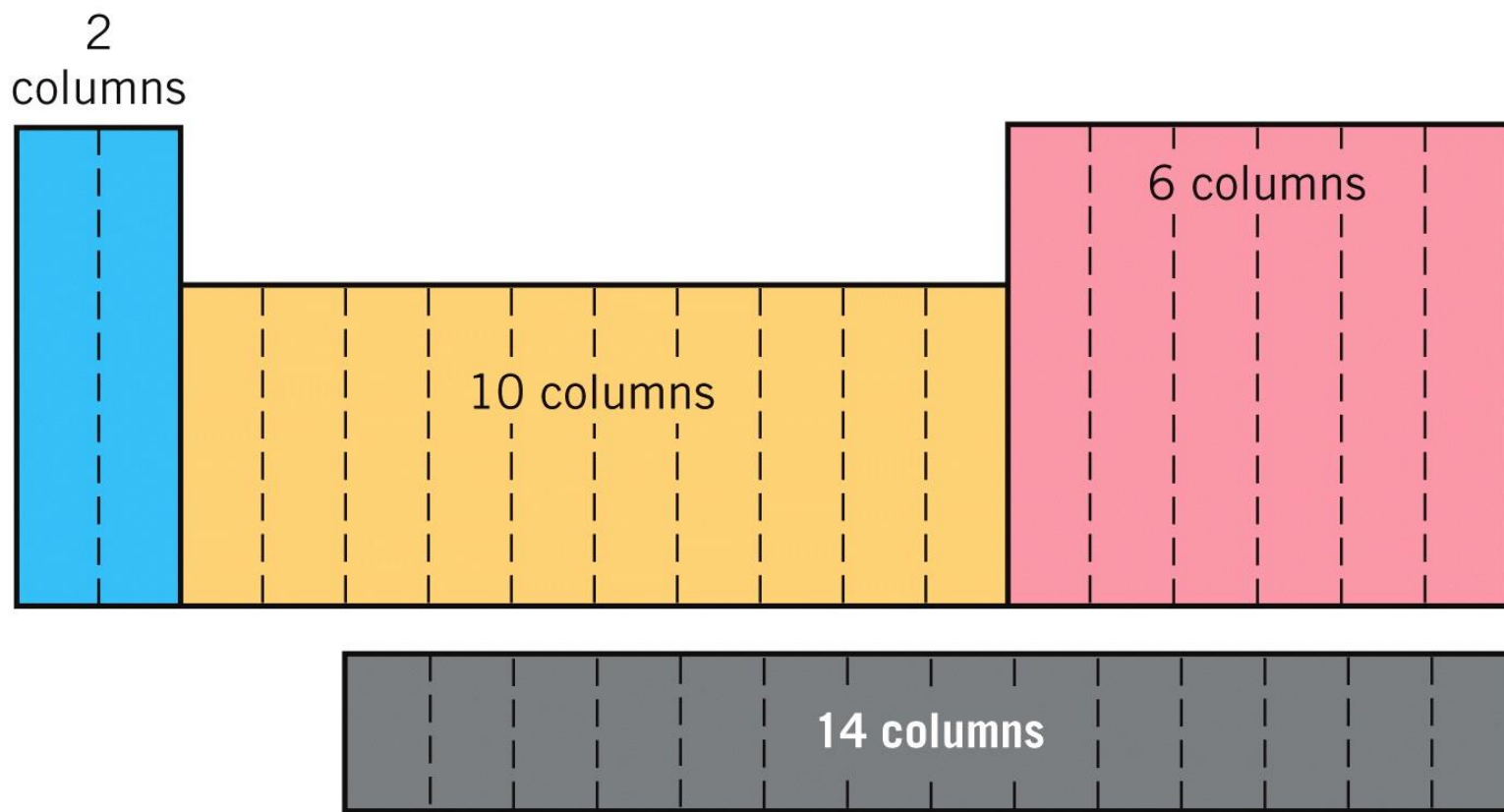
Is ruthenium diamagnetic or paramagnetic, and if paramagnetic, how many unpaired electrons does it have?

- A. paramagnetic, 4
- B. paramagnetic, 6
- C. diamagnetic, 0
- D. diamagnetic, 4
- E. diamagnetic, 6



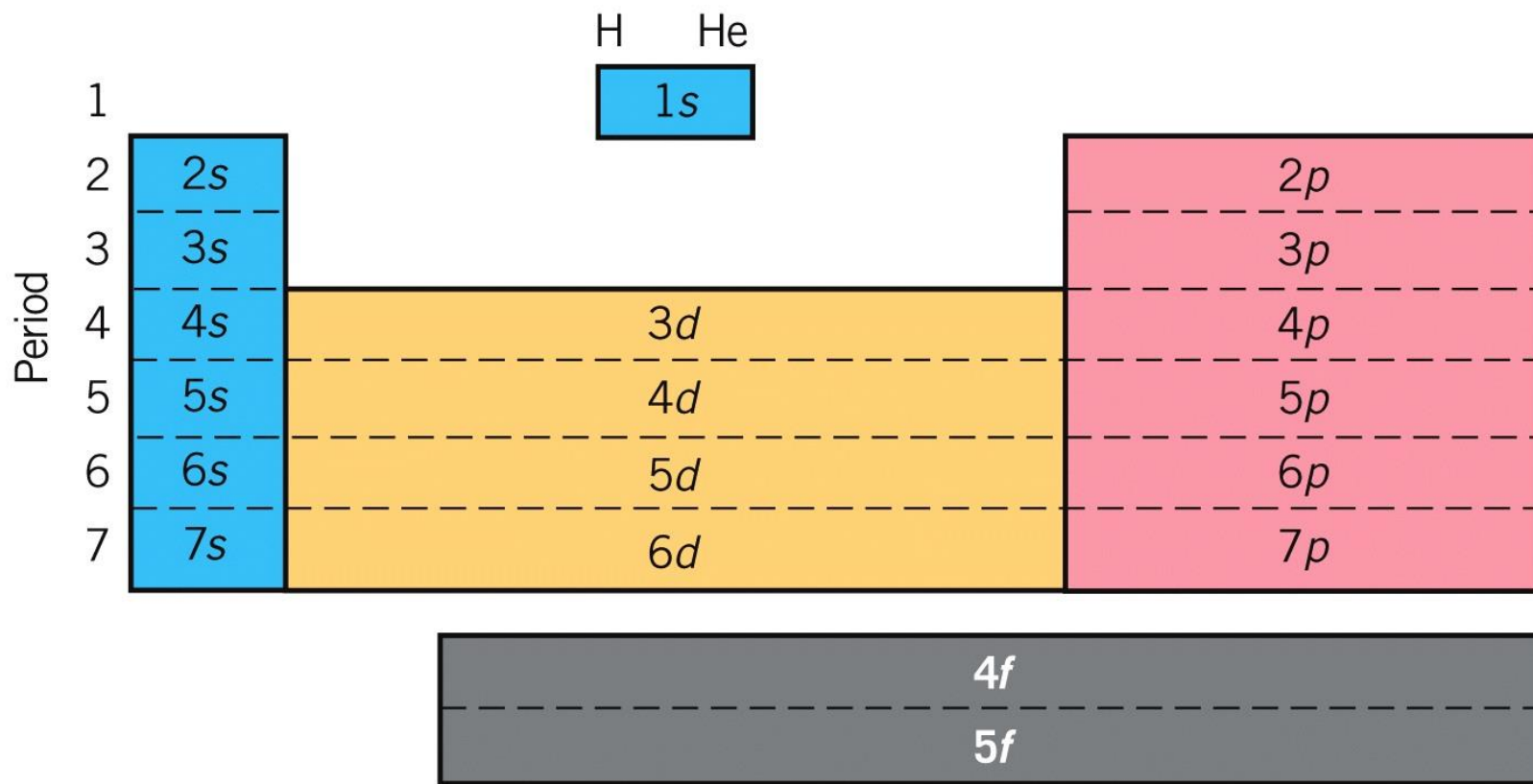
Periodic Table

- Divided into regions of 2, 6, 10, and 14 columns
- This equals maximum number of electrons in s , p , d , and f sublevels



Sublevels and the Periodic Table ^{7.8}

- Each row (period) represents different energy level
- Each region of chart represents different type of sublevel



Now Ready to Put Electrons into Atoms ^{7.8}

Electron configurations must be consistent with:

Pauli Exclusion principle

- Two electrons per orbital, spins opposite

Aufbau principle

- Start at lowest energy orbital
- Fill, then move up

Hund's rule

- One electron in each orbital of same energy, spins parallel
- Only pair up if have to

Read Periodic Table to Determine Electron Configuration – He

- Read from left to right
- First electron goes into period 1
- First type of sublevel to fill = “ $1s$ ”
- He has 2 two electrons
- electron configuration for He is: $1s^2$

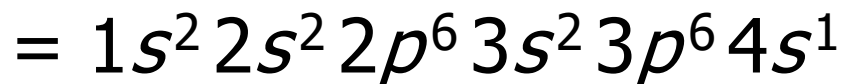
$n=1$	1 H																				2 He	
$n=2$	3 Li	4 Be																				
$n=3$	11 Na	12 Mg																				
$n=4$	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni												
$n=5$	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd												
$n=6$	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt												
$n=7$	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds												

“ ns ” orbital being filled
 “ np ” orbital being filled
 “ $(n-1)d$ ” orbital being filled
 “ $(n-2)f$ ” orbital being filled

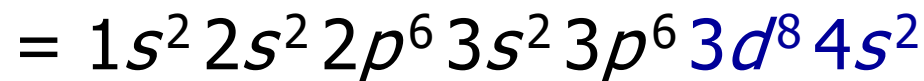
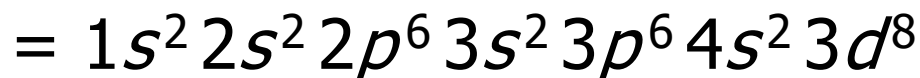
Learning Check

Write the correct ground state electron configuration for each of the following elements. List in order of increasing n and within each shell, increasing ℓ .

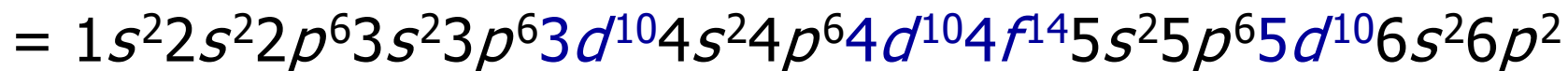
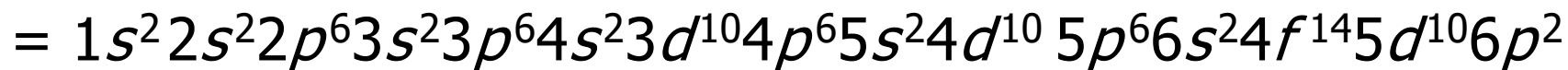
1. K Z = 19



2. Ni Z = 28



3. Pb Z = 82

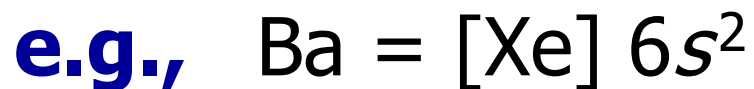


Chemical Reactivity

- Periodic table arranged by chemical reactivity
 - Depends on outer shell electrons (highest n)
 - Each row is different n
- Core electrons
 - Inner electrons are those with $n < n_{\max}$
 - Buried deep in atom

Abbreviated Electron Configurations^{7.8} - Noble Gas Notation

- [noble gas of previous row] and electrons filled in next row
- Represents core + outer shell electrons
- Use to emphasize that only outer shell electrons react



Noble Gas Core Notation for Mn

- Find last noble gas that is filled before Mn
- Next fill sublevels that follow



n=1	1 H																			2 He
n=2	3 Li	4 Be										5 B	6 C	7 N	8 O	9 F	10 Ne			
n=3	11 Na	12 Mg										13 Al	14 Si	15 P	16 S	17 Cl	18 Ar			
n=4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr		
n=5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe		
n=6	55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn		
n=7	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg									

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Your Turn!

The ground state electron configuration for Ca is:

- A. $[\text{Ar}] 3s^1$
- B. $1s^2 2s^2 2p^6 3s^2 3p^5 4s^2$
- C. $[\text{Ar}] 4s^2$
- D. $[\text{Kr}] 4s^1$
- E. $[\text{Kr}] 4s^2$

Look at Group 2A

	Z	Electron Configuration	Abbrev
Be	4	$1s^2 2s^2$	[He] $2s^2$
Mg	12	$1s^2 2s^2 2p^6 3s^2$	[Ne] $3s^2$
Ca	20	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	[Ar] $4s^2$
Sr	38	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$	[Kr] $5s^2$
Ba	56	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2$	[Xe] $6s^2$
Ra	88	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^{14} 5s^2 5p^6$ $5d^{10} 6s^2 6p^6 7s^2$	[Rn] $7s^2$

- All have ns^2 outer shell electrons
- Only difference is value of n

Your Turn!

An element with the electron configuration $[\text{Xe}]4f^{14}5d^76s^2$ would belong to which class (e.g., halogen, alkali metal, etc.) on the periodic table?

- A. transition elements
- B. alkaline earth elements
- C. halogens
- D. lanthanide elements
- E. alkali metals

Valence Shell Electron Configurations

- One last type of electron configuration
- Use with representative elements (s and p block elements) – longer columns
- Here only electrons in outer shell important for bonding
- Only electrons in s and p subshells
- **Valence shell** = outer shell
= occupied shell with highest n
- Result – use even more abbreviated notation for electron configurations
- $\text{Sn} = 5s^25p^2$

Electronic Configurations

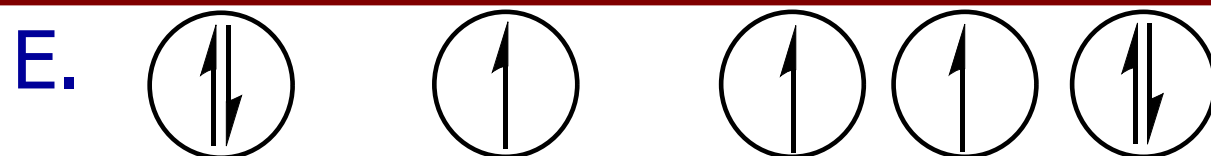
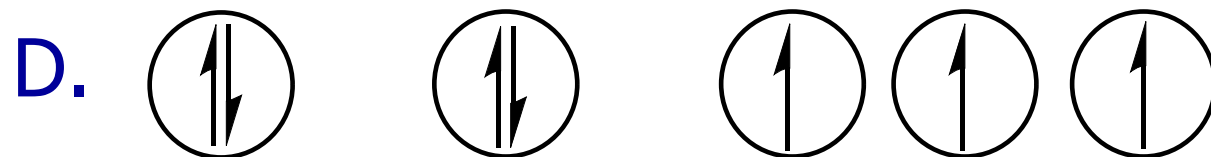
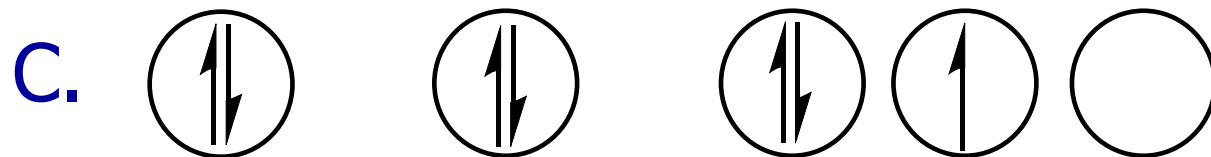
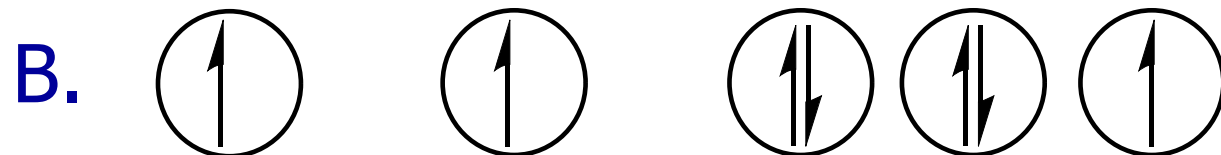
- A few exceptions to rules

Element	Expected	Experimental
Cr	[Ar] 3d⁴4s²	[Ar] 3d ⁵ 4s ¹
Cu	[Ar] 3d⁹4s²	[Ar] 3d ¹⁰ 4s ¹
Ag	[Kr] 4d⁹5s²	[Kr] 4d ¹⁰ 5s ¹
Au	[Xe] 5d⁹6s²	[Xe] 5d ¹⁰ 6s ¹

- Exactly filled and exactly half-filled subshells have extra stability
- Promote one electron into *ns* orbital to gain this extra stability

Your Turn!

The orbital diagram corresponding to the ground state electron configuration for **nitrogen** is:



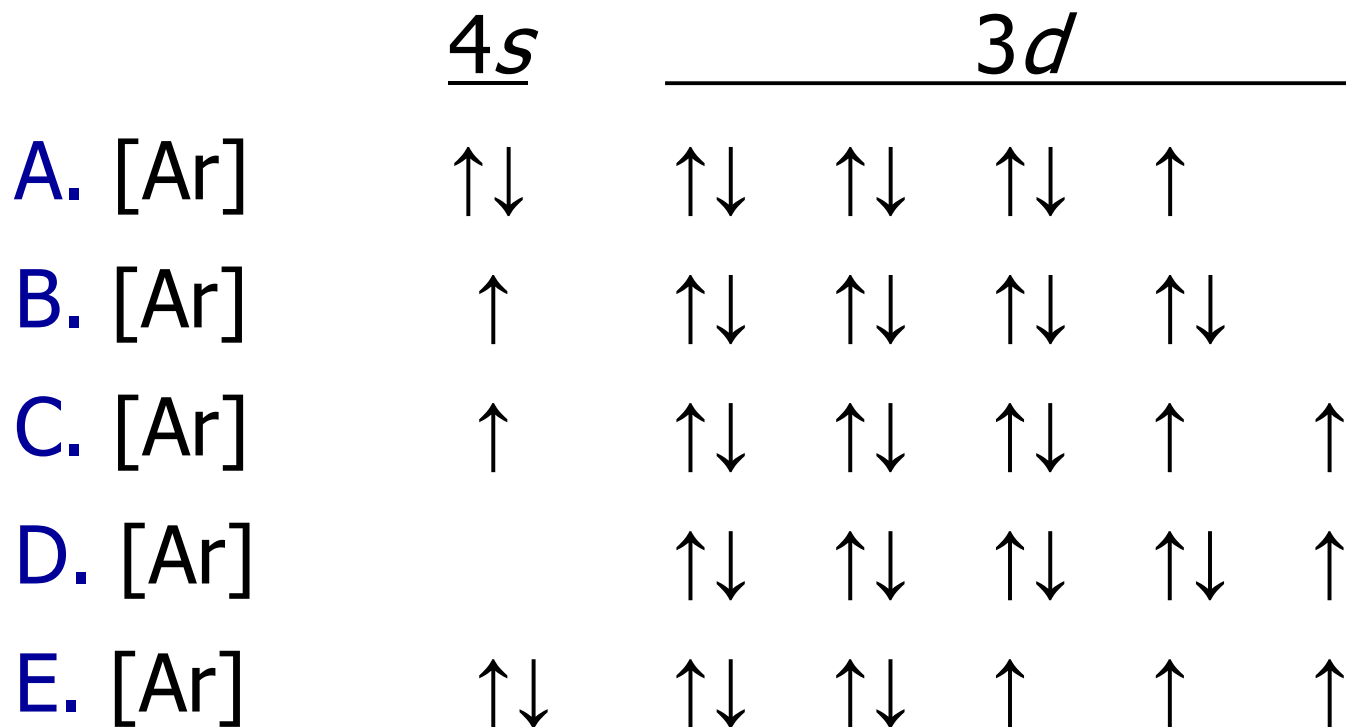
1s

2s

2p

Your Turn!

Which of the following choices is the correct orbital diagram for a cobalt atom?



Heisenberg's Uncertainty Principle^{7.9}

- Can't know both exact position and exact speed of subatomic particle simultaneously
- Such measurements always have certain minimum uncertainty associated with them

Heisenberg's Uncertainty Principle^{7.9}

Macroscopic scale

- Errors in measurements much smaller than measured value

Subatomic scale

- Errors in measurements equal to or greater than measured value
 - If you know **position** exactly, know nothing about **velocity**
 - If you know **velocity** exactly, know nothing about **position**

Consequence of Heisenberg's Uncertainty Principle

- Can't talk about absolute position
- Can only talk about electron probabilities
 - Where is e^- likely to be?
- ψ = wavefunction
 - Amplitude of electron wave
- ψ^2 = probability of finding electron at given location
- Probability of finding an electron in given region of space equals the square of the amplitude of wave at that point

Electron Cloud

Electron dot picture = snapshots

- Lots of dots shown by large amplitude of wave function
 - High probability of finding electrons

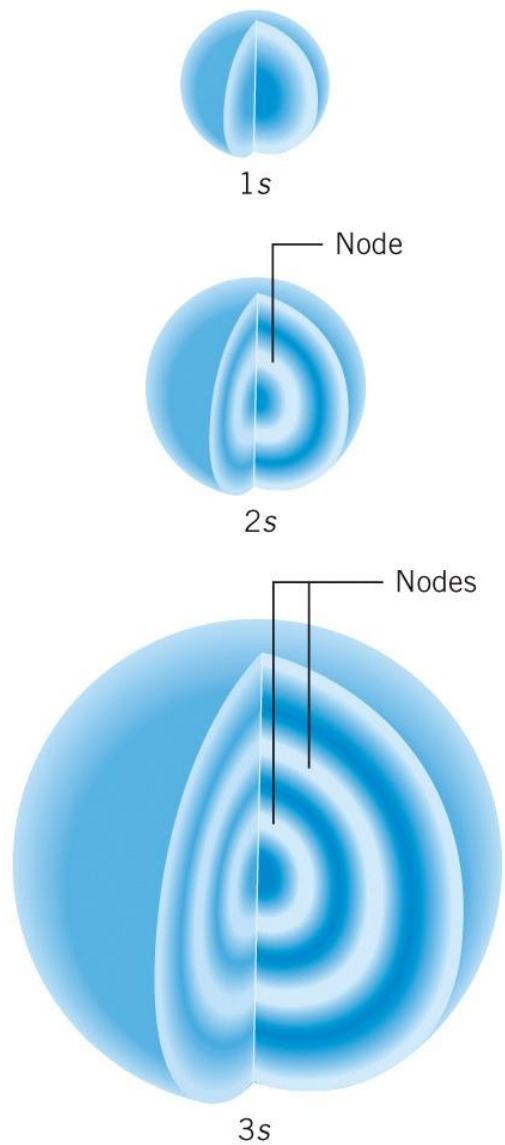
Electron density

- How much of electrons charge packed into given volume
- **High** Probability
 - **High** electron density or **Large** electron density
- **Low** Probability
 - **Low** electron density or **Small** electron density

Electron Density Distribution

- Determined by
 - Shape l
 - Size n
 - Orientation m_l
- Electron density
 - No sharp boundary
 - Gradually fades away
- “Shape”
 - Imaginary surface enclosing 90% of electron density of orbital
 - Probability of finding electrons is same everywhere on surface

Effect of n on s Orbital^{7,9}



- In any given direction probability of finding electron same
- All s orbitals are **spherically** shaped
- **Size** increases as n increases

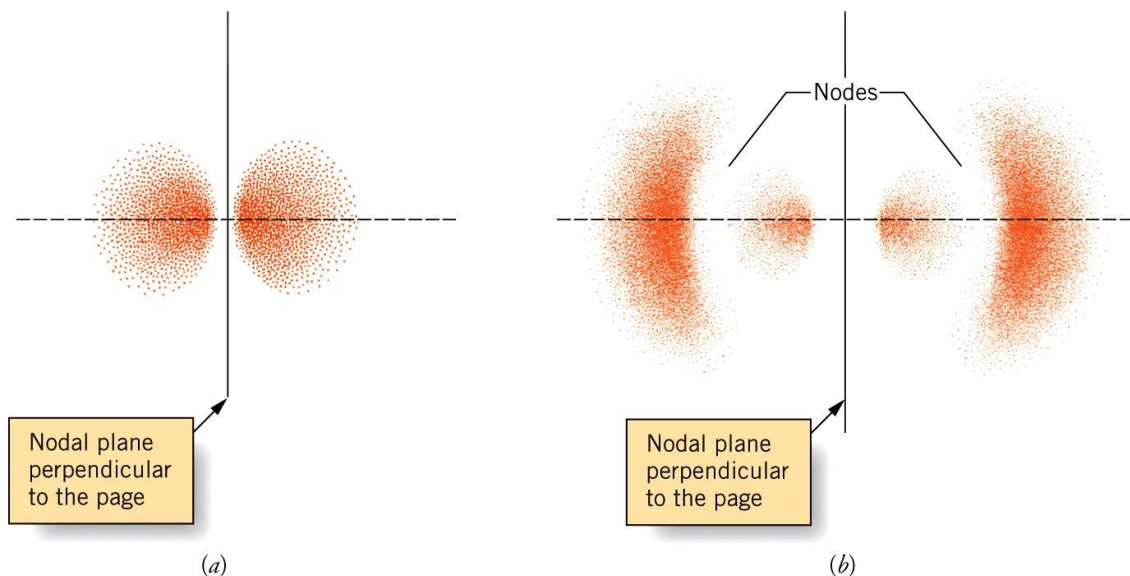
Spherical Nodes

- At higher n , now have spherical nodes
 - Spherical regions of zero probability, inside orbital
- Node for electron wave
 - Imaginary surface where electron density = 0
 - $2s$, one spherical node, size larger
 - $3s$, two spherical nodes, size larger yet

In general:

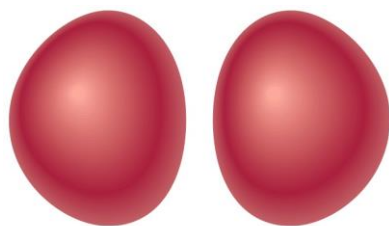
- Number of spherical nodes
= $n - 1$

p Orbitals

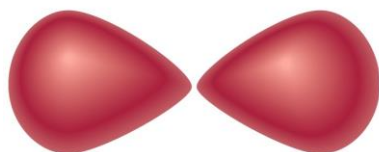


- Possess one **nodal plane** through nucleus
 - Electron density only on two sides of nucleus
 - Two lobes of electron density
- All p orbitals have same overall shape
 - **Size** increases as n increases
 - Note that the **$3p$** orbital has an additional node

Representations of p Orbitals

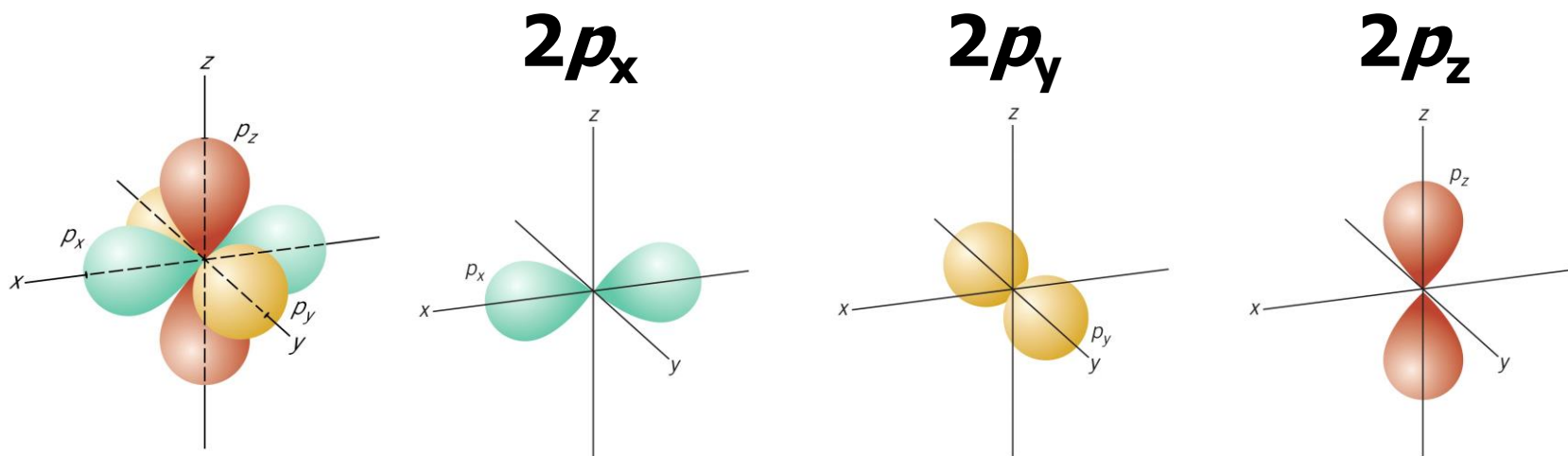


(a)

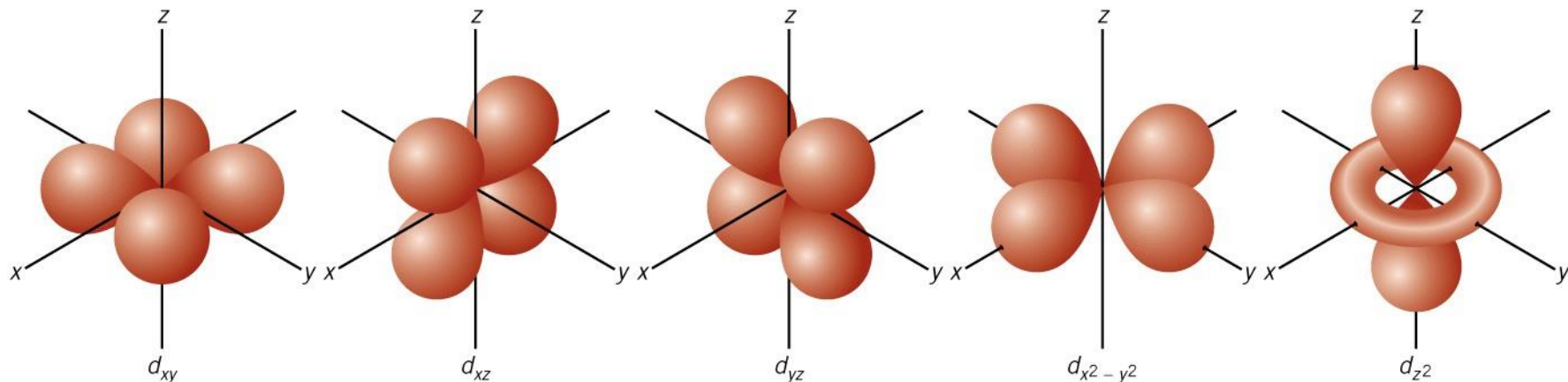


(b)

- Constant probability surface for $2p$ orbital
- Simplified p orbital emphasizing directional nature of orbital
- All $2p$ orbitals in p sub shell
 - One points along each axis



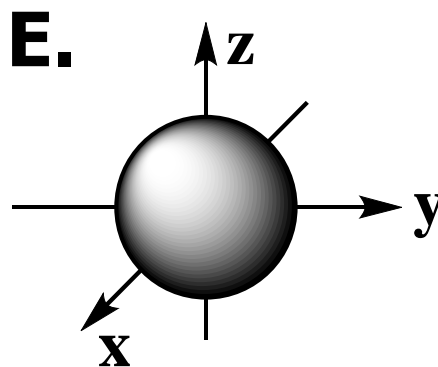
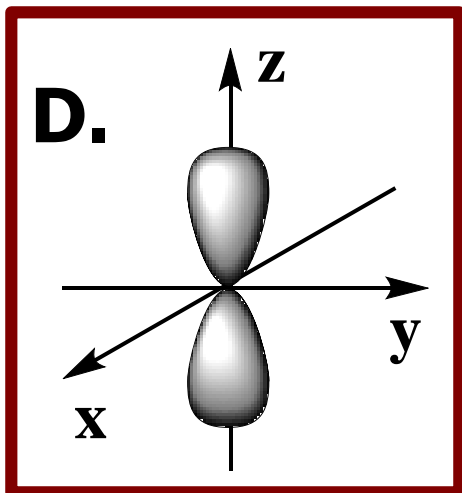
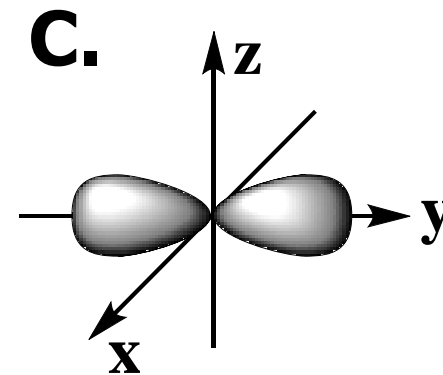
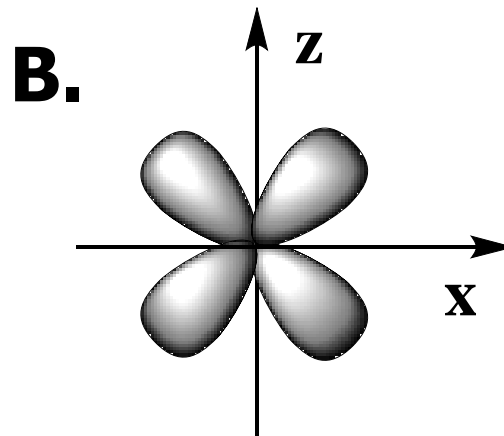
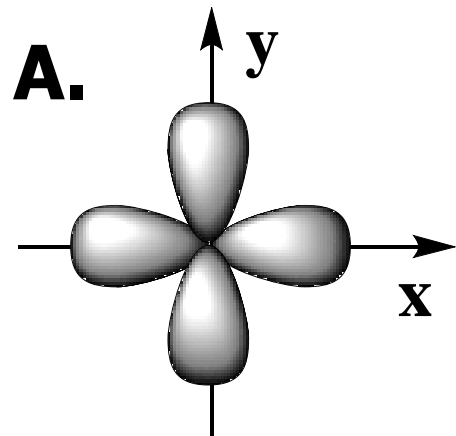
There Are Five Different *d* Orbitals^{7.9}



- Four with four lobes of electron density
- One with two lobes and ring of electron density
- Result of two nodal planes through nucleus
- Number of nodal planes through nucleus = l

Your Turn!

Which sketch represents a p_z orbital?



Your Turn!

Compared to electrons with lower principal quantum numbers, electrons with higher principal quantum numbers:

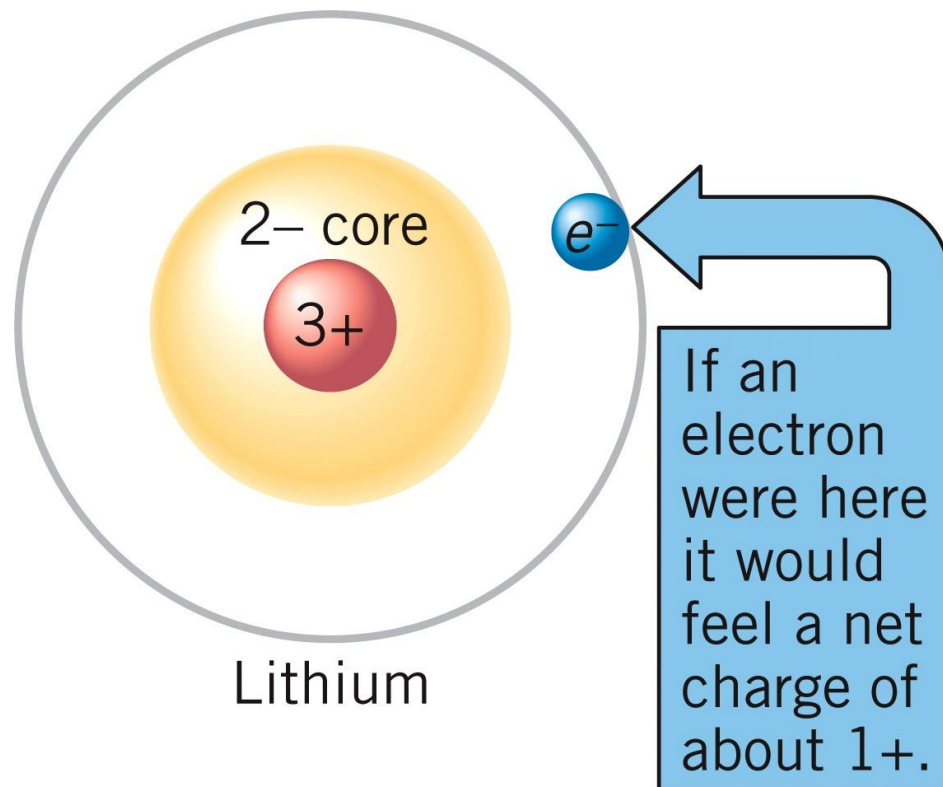
- A. all exist is orbitals that do not have lobes
- B. move in circular orbits around the nucleus
- C. have fewer subshells that they can occupy
- D. have lower energy because they are closer to the nucleus
- E. have higher energy because they are further away from the nucleus

Periodic Properties: Consequences of Electron Configuration

- Chemical and physical properties of elements
 - Vary **systematically** with position in periodic table
 - i.e., with element's electron configuration
- To explain, must first consider amount of positive charge felt by outer electrons (valence electrons)
 - Core electrons spend most of their time closer to nucleus than valence (outer shell) electrons
 - Shield or cancel out (screen out, neutralize) some of positive charge of nucleus

Learning check: Li $1s^2 2s^1$

- Three protons in nucleus
- Two core electrons in close ($1s$)
- Net positive charge felt by outer electron
 - Approximately one proton



Effective Nuclear Charge (Z_{eff})

- Net positive charge outer **electron** feels
- Core **electrons** shield valence **electrons** from full nuclear charge

Shielding

- Electrons in same subshell don't shield each other
 - Same average distance from nucleus
 - Trying to stay away from each other
 - Spend very little time one below another
- Effective nuclear charge determined primarily by
 - Difference between charge on nucleus (Z) and charge on core (number of inner electrons)

Your Turn!

What value is the closest estimate of Z_{eff} for a valence electron of the calcium atom?

A. 1

B. 2 $Z_{\text{eff}} = 20 \text{ protons} - 18 \text{ core electrons} = 2$

C. 6

D. 20

E. 40

Atomic Size

- Theory suggests sizes of atoms and ions indistinct
- Experiment shows atoms/ions behave as if they have definite size
 - C and H have \sim same distance between nuclei in large number of compounds

Atomic Radius (r)

- Half of distance between two like atoms
 - H—H C—C etc.
 - Usually use units of picometer
 - $1 \text{ pm} = 1 \times 10^{-12} \text{ m}$
 - Range 37 – 270 pm for atoms

Trends in Atomic Radius (r)

Increases down Column (group)

- Z_{eff} essentially constant
- n increases, outer electrons farther away from nucleus and radius increase

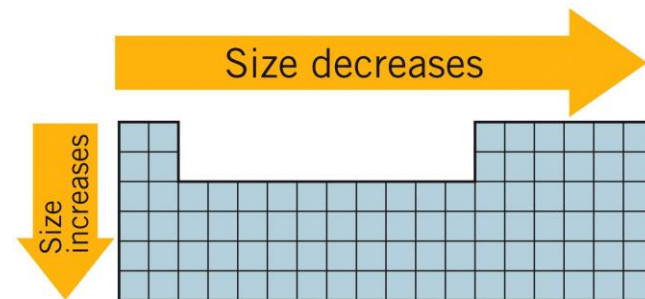
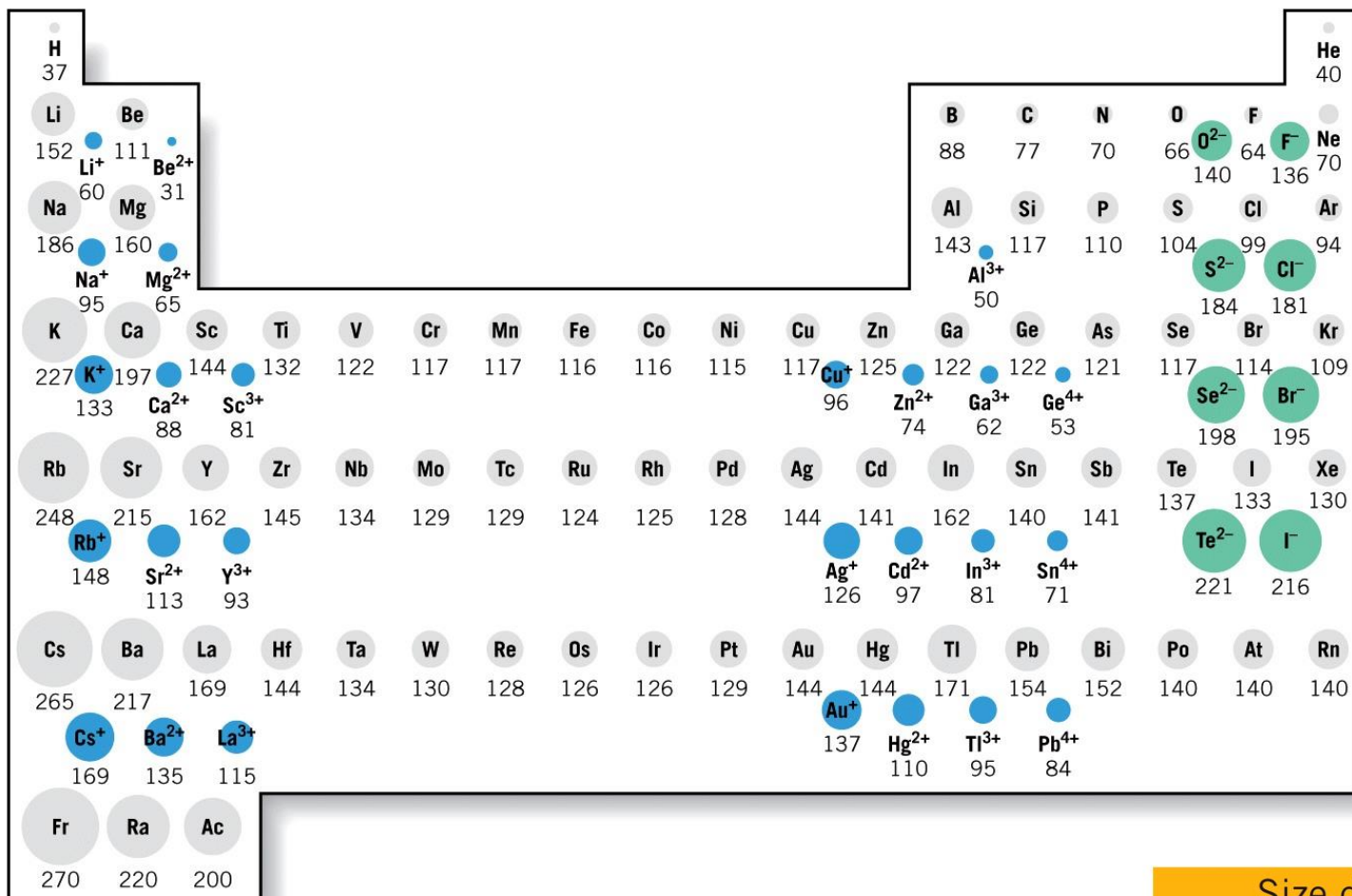
Decreases across row (period)

- n constant
- Z_{eff} increases, outer electrons feel larger Z_{eff} and radius decreases

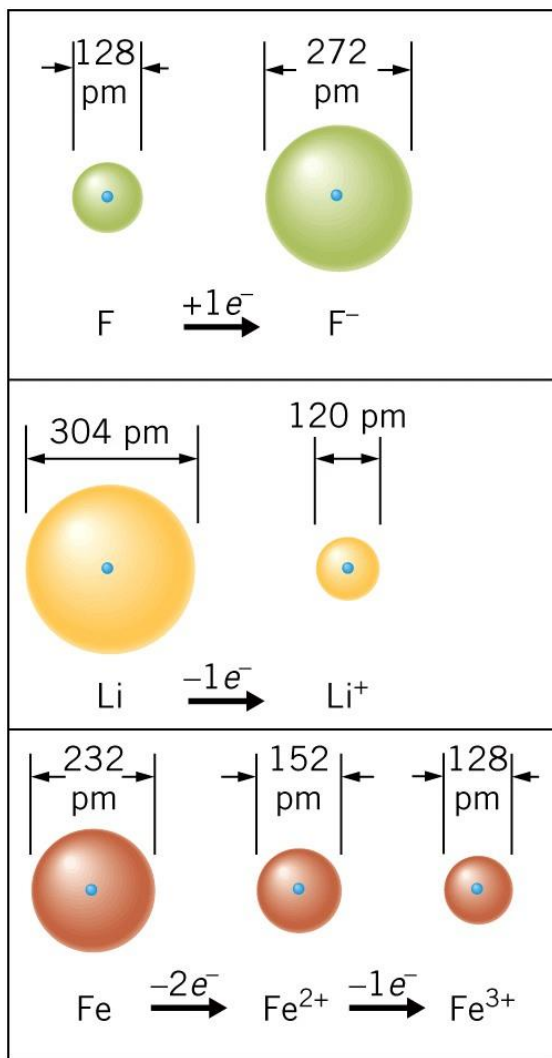
Transition Metals and Inner Transition Metals

- Size variations less pronounced as filling core
- n same (outer electrons) across row
- Decrease in Z_{eff} and r more gradually

Atomic and Ionic Radii (in pm)



Ionic Radii



- Increases down column (group)
 - Decreases across row (period)
- Anions** *larger* than parent atom
- Same Z_{eff} , more electrons
 - Radius expands
- Cations** *smaller* than parent atom
- Same Z_{eff} , less electrons,
 - Radius contracts

Your Turn!

Which of the following has the **smallest** radius?

A. Ar

B. K^+

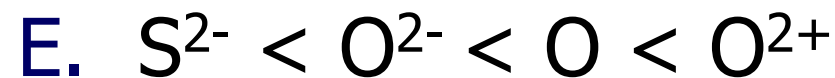
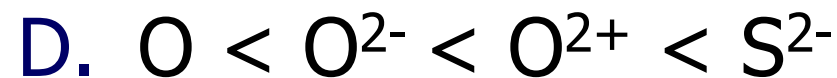
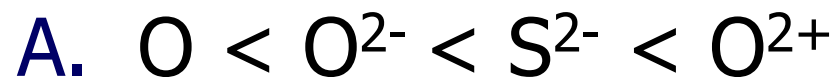
C. Cl^-

D. Ca^{2+}

E. S^{2-}

Your Turn!

Arrange the following from smallest to largest
radius: O , O^{2-} , S^{2-} , O^{2+}



Ionization Energy

- Energy required to remove electron from gas phase atom
- Corresponds to taking electron from n to $n = \infty$
- First ionization energy $M(g) \longrightarrow M^+(g) + e^-$

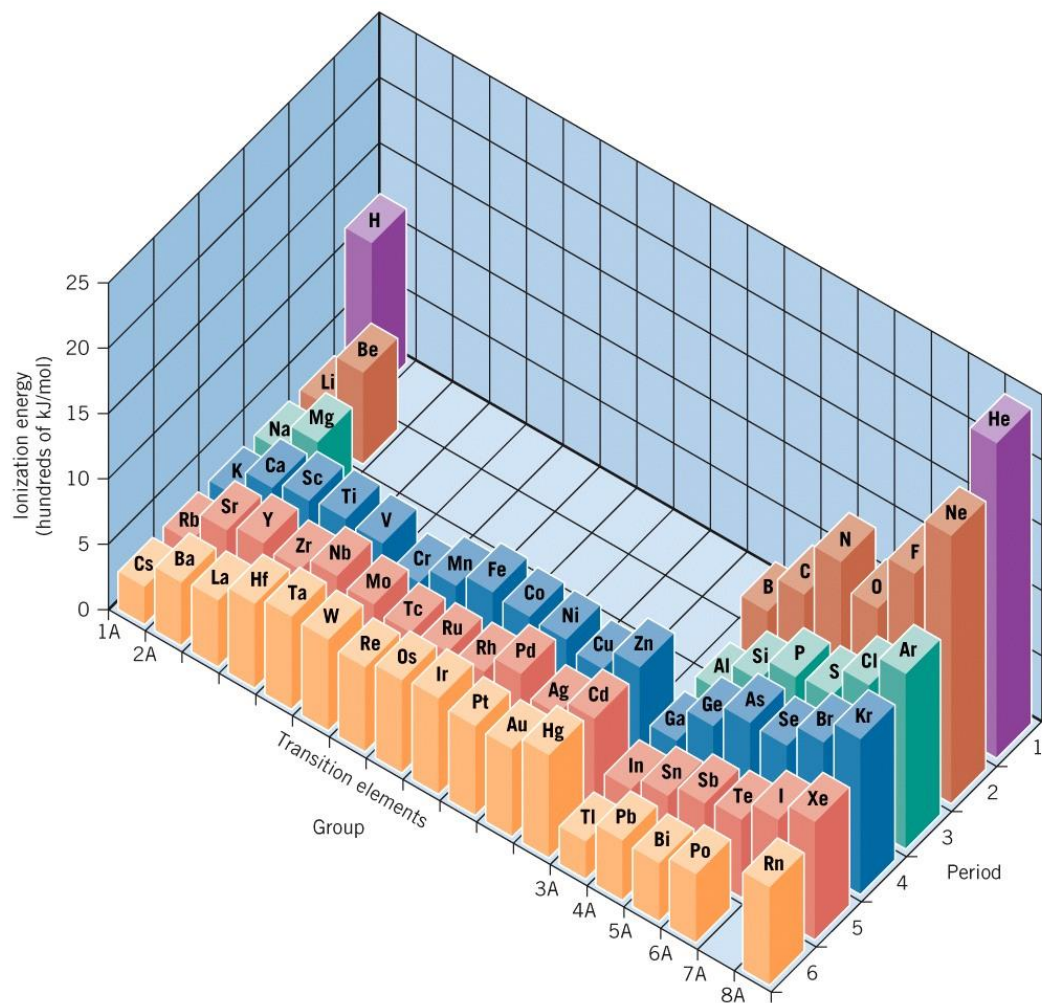
- $IE = \Delta E$

$$IE = \frac{R_{\text{H}} hc Z_{\text{eff}}^2}{n^2}$$

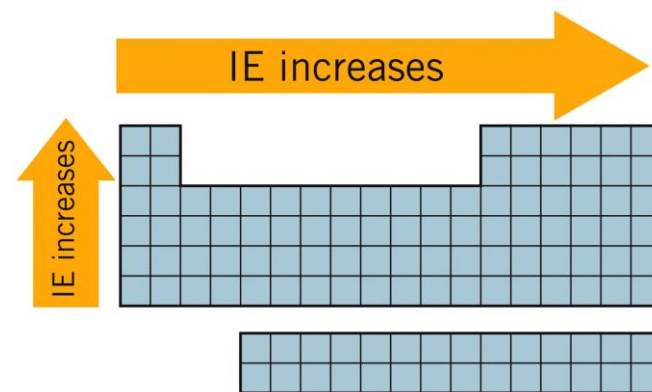
Trends:

- Ionization energy decreases down column (group) as n increases
- Ionization energy increases across row (period) as Z_{eff} increases

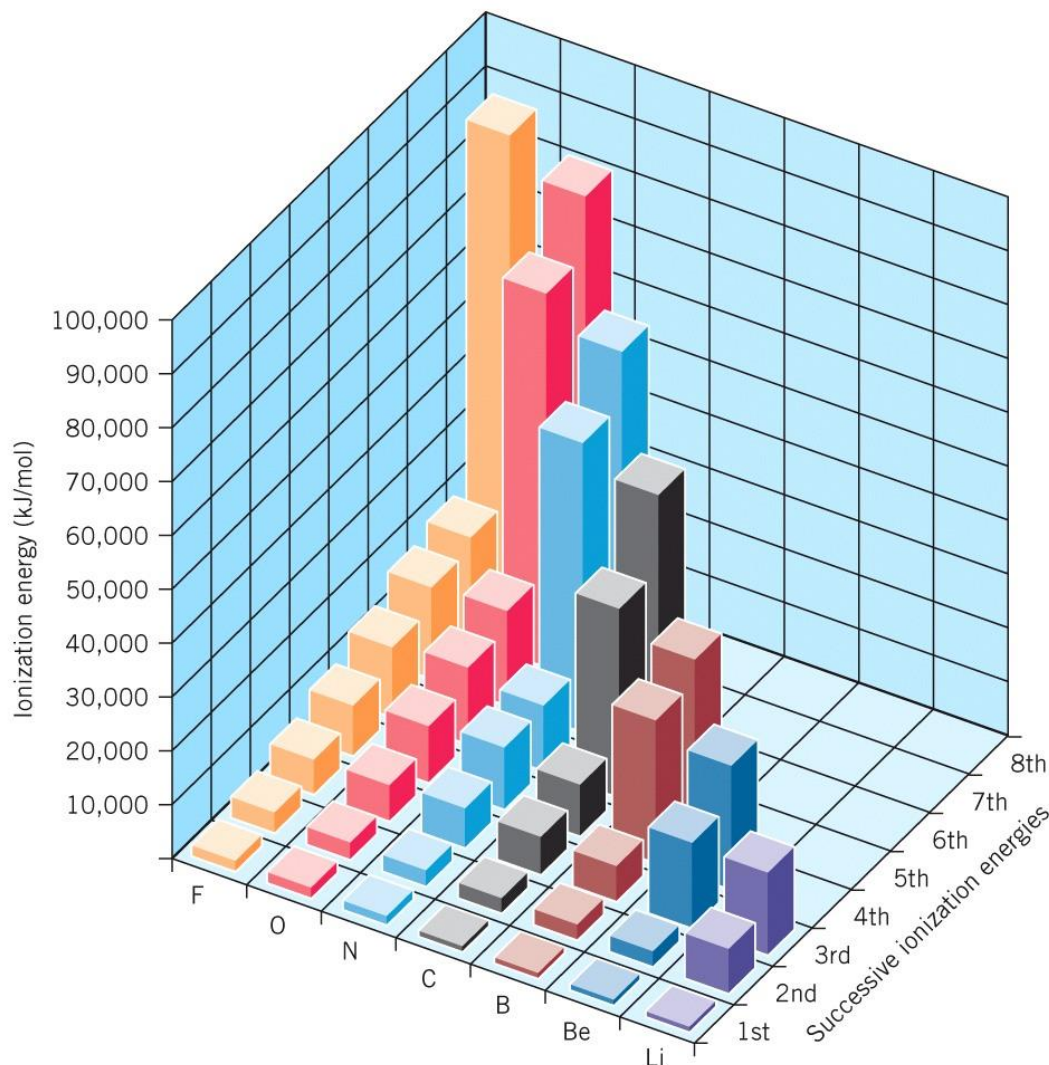
Comparing First Ionization Energies



- Largest first ionization energies are in upper right
- Smallest first ionization energies are in lower left



Successive Ionization Energies



- Increases slowly as remove each successive electron
- See big increase in ionization energy
 - When break into exactly filled or half filled subshell

Successive Ionization Energies in kJ/mol for H through Mg ^{7.10}

TABLE 7.2

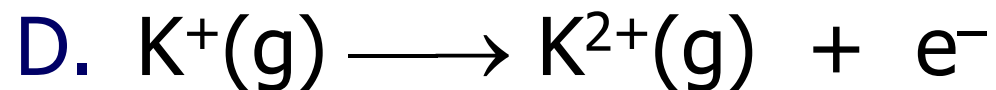
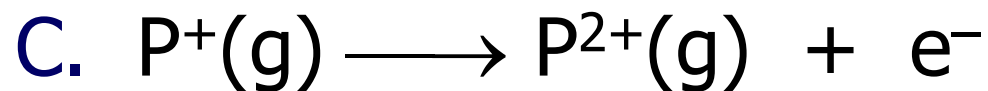
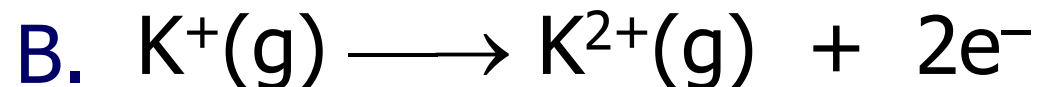
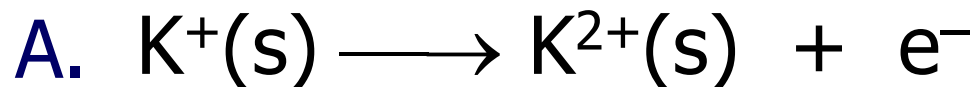
Successive Ionization Energies in kJ/mol for Hydrogen through Magnesium

	1st	2nd	3rd	4th	5th	6th	7th	8th
H	1312							
He	2372	5250						
Li	520	7297	11,810					
Be	899	1757	14,845	21,000				
B	800	2426	3659	25,020	32,820			
C	1086	2352	4619	6221	37,820	47,260		
N	1402	2855	4576	7473	9442	53,250	64,340	
O	1314	3388	5296	7467	10,987	13,320	71,320	84,070
F	1680	3375	6045	8408	11,020	15,160	17,860	92,010
Ne	2080	3963	6130	9361	12,180	15,240	–	–
Na	495	4563	6913	9541	13,350	16,600	20,113	25,666
Mg	737	1450	7731	10,545	13,627	17,995	21,700	25,662

Note the sharp increase in ionization energy when crossing the “staircase,” indicating that the last of the valence electrons has been removed.

Your Turn!

Write the chemical equation that corresponds to the second ionization energy of potassium



Your Turn!

Place the elements C, N, and O in order of increasing ionization energy.

A. $C < N < O$

B. $O < N < C$

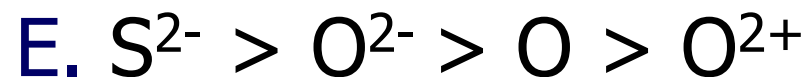
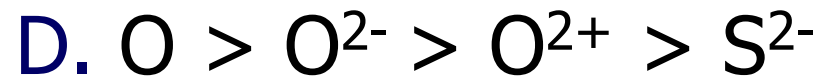
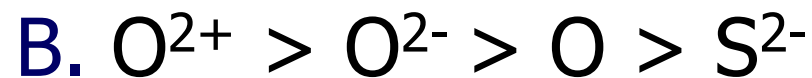
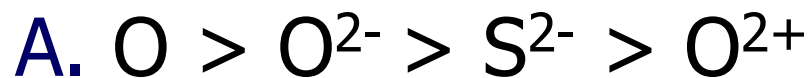
C. $C < O < N$

D. $N < O < C$

E. $N < C < O$

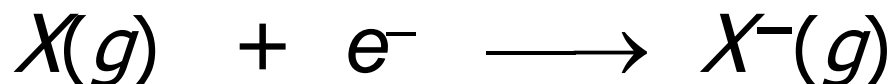
Your Turn!

Arrange the following from highest ionization energy to the lowest:



Electron Affinity (EA)

- Potential energy change associated with addition of one electron to gas phase atom or ion in the ground state



- O and F very favorable to add electrons
- Comparing first electron affinities usually negative (exothermic)
- Larger negative value means more favorable to add electron

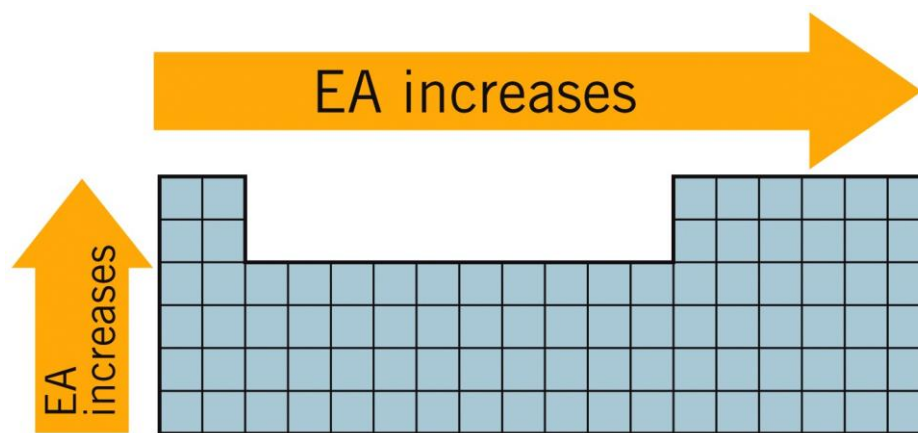
Electron Affinities of Representative Elements^{7.10}

TABLE 7.3**Electron Affinities of the Representative Elements (kJ/mol)**

1A	2A	3A	4A	5A	6A	7A
H						
-73						
Li	Be	B	C	N	O	F
-60	+238	-27	-122	~+9	-141	-328
Na	Mg	Al	Si	P	S	Cl
-53	+230	-44	-134	-72	-200	-349
K	Ca	Ga	Ge	As	Se	Br
-48	+155	-30	-120	-77	-195	-325
Rb	Sr	In	Sn	Sb	Te	I
-47	+167	-30	-121	-101	-190	-295
Cs	Ba	Tl	Pb	Bi	Po	At
-45	+50	-30	-110	-110	-183	-270

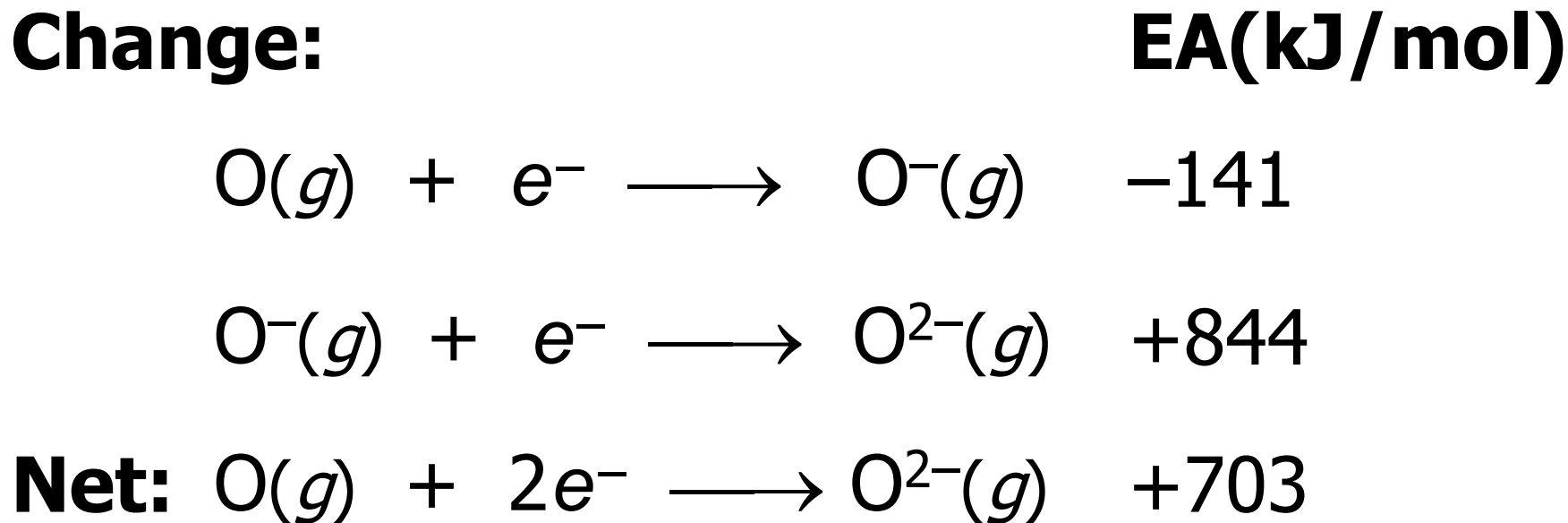
Trends in Electron Affinity (EA)

- Electron affinity becomes less exothermic down column (group) as n increases
 - Electron harder to add as orbital farther from nucleus and feels less positive charge
- Electron affinity becomes more exothermic across row (period) as Z_{eff} increases
- Easier to attract electrons as positive charge increases



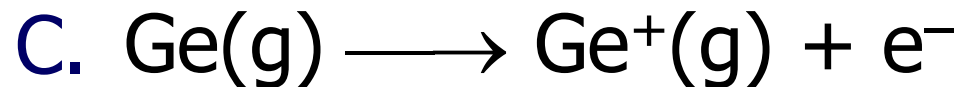
Successive Electron Affinities

- Addition of first electron – often exothermic
- Addition of more than one electron requires energy
- Consider addition of electrons to oxygen:



Your Turn!

Write the chemical equation that corresponds to the electron affinity of germanium



Your Turn!

Which of the following has the **largest** (i.e., most negative) electron affinity?

A. O

B. F

C. As

D. Cs

E. Ba

Your Turn!

Which of the following has the **smallest** electron affinity?

A. Si

B. F

C. S

D. P

E. Cl