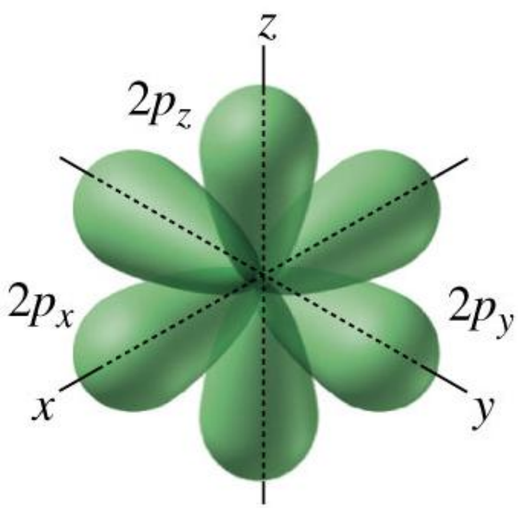
Chapter 5 Electronic Structure and Periodic Trends

5.3

Sublevels and Orbitals



A *p* sublevel consists of three *p* orbitals.

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Energy Levels

Energy levels

- are assigned quantum numbers n = 1, 2, 3, 4, and so on
- increase in energy as the value of n increases
- have a maximum number of electrons equal to $2n^2$

TABLE 5.1 Maximum Number of Electrons Allowed in Energy Levels 1–4

Energy Level (n)	1	2	3	4
$2n^2$	$2(1)^2$	$2(2)^2$	$2(3)^2$	$2(4)^2$
Maximum Number of Electrons	2	8	18	32

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Sublevels

A **sublevel**

- contains electrons with the same energy
- has the same shape but increases in volume at higher energy levels
- is found within each energy level
- is designated by the letters *s, p, d,* or *f*

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Energy of Sublevels

In any energy level

- the *s* sublevel has the lowest energy
- the *s* sublevel is followed by the *p, d, f* sublevels
- higher sublevels are possible, but only *s, p, d, f* sublevels are needed to hold the number of electrons in the atoms known today

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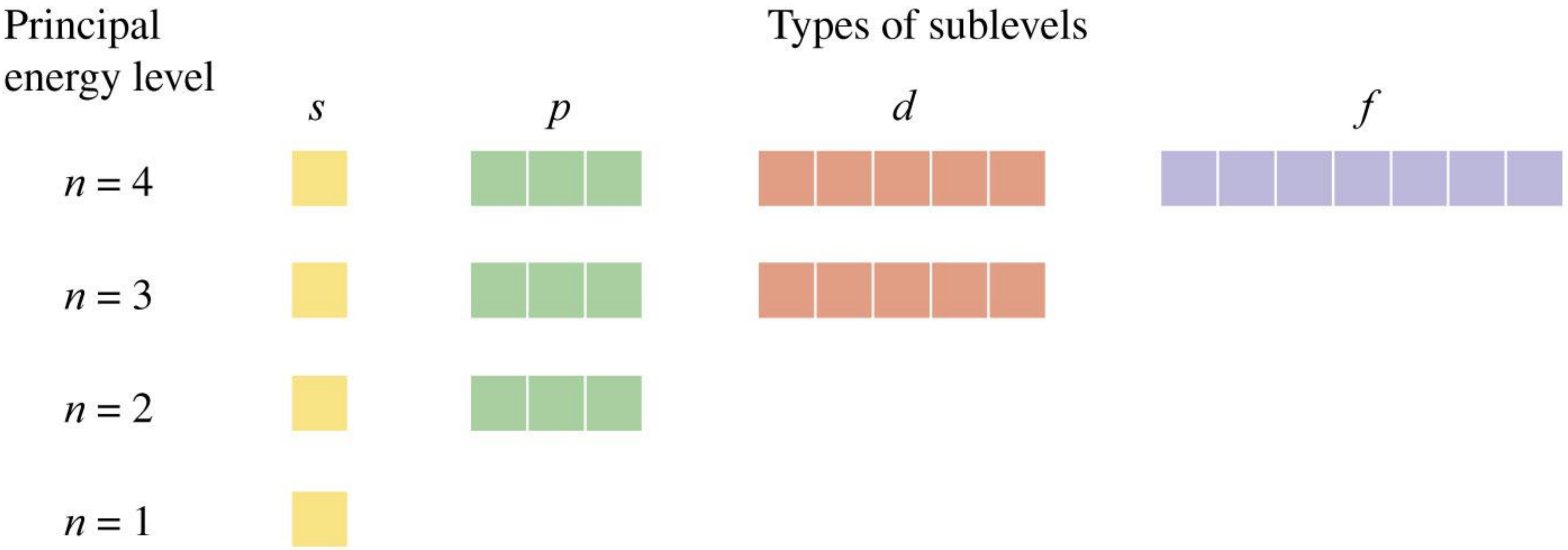
Number of Sublevels



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The number of sublevels in an energy level is the same as the principal quantum number, *n*.

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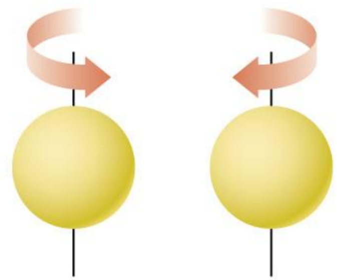
Orbitals

An **orbital**

- is a three-dimensional space around a nucleus where an electron is found most of the time
- has a shape that represents electron density (*not a path the electron follows*)
- can hold up to two electrons
- contains two electrons that spin in opposite directions



Opposite spins of electrons in an orbital

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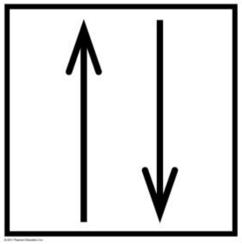

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s Orbitals

An ***s* orbital**

- has a spherical shape around the nucleus
- increases in size around the nucleus as the energy level *n* value increases
- is a single orbital found in each *s* sublevel

All *s* orbitals have spherical shapes that increase in volume at higher energy levels.

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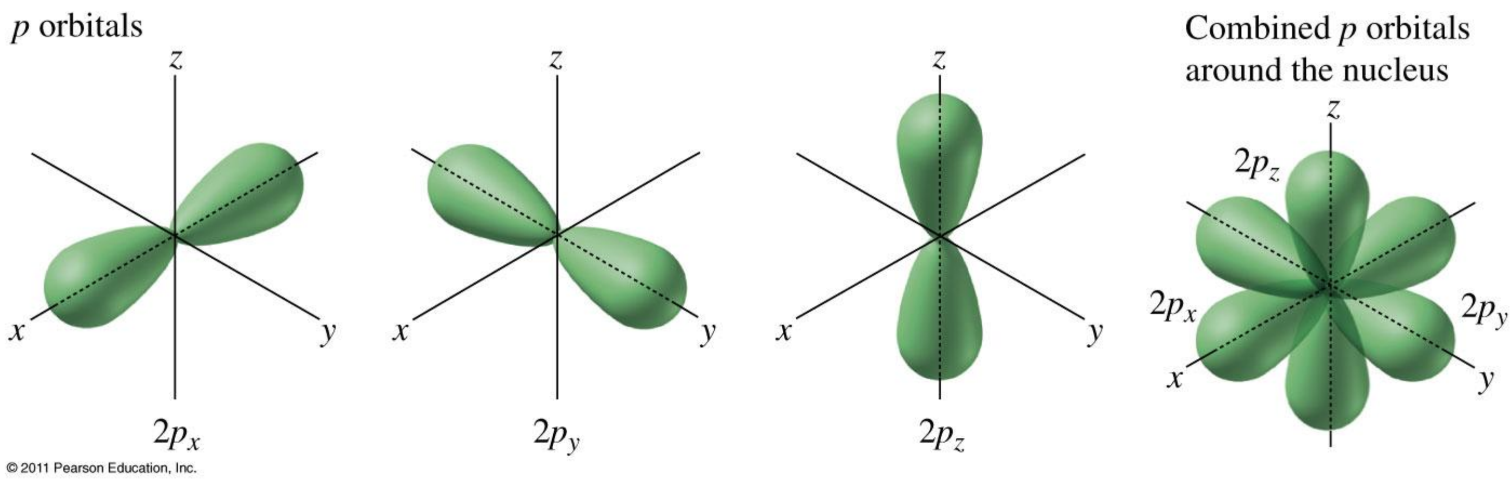
p Orbitals

A ***p* orbital**

- has a two-lobed shape
- is one of three *p* orbitals that make up each *p* sublevel, each aligned along a different axis
- increases in size as the value of *n* increases



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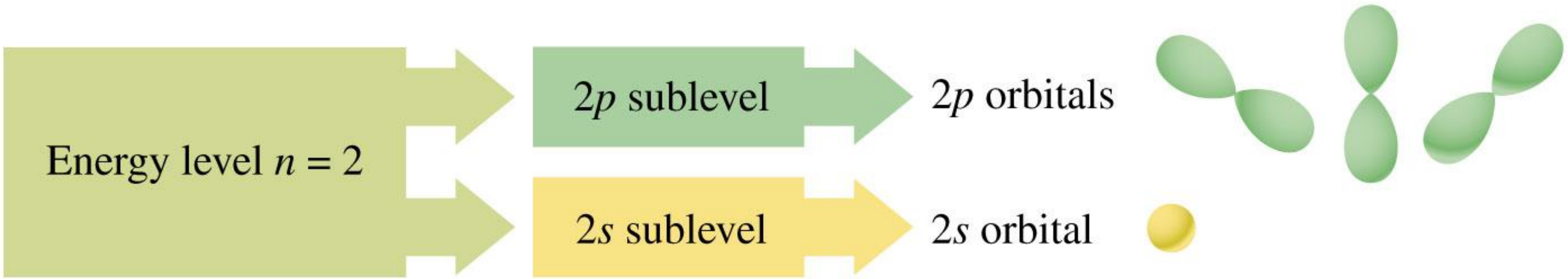
Sublevels and Orbitals

Each **sublevel** consists of a specific number of orbitals.

- an *s* sublevel contains one *s* orbital
- a *p* sublevel contains three *p* orbitals
- a *d* sublevel contains five *d* orbitals
- an *f* sublevel contains seven *f* orbitals



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Electron Capacity

TABLE 5.2 Electron Capacity in Sublevels for Energy Levels 1–4

Energy Level (n)	Number of Sublevels	Type of Sublevels	Number of Orbitals	Maximum Number of Electrons	Total Electrons ($2n^2$)
4	4	4*f*	7	14	32
		4*d*	5	10	
		4*p*	3	6	
		4*s*	1	2	
3	3	3*d*	5	10	18
		3*p*	3	6	
		3*s*	1	2	
2	2	2*p*	3	6	8
		2*s*	1	2	
1	1	1*s*	1	2	2

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The total number of electrons in all the sublevels adds up to give the maximum number of electrons ($2n^2$) allowed in an energy level.

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Learning Check

Indicate the number and type of orbitals in each of the following:

A. 4*s* sublevel

B. 3*d* sublevel

C. $n = 3$

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Solution

Indicate the number and type of orbitals in each of the following:

A. 4*s* sublevel

one 4*s* orbital

B. 3*d* sublevel

five 3*d* orbitals

C. $n = 3$

one 3*s* orbital, three 3*p* orbitals, and five 3*d* orbitals

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Learning Check

The number of

A. electrons that can occupy a *p* orbital is

1) 1 2) 2 3) 3

B. *p* orbitals in the 2*p* sublevel is

1) 1 2) 2 3) 3

C. *d* orbitals in the n = 4 energy level is

1) 1 2) 3 3) 5

D. electrons that can occupy the 4*f* sublevel is

1) 2 2) 6 3) 14

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Solution

The number of

A. electrons that can occupy a *p* orbital is

2) 2

B. *p* orbitals in the 2*p* sublevel is

3) 3

C. *d* orbitals in the *n* = 4 energy level is

3) 5

D. electrons that can occupy the 4*f* sublevel is

3) 14

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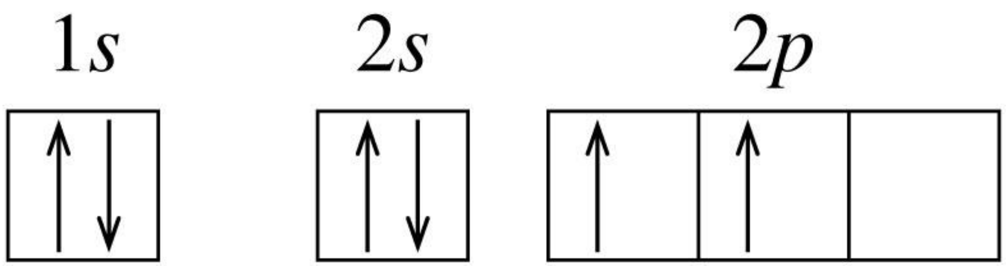
Chapter 5 Electronic Structure and Periodic Trends

5.4

Drawing Orbital Diagrams and Writing Electron Configurations



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Orbital diagram of carbon

In the orbital diagram of carbon, two electrons occupy the 1*s* orbital, two electrons occupy the 2*s* orbital, and two electrons each occupy a 2*p* orbital in the 2*p* sublevel.

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Order of Filling

Energy levels fill with electrons

- in order of increasing energy
- beginning with quantum number $n = 1$
- beginning with *s* followed by *p*, *d*, and *f*

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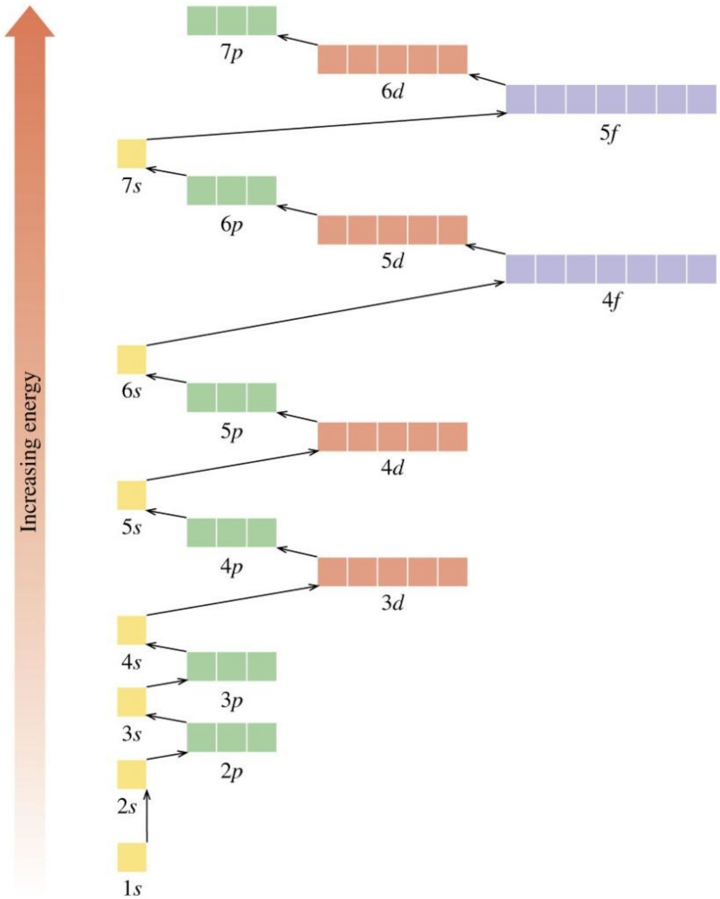
Energy Diagram for Sublevels



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The orbitals of an atom fill in order of increasing energy of the sublevels beginning with 1*s*.

Orbital Diagrams

An **orbital diagram** shows

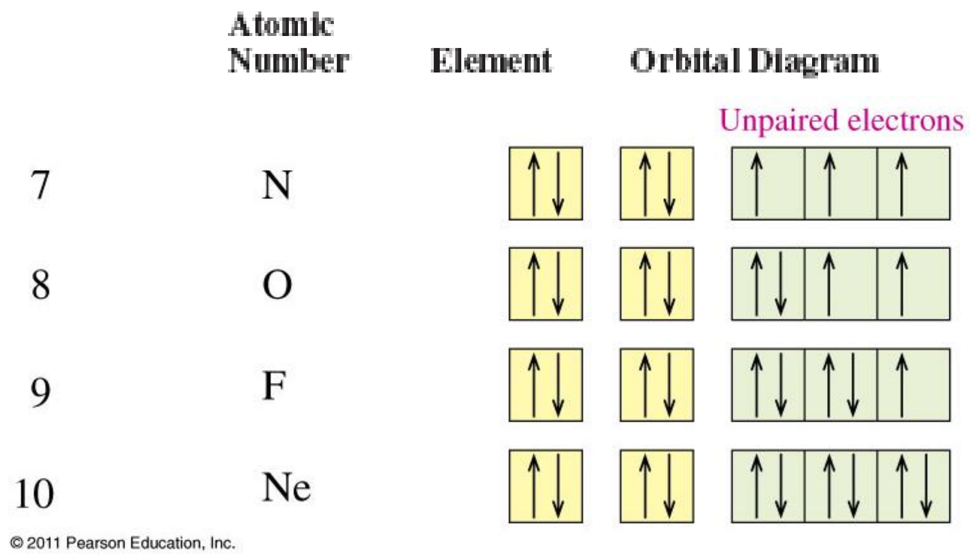
- orbitals as boxes in each sublevel
- electrons in orbitals as vertical arrows
- electrons in the same orbital with opposite spins (up and down vertical arrows)

Atomic Number	Element	Orbital Diagram	
		$1s$	$2s$
3	Li	$\boxed{\uparrow\downarrow}$	$\boxed{\uparrow}$

Order of Filling

Electrons in an atom

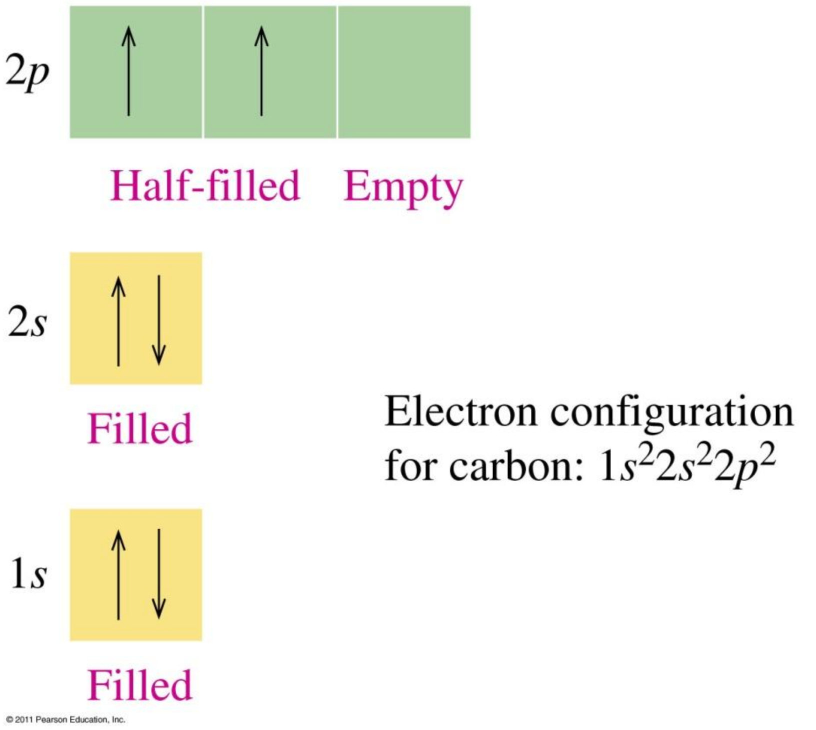
- fill each orbital in a sublevel with one electron until half full
- then pair up with an electron of opposite spin



Writing Orbital Diagrams

The **orbital diagram** for carbon consists of

- two electrons in the 1*s* orbital
- two electrons in the 2*s* orbital
- one electron each in two of the 2*p* orbitals



Learning Check

Write the orbital diagrams for

A. nitrogen

B. oxygen

C. magnesium

Solution

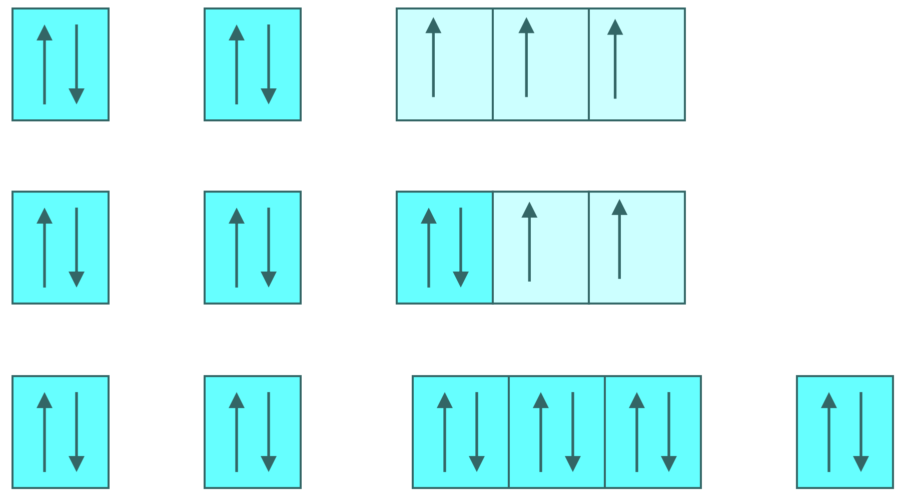
Write the orbital diagrams for

$1s$ $2s$ $2p$ $3s$

A. nitrogen

B. oxygen

C. magnesium



Electron Configuration

An **electron configuration**

- lists the sublevels filling with electrons in order of increasing energy
- uses superscripts to show the number of electrons in each sublevel
- for carbon is as follows:

Type of orbital

Number of electrons

$1s^2 2s^2 2p^2$ Read as "one *s* two, two *s* two, two *p* two"

Period 1 Configurations

In Period 1, the first two electrons go into the 1*s* orbital.

Atomic Number	Element	Orbital Diagram	Electron Configuration
		1*s*	
1	H	↑	$1s^1$
2	He	↑↓	$1s^2$

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Abbreviated Configurations

An **abbreviated configuration** shows

- the symbol of the noble gas in brackets that represents completely filled sublevels
- the remaining electrons in order of their sublevels

Example: Fluorine has a configuration and abbreviated electron configuration of

Element	Orbital Diagram	Electron Configuration	Abbreviated Electron Configuration
F	[↑↓] [↑↓] [↑↓ | ↑↓ | ↑]	$1s^2 2s^2 2p^5$	$[He]2s^2 2p^5$

Period 2 Configurations

Atomic Number	Element	Orbital Diagram 1s	2s	2p	Electron Configuration	Abbreviated Electron Configuration
3	Li	↑↓	↑		$1s^2 2s^1$	$[He]2s^1$
4	Be	↑↓	↑↓		$1s^2 2s^2$	$[He]2s^2$
5	B	↑↓	↑↓	↑ | |	$1s^2 2s^2 2p^1$	$[He]2s^2 2p^1$
6	C	↑↓	↑↓	↑ | ↑ |	$1s^2 2s^2 2p^2$	$[He]2s^2 2p^2$
7	N	↑↓	↑↓	Unpaired electrons ↑ | ↑ | ↑	$1s^2 2s^2 2p^3$	$[He]2s^2 2p^3$
8	O	↑↓	↑↓	↑↓ | ↑ | ↑	$1s^2 2s^2 2p^4$	$[He]2s^2 2p^4$
9	F	↑↓	↑↓	↑↓ | ↑↓ | ↑	$1s^2 2s^2 2p^5$	$[He]2s^2 2p^5$
10	Ne	↑↓	↑↓	↑↓ | ↑↓ | ↑↓	$1s^2 2s^2 2p^6$	$[He]2s^2 2p^6$

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Period 3 Configurations

Atomic Number	Element	Orbital Diagram (3*s* and 3*p* orbitals only) 3*s*	3*p*	Electron Configuration	Abbreviated Electron Configuration
11	Na	[Ne] ↑					$1s^22s^22p^63s^1$	$[Ne]3s^1$
12	Mg	[Ne] ↑↓					$1s^22s^22p^63s^2$	$[Ne]3s^2$
13	Al	[Ne] ↑↓		↑				$1s^22s^22p^63s^23p^1$	$[Ne]3s^23p^1$
14	Si	[Ne] ↑↓		↑	↑			$1s^22s^22p^63s^23p^2$	$[Ne]3s^23p^2$
15	P	[Ne] ↑↓		↑	↑	↑		$1s^22s^22p^63s^23p^3$	$[Ne]3s^23p^3$
16	S	[Ne] ↑↓		↑↓	↑	↑		$1s^22s^22p^63s^23p^4$	$[Ne]3s^23p^4$
17	Cl	[Ne] ↑↓		↑↓	↑↓	↑		$1s^22s^22p^63s^23p^5$	$[Ne]3s^23p^5$
18	Ar	[Ne] ↑↓		↑↓	↑↓	↑↓		$1s^22s^22p^63s^23p^6$	$[Ne]3s^23p^6$

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Learning Check

A. The correct electron configuration for nitrogen is

1) $1s^22p^5$ 2) $1s^22s^22p^6$ 3) $1s^22s^22p^3$

B. The correct electron configuration for oxygen is

1) $1s^22p^6$ 2) $1s^22s^22p^4$ 3) $1s^22s^22p^6$

C. The correct electron configuration for calcium is

1) $1s^22s^22p^63s^23p^63d^2$
2) $1s^22s^22p^63s^23p^64s^2$
3) $1s^22s^22p^63s^23p^8$

Solution

A. The correct electron configuration for nitrogen is

3) $1s^22s^22p^3$

B. The correct electron configuration for oxygen is

2) $1s^22s^22p^4$

C. The correct electron configuration for calcium

2) $1s^22s^22p^63s^23p^64s^2$

Learning Check

Write the electron configuration and abbreviated configuration for each of the following elements:

A. Cl

B. S

C. K

Solution

A. Cl

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

$[Ne]3s^2 3p^5$

B. S

$1s^2 2s^2 2p^6 3s^2 3p^4$

$[Ne]3s^2 3p^4$

C. K

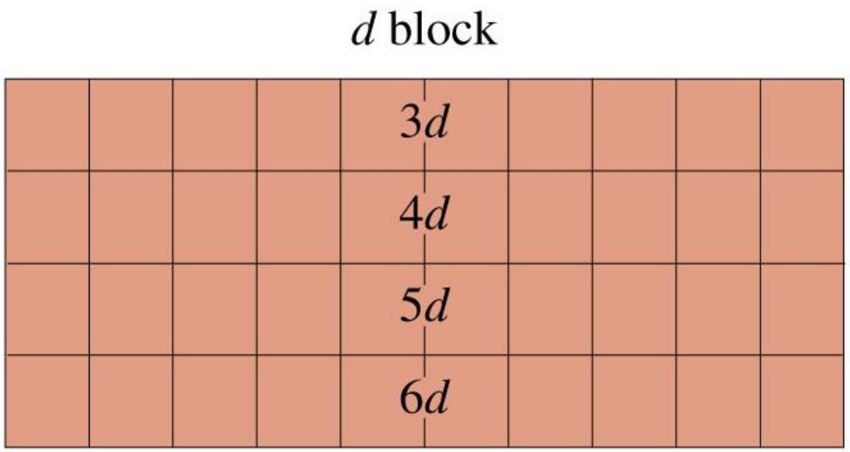
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

$[Ar]4s^1$

Chapter 5 Electronic Structure and Periodic Trends

5.5

Electron Configurations and the Periodic Table



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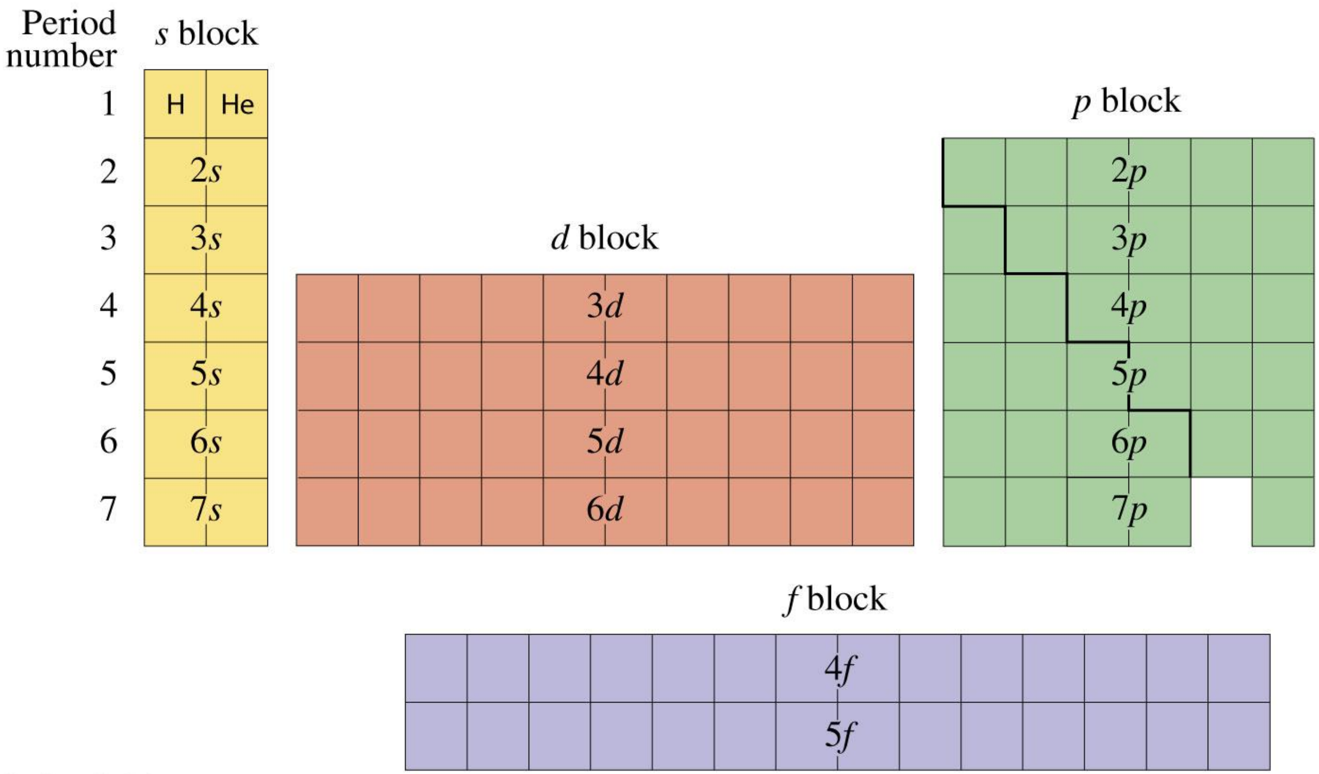
Sublevel Blocks on the Periodic Table

The periodic table consists of **sublevel blocks** arranged in order of increasing energy.

- Groups 1A(1)-2A(2) = *s* level
- Groups 3A(13)-8A(18) = *p* level
- Groups 3B(3) to 2B(12) = *d* level
- Lanthanides/Actinides = *f* level

Sublevel Blocks



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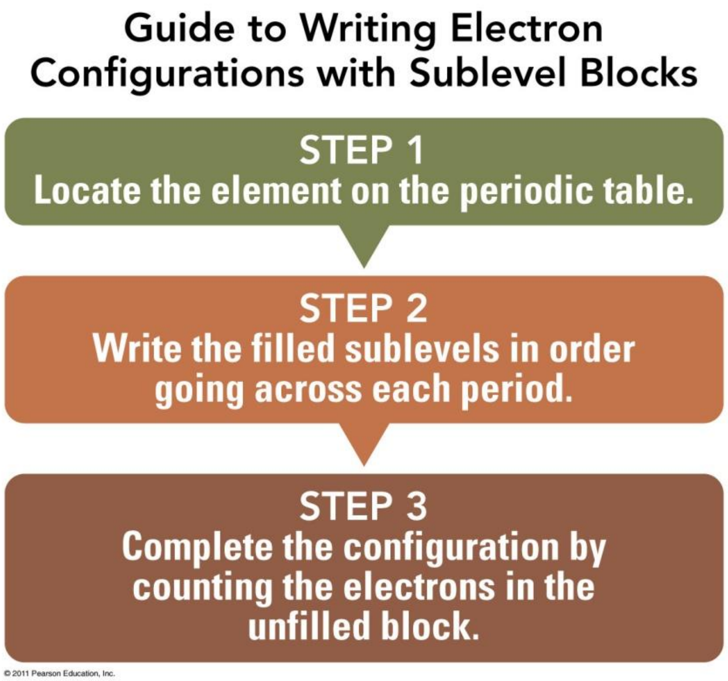
Electron configurations follow the order of sublevels on the periodic table.

Using Sublevel Blocks

To write an electron configuration using **Sublevel blocks**,

- locate the element on the periodic table
- starting with H in 1*s*,write each sublevel block in order going from left to right across each period
- write the number of electrons in each block



Writing Electron Configurations

Using the periodic table, write the electron configuration for silicon.

Solution

Period 1	1s block	$1s^2$
Period 2	2s $\rightarrow$ 2p blocks	$2s^2\,2p^6$
Period 3	3s $\rightarrow$ 3p blocks	$3s^23p^2$ (Si)

Writing all the sublevel blocks in order gives

$$1s^22s^22p^63s^23p^2$$

Electron Configurations *d* Sublevel

- The 4*s* orbital has a lower energy that the 3*d* orbitals.
- In potassium, K, the last electron enters the 4*s* orbital, not the 3*d* (*as shown below*).

	1*s*	2*s* 2p	3*s* 3*p* 3*d*	4s
Ar	$1s^2$	$2s^2\,2p^6$	$3s^2\,3p^6$	
K	$1s^2$	$2s^2\,2p^6$	$3s^2\,3p^6$	$4s^1$
Ca	$1s^2$	$2s^2\,2p^6$	$3s^2\,3p^6$	$4s^2$
Sc	$1s^2$	$2s^2\,2p^6$	$3s^2\,3p^6\,3d^1$	$4s^2$
Ti	$1s^2$	$2s^2\,2p^6$	$3s^2\,3p^6\,3d^2$	$4s^2$

Writing Electron Configurations

Using the periodic table, write the electron configuration for manganese.

Solution

Period 1	1*s* block	$1s^2$
Period 2	2*s* $\rightarrow$ 2*p* blocks	$2s^2$ $2p^6$
Period 3	3*s* $\rightarrow$ 3*p* blocks	$3s^2$ $3p^6$
Period 4	4*s* $\rightarrow$ 3*d* blocks	$4s^2$ $3d^5$ *(at Mn)*

Writing all the sublevel blocks in order gives

$$1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6\ 4s^2\ 3d^5$$

Writing Electron Configurations

Using the periodic table, write the electron configuration for iodine.

Solution

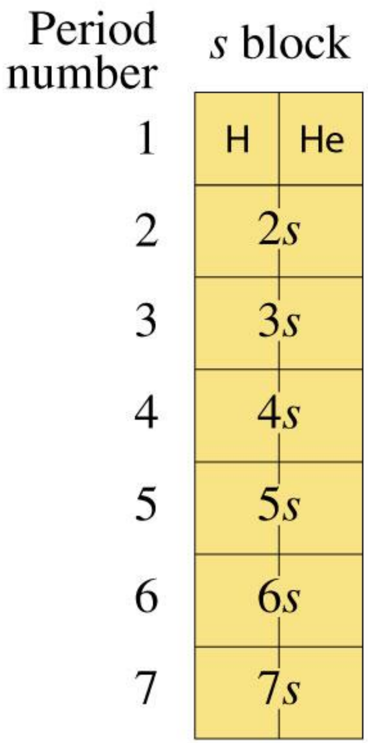
Period 1	1*s* block	$1s^2$
Period 2	2*s* → 2*p* blocks	$2s^2$ $2p^6$
Period 3	3*s* → 3*p* blocks	$3s^2$ $3p^6$
Period 4	4*s* → 3*d* → 3*p* blocks	$4s^2$ $3d^{10}$ $4p^6$
Period 5	5*s* → 4*d* → 5*p* blocks	$5s^2$ $4d^{10}$ $5p^5$

Writing all the sublevel blocks in order gives

$1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$

(iodine)

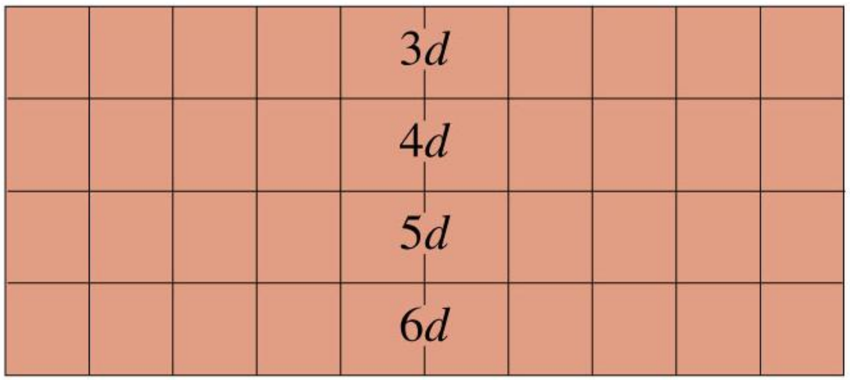
4s Block



Atomic Number	Element	Electron Configuration	Abbreviated Electron Configuration
4*s* Block			
19	K	$1s^22s^22p^63s^23p^64s^1$	$[Ar]4s^1$
20	Ca	$1s^22s^22p^63s^23p^64s^2$	$[Ar]4s^2$

3d Block

d block



3d Block

21	Sc	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$	$[Ar]4s^2 3d^1$
22	Ti	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$	$[Ar]4s^2 3d^2$
23	V	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$	$[Ar]4s^2 3d^3$
24	Cr*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$	$[Ar]4s^1 3d^5$ (half-filled *d* sublevel is stable)
25	Mn	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$	$[Ar]4s^2 3d^5$
26	Fe	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$	$[Ar]4s^2 3d^6$
27	Co	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$	$[Ar]4s^2 3d^7$
28	Ni	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$	$[Ar]4s^2 3d^8$
29	Cu*	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$	$[Ar]4s^1 3d^{10}$ (filled *d* sublevel is stable)
30	Zn	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$	$[Ar]4s^2 3d^{10}$

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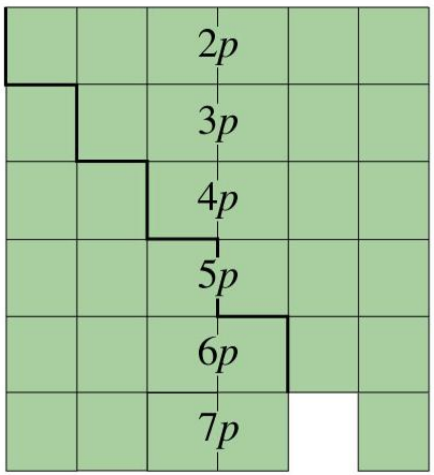
4*p* Block

p block



Atomic Number	Element	Electron Configuration	Abbreviated Electron Configuration
4*p* Block			
31	Ga	$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$	$[Ar]4s^23d^{10}4p^1$
32	Ge	$1s^22s^22p^63s^23p^64s^23d^{10}4p^2$	$[Ar]4s^23d^{10}4p^2$
33	As	$1s^22s^22p^63s^23p^64s^23d^{10}4p^3$	$[Ar]4s^23d^{10}4p^3$
34	Se	$1s^22s^22p^63s^23p^64s^23d^{10}4p^4$	$[Ar]4s^23d^{10}4p^4$
35	Br	$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$	$[Ar]4s^23d^{10}4p^5$
36	Kr	$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$	$[Ar]4s^23d^{10}4p^6$

*Exceptions to the order of filling.

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Learning Check

A. The last two sublevel blocks in the electron configuration for Co are

1) $3p^6 4s^2$
2) $4s^2 4d^7$
3) $4s^2 3d^7$

B. The last three sublevel blocks in the electron configuration for Sn are

1) $5s^2 5p^2 4d^{10}$
2) $5s^2 4d^{10} 5p^2$
3) $5s^2 5d^{10} 5p^2$

Solutions

A. The last two sublevel blocks in the electron configuration for Co are
3) $4s^2 3d^7$

B. The last three sublevel blocks in the electron configuration for Sn are
2) $5s^2 4d^{10} 5p^2$

Learning Check

Using the periodic table, write the electron configuration and abbreviated configuration for each of the following elements:

A. Zn

B. Sr

C. I

Solution

A. Zn

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

$[Ar]\ 4s^2 3d^{10}$

B. Sr

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$

$[Kr] 5s^2$

C. I

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^5$

$[Kr] 5s^2 4d^{10} 5p^5$

Learning Check

Give the symbol of the element that has

A. $[Ar]4s^2 3d^6$

B. Four $3p$ electrons

C. Two electrons in the $4d$ sublevel

D. Electron configuration

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$$

Solution

Give the symbol of the element that has

A. $[Ar]4s^2 3d^6$ Fe

B. Four 3*p* electrons S

C. Two electrons in the 4*d* sublevel Zr

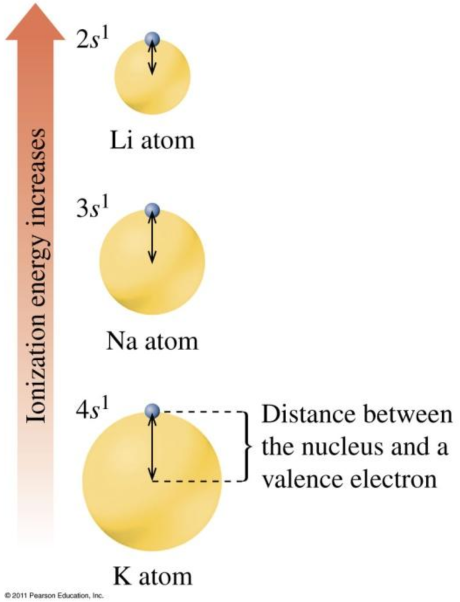
D. Electron configuration Ti

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

Chapter 5 Electron Configuration and Periodic Trends

5.6

Periodic Trends of the Elements



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Valence Electrons

The **valence electrons**

- determine the chemical properties of an element
- are the electrons in the *s* and *p* sublevels in the highest energy level
- are related to the group number of the element

Example: Phosphorus has 5 valence electrons

P Group 5A(15) $1s^22s^22p^6\overbrace{3s^23p^3}^{\textit{5 valence electrons}}$

Group Number and Valence Electrons

All the elements in a group have the same number of valence electrons.

Example:

Elements in Group 2A (2) have two (2) valence electrons.

Be	$1s^2\ 2s^2$
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
Sr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$

Periodic Table and Valence Electrons

TABLE 5.3 Valence Electrons for Representative Elements in Periods 1–4

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
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 $3s^23p^6$
19 K $4s^1$	20 Ca $4s^2$	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 $4s^24p^6$

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Learning Check

State the number of valence electrons for each:

A. O

1) 4 2) 6 3) 8

B. Al

1) 13 2) 3 3) 1

C. Cl

1) 2 2) 5 3) 7

Solution

State the number of valence electrons for each.

A. O

2) 6

B. Al

2) 3

C. Cl

3) 7

Learning Check

State the number of valence electrons for each.

A. Calcium

1) 1 2) 2 3) 3

B. Group 6A (16)

1) 2 2) 4 3) 6

C. Tin

1) 2 2) 4 3) 14

Solution

State the number of valence electrons for each.

A. Calcium

2) 2

B. Group 6A (16)

3) 6

C. Tin

2) 4

Learning Check

State the number of valence electrons for each.

A. $1s^2 2s^2 2p^6 3s^2 3p^3$

B. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$

C. $1s^2 2s^2 2p^5$

Solution

State the number of valence electrons for each.

A. $1s^22s^22p^63s^23p^3$ 5

B. $1s^22s^22p^63s^23p^64s^23d^{10}4p^4$ 6

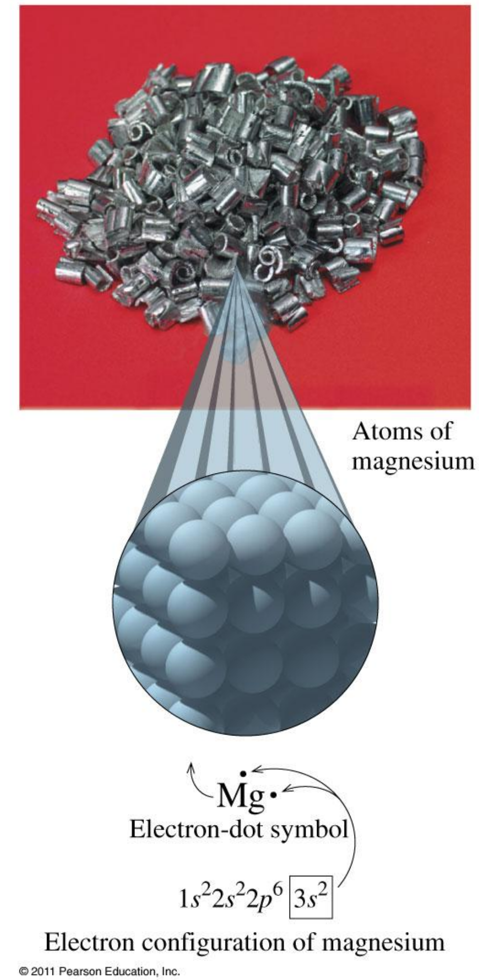
C. $1s^22s^22p^5$ 7

Electron-Dot Symbols

An **electron-dot symbol**

- indicates the valence electrons as dots around the symbol of the element
- for Mg shows two valence electrons placed as single dots on the sides of the symbol Mg

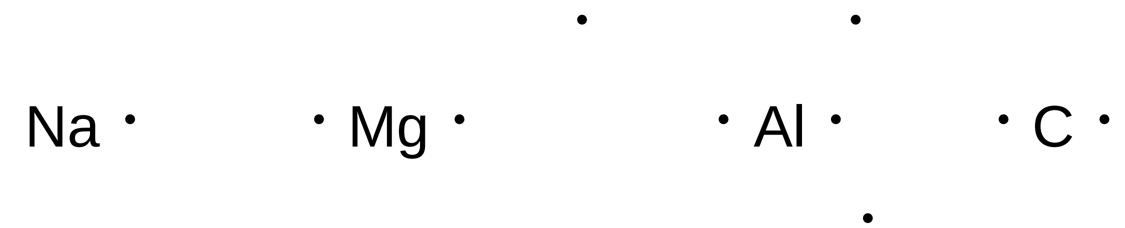
$$\cdot\text{Mg}\cdot \quad \text{or} \quad \overset{\cdot}{\text{Mg}}\cdot \quad \text{or} \quad \cdot\overset{\cdot}{\text{Mg}} \quad \text{or} \quad \cdot\,\text{Mg}\,\cdot$$



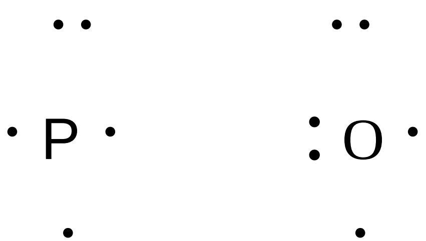
Writing Electron-Dot Symbols

The **electron-dot symbols** for

- Groups 1A (1) to 4A (14) use single dots



- Groups 5A (15) to 7A (17) use pairs and single dots



Groups and Electron-Dot Symbols

- In a group, all the electron-dot symbols have the same number of valence electrons (dots).

Example: Atoms of elements in Group 2A (2) each have two valence electrons.

2A (2)

· Be ·

· Mg ·

· Ca ·

· Sr ·

· Ba ·

Periodic Table and Electron-Dot Symbols

TABLE 5.4 Electron-Dot Symbols for Selected Elements in Periods 1–4

	Group Number							
	1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
Number of Valence Electrons	1	2	3	4	5	6	7	8
Electron-Dot Symbols	H·							He:
	Li·	Be·	·B·	·C·	·N·	·O:	·F:	:Ne:
	Na·	Mg·	·Al·	·Si·	·P·	·S:	·Cl:	:Ar:
	K·	Ca·	·Ga·	·Ge·	·As·	·Se:	·Br:	:Kr:

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Learning Check

A. $\dot{\mathrm{X}}$ is the electron-dot symbol for

1) Na 2) K 3) Al

B.
  • •
• X •
   •
is the electron-dot symbol of

1) B 2) N 3) P

Solution

A. $\dot{\mathrm{X}}$ is the electron-dot symbol for

1) Na   2) K

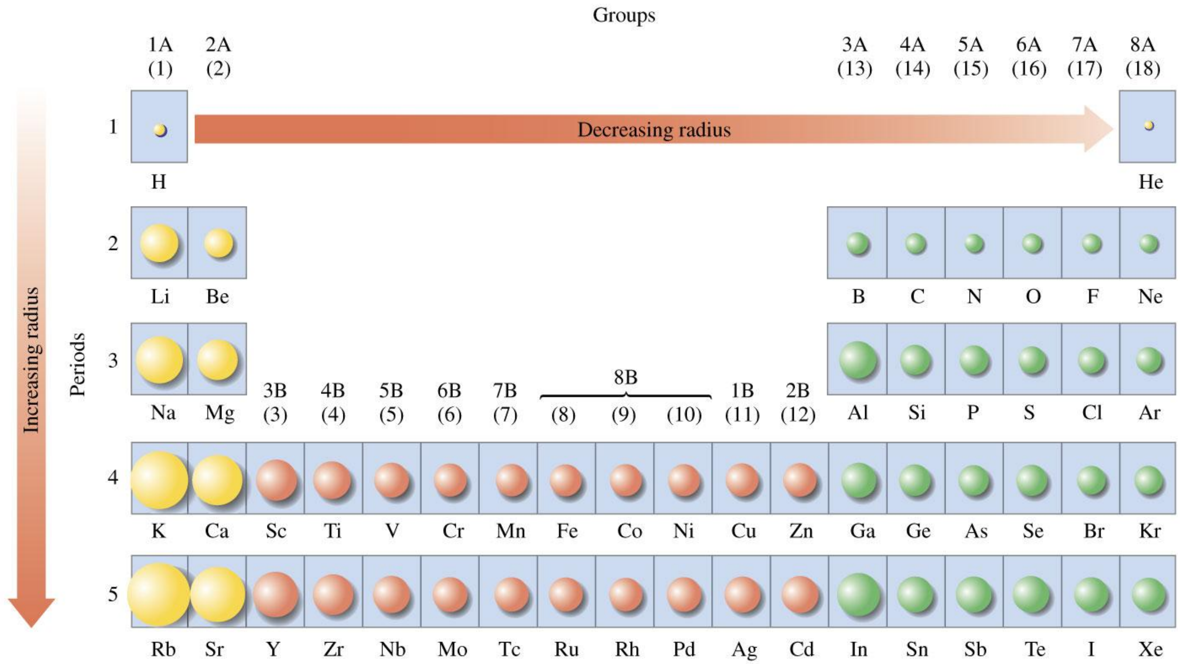
B. $\cdot\overset{\cdot\cdot}{\underset{\cdot}{\mathrm{X}}}\cdot$ is the electron-dot symbol of

2) N   3) P

Atomic Radius

The **atomic radius**

- is the distance from the nucleus to the valence electrons



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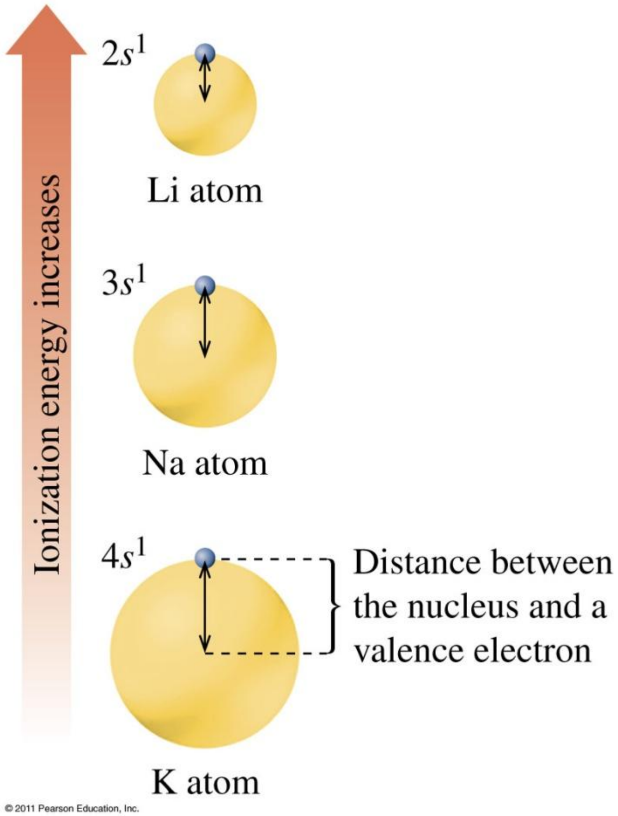
Atomic Radius within a Group

The **atomic radius** increases

- going down each group of representative elements
- as the number of energy levels increases



Atomic Radius across a Period

The **atomic radius** decreases

- going from left to right across a period
- as more protons increase the nuclear attraction for valence electrons

Learning Check

Select the element in each pair with the larger atomic radius.

A. Li or K
B. K or Br
C. P or Cl

Solution

Select the element in each pair with the larger atomic radius.

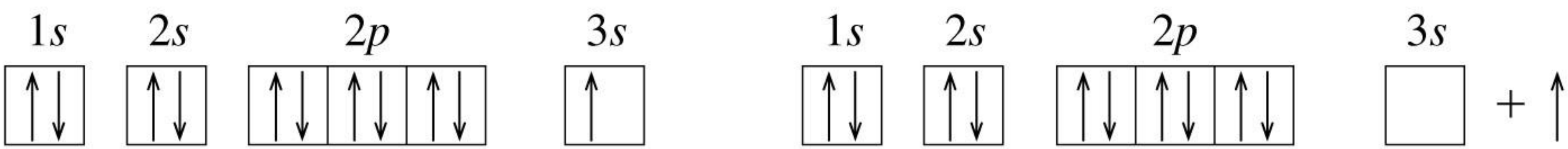
A. K is larger than Li
B. K is larger than Br
C. P is larger than Cl

Ionization Energy

Ionization energy

- is the energy it takes to remove a valence electron

$$Na(g) + \text{energy (ionization)} \longrightarrow Na^+ + e^-$$



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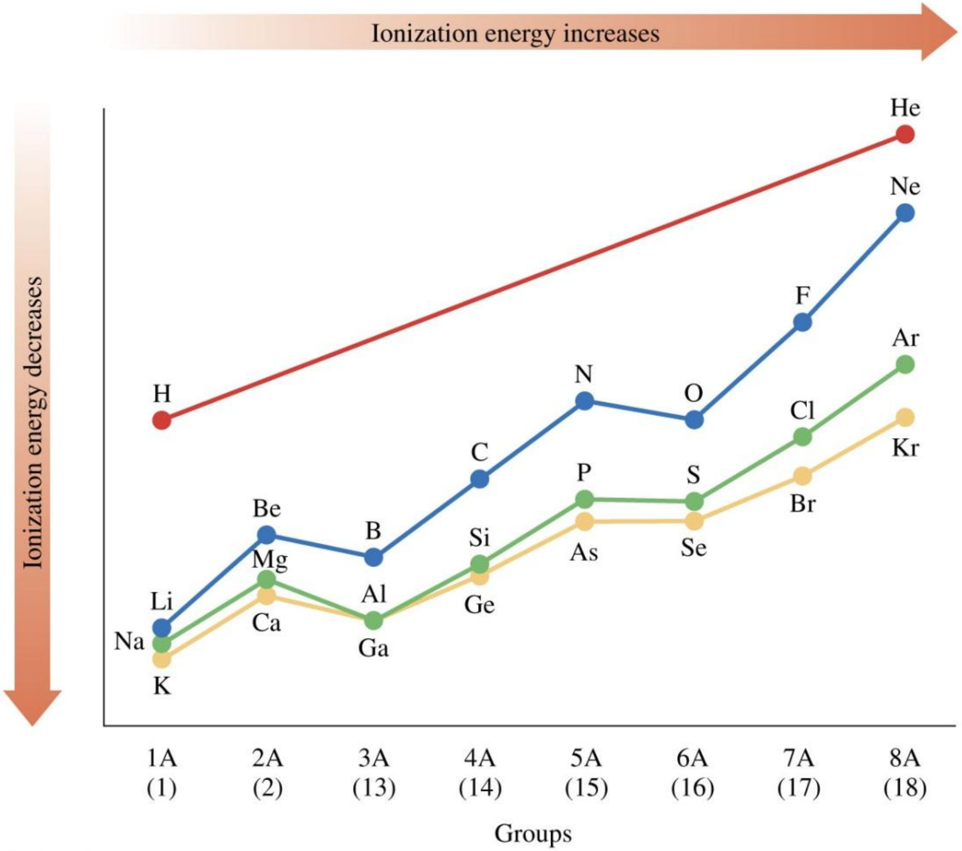
Ionization Energy

Metals have

- 1-3 valence electrons
- lower ionization energies



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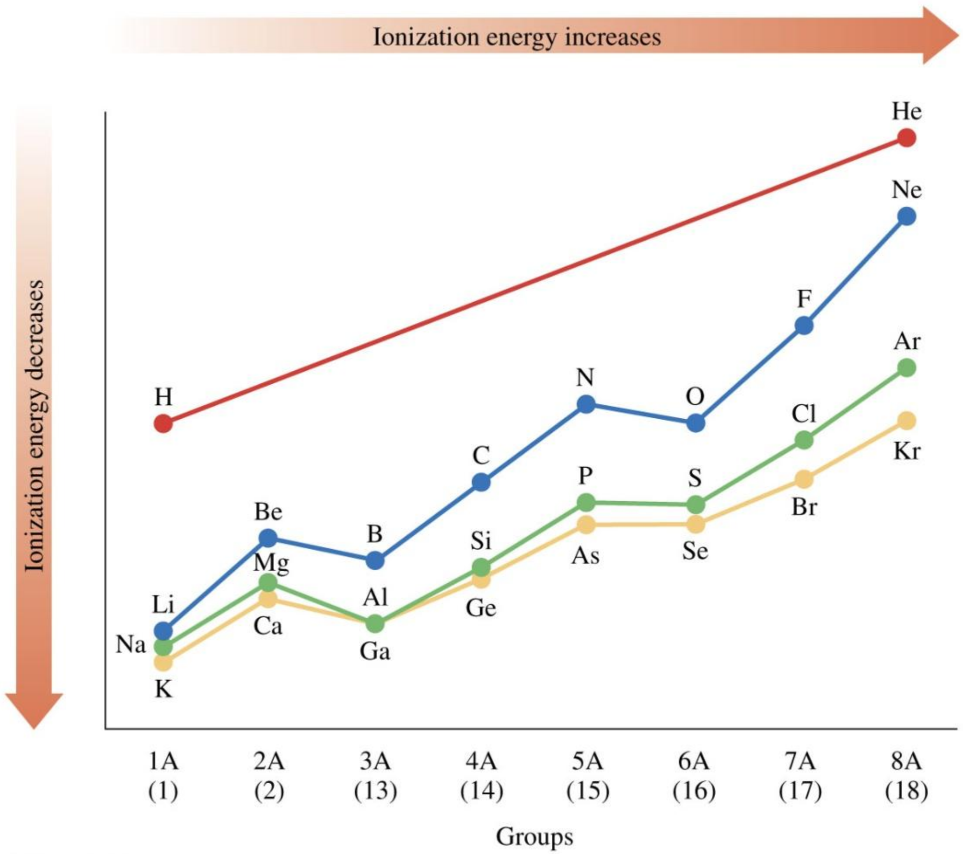
Ionization Energy

Nonmetals have

- 5-7 valence electrons
- higher ionization energies



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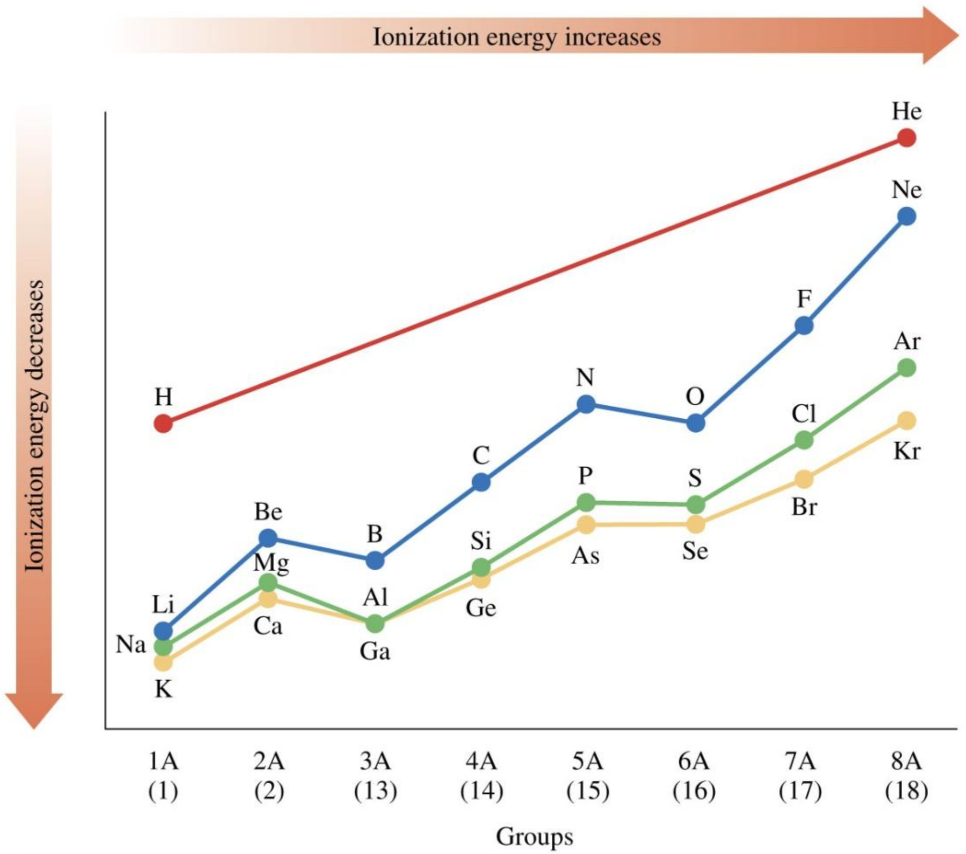
Ionization Energy

Noble gases have

- complete octets (He has two valence electrons)
- the highest ionization energies in each period



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Learning Check

Select the element in each pair with the higher ionization energy.

A. Li or K
B. K or Br
C. P or Cl

Solution

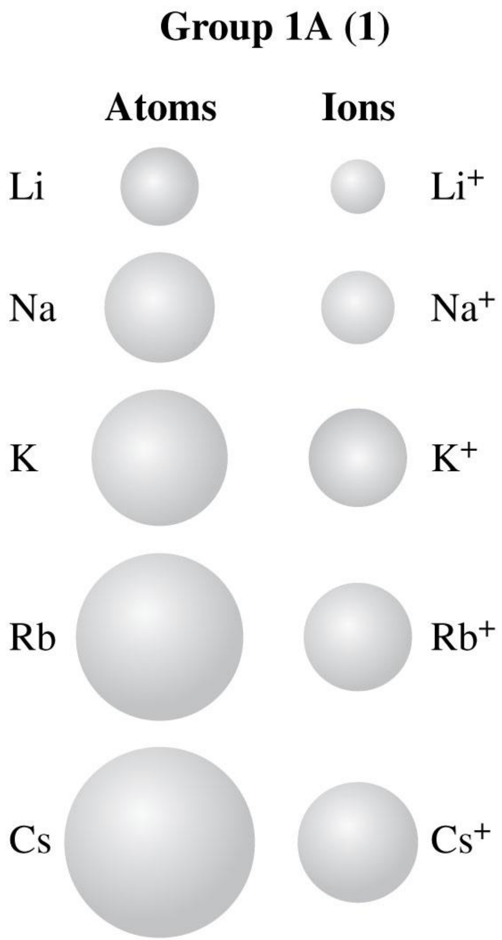
Select the element in each pair with the higher ionization energy.

A. Li
B. Br
C. Cl

Sizes of Metal Atoms and Ions

A **positive ion**

- has lost its valence electrons
- is smaller than the corresponding metal atom (about half the size)



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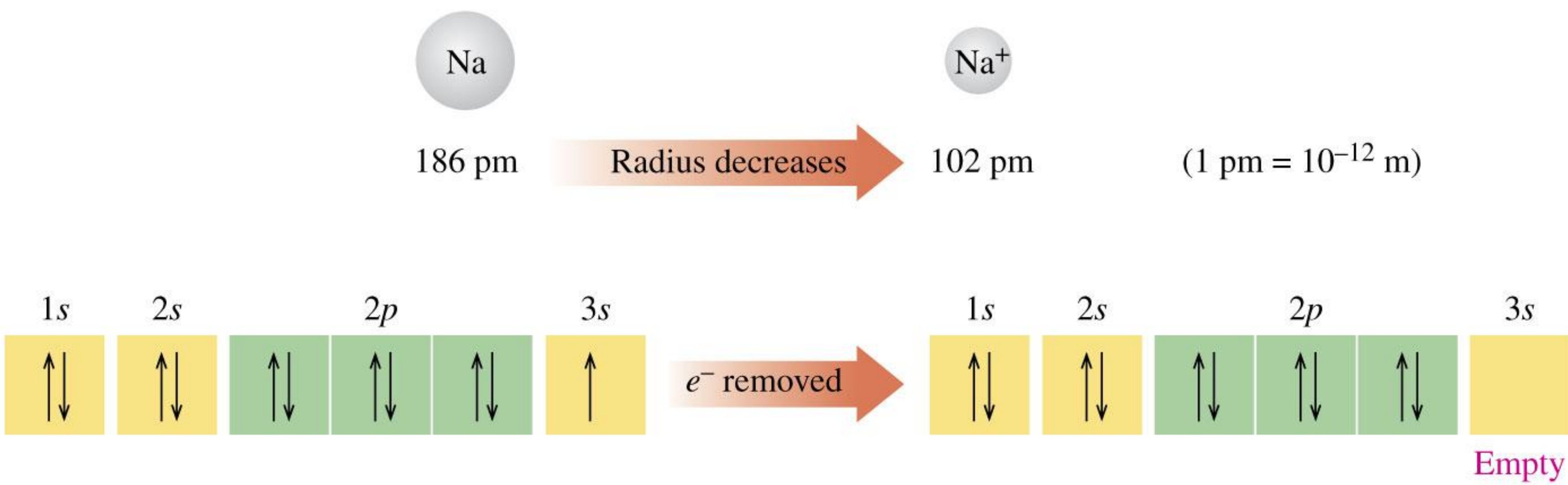
Size of Sodium Ion

The sodium ion Na^+

- forms when the Na atom loses one electron from the third energy level
- is smaller than a Na atom



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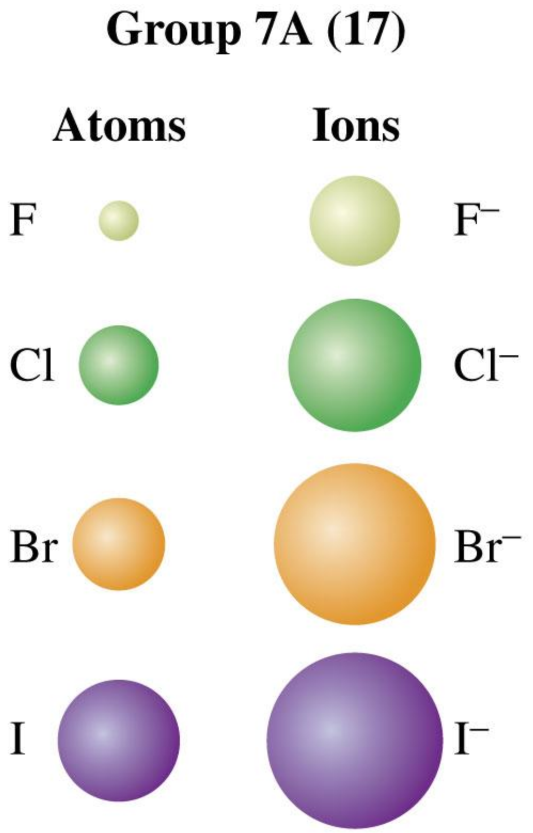
Sizes of Nonmetal Atoms and Ions

A **negative ion**

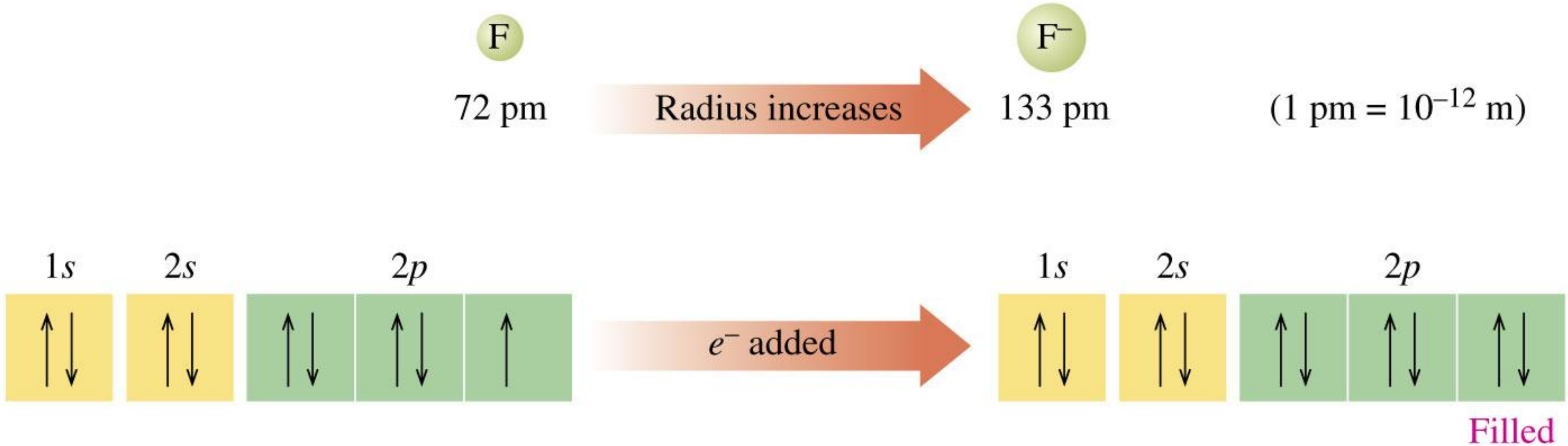
- has a complete octet
- increases the number of valence electrons
- is larger than the corresponding nonmetal atom (about twice the size)



Size of Fluoride Ion

The fluoride ion F^-

- forms when a valence electron is added
- has increased repulsions due to the added valence electron
- is larger than a F atom



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Learning Check

1. Which is larger in each of the following?
   A. K or K^+
   B. Al or Al^{3+}
   C. S^{2-} or S

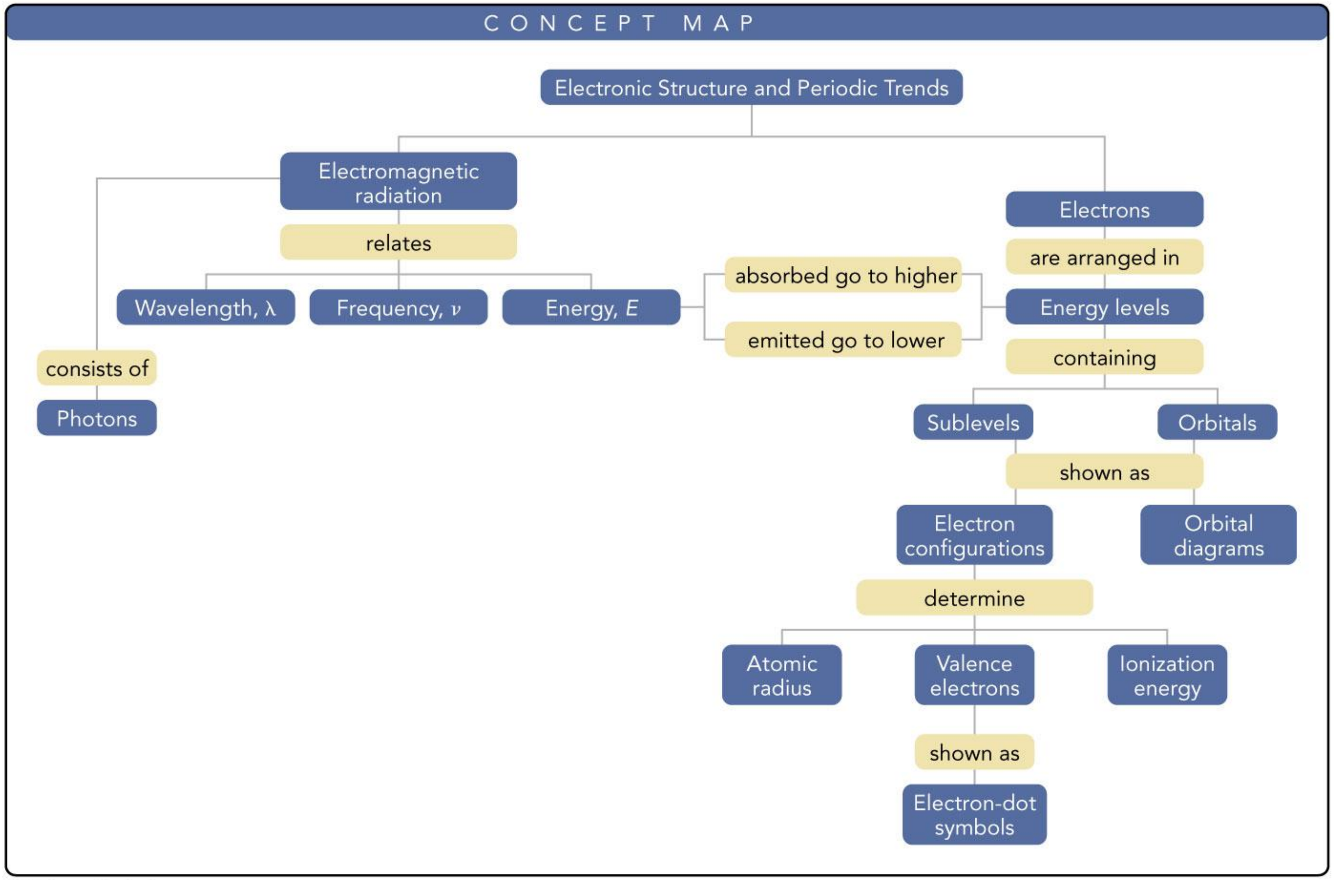
2. Which is smaller in each of the following?
   A. N^{3-} or N
   B. Cl or Cl^-
   C. Sr^{2+} or Sr

Solution

1. Which is larger in each of the following?
   A. $K > K^{+}$
   B. $Al > Al^{3+}$
   C. $S^{2-} > S$

2. Which is smaller in each of the following?
   A. $N < N^{3-}$
   B. $Cl < Cl^{-}$
   C. $Sr^{2+} < Sr$

Concept Map



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