

Chemistry 10: Electron “Orbitals”

March 23-27

Time Allotment: 40 minutes per day

Student Name: _____

Teacher Name: _____

Academic Honesty

I certify that I completed this assignment
independently in accordance with the GHNO
Academy Honor Code.

Student signature:

I certify that my student completed this
assignment independently in accordance with
the GHNO Academy Honor Code.

Parent signature:

Packet Overview

Date	Objective(s)	Page Number
Monday, March 23	<i>Identify an energy-based understanding of orbitals and begin to intuit that electrons orbit atoms in energetic levels.</i>	2
Tuesday, March 24	<i>Begin to notice the distinction between “inner” and “outer” electrons.</i>	5
Wednesday, March 25	<i>Identify different orbital shapes and begin to name the four different “types” of orbital by shape.</i>	9
Thursday, March 26	<i>Identify how many electrons an atom has in each of the orbital level types in the first 3 rows of the periodic table (levels 1, 2, and 3).</i>	13
Friday, March 27	<i>Identify how many electrons an atom has in each of the orbital level types in all rows of the periodic table.</i>	18

Additional Notes: To parents (and you wonderful older siblings): **this packet is designed to be accessible to any student of chemistry.** An adult simply using the information contained in this packet should be able to understand and work the problems contained herein. **However, it is a progressive packet, and so if your student asks for help on day 4, you may have to work through days 1-3 before day 4 makes sense to you.** Thank you for working with this under these non-ideal conditions. It is my belief that such a packet can and will substitute for in-person teaching. The wording and presentation style is consistent with what your students are used to in Chemistry this year. Any requisite information that was covered at an earlier point in the year is referred to explicitly or simply re-stated, to minimize the amount of implicit background information required by the student.

The textbook *can* be a good resource here. However, it involves more material which is not appropriate for a general-level course, and so primarily using this packet as a didactic tool is recommended for time-efficient reasons. Helpful section summaries related to the material in this packet are found on page 250 of our textbook, and they can be used to reference full sections in the book.

As far as online resources, Chemistry LibreTexts as a variety of presentation modes that are worth exploring for the interested student looking for alternative ways to approach the material addressed in this packet. I’d encourage simply searching “Chemistry LibreTexts [key-word]” to find relevant material, and can provide links to supplemental information upon request. Khan Academy is also an excellent resource, though again the approach that Khan Academy’s instructors take may be different from the one taken in this packet, which has been designed specifically for students taking Chemistry at Northern Oaks in the year 2019-2020.

Monday, March 23

Chemistry Unit: Electrons and orbitals
Lesson 1: “Ionization” Energy

Unit Overview

Ionization energy is defined as the minimum motive force required to remove a specific electron from an atom (thereby making it an ion).

The closer an electron is to the (positive) nucleus, the higher its ionization energy.

All other things remaining equal, the more protons an atom has, the higher the ionization energy it has.

Lesson 1 Socratic Question: Keep these questions in mind as you study this lesson!

How do electrons orbit the (positive) nucleus?

What sort of patterns do they make? Would it just be a sort of clump? Since they are moving, can they all be the same distance from the nucleus?

Objective: Be able to do this by the end of this lesson.

Be able to explain what ionization energy is to another adult and how distance from the nucleus affects an electron’s ionization energy.

Introduction to Lesson 1

“Ionization Energy”

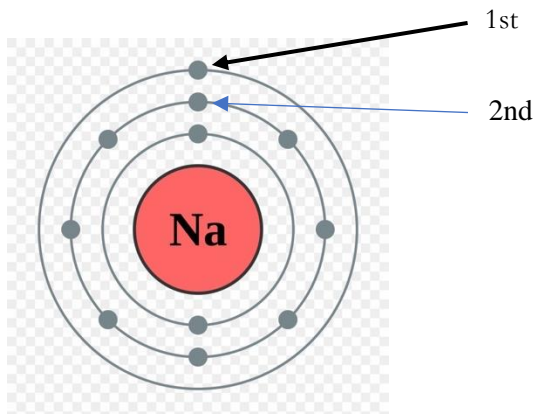
Ionization energy is defined as the minimum motive force required to remove a specific electron from an atom (thereby making it an ion). This can be done in an almost-infinite variety of ways.

Question: How might you remove electrons from a substance? (Think of the ways you can “add” energy to a substance, say in a lab or kitchen setting).

For instance, you can convert Cesium (Cs) into its common ion (Cs^+), removing one of its electrons, by shining a high-frequency UV light onto it.

Each electron has a specific ionization energy which changes based on a few factors. We’ll now consider the question: **what are those factors?**

A picture of a sodium atom (Na) is shown below. Observe the electrons (dots) surrounding the (positive) nucleus (ignore the lines right now—just use them to gauge distance from the nucleus).



Q: Remembering that positive objects attract negative objects, would you expect the outermost electron to take more energy or less energy to remove than the electron closer to the center (2nd)?

Write your answer here:

The outermost electron's ionization energy would be higher/lower (circle one).

The innermost electron's ionization energy would be higher/lower (circle one).

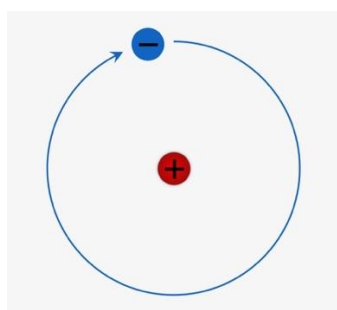
The electrons closer to the center of the atom will require **more** energy to remove because, being closer, the nucleus exerts a stronger pull on their negative charges.

Memorize this definition of ionization energy: the minimum motive force required to remove a specific electron from an atom.

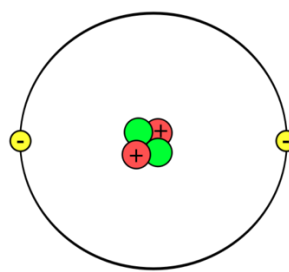
Memorize this key fact about ionization energy: the closer an electron is to the (positive) nucleus, the more energy it takes to remove.

Other Effects

Observe the Bohr models of the first two elements below.



Hydrogen



Helium

The ionization energy of Hydrogen’s electron is 1312 kJ/mol. On the other hand, the ionization energy(s) for Helium’s electrons are much higher (2372 kJ/mol for the first one, 5276 kJ/mol for the second one).

Q: Why is the ionization energy of Helium’s electron(s) so much higher than Hydrogen’s ionization energy, even though they are the same distance from the nucleus?

Write your answer here:

In addition to distance, the **number of protons** an atom has will affect the ionization energy of an electron. Because Helium has twice as many protons, its ionization energy is significantly (though not exactly two times) higher.

Memorize this key fact pertaining to electron pull: the more protons an atom has, the stronger the pull it exerts on its electrons.

Closing: Check your understanding of the lesson by articulating an answer to the following questions.

1. It takes 495 kJ/mol of energy to remove one electron from Na, but it takes 4562 kJ/mol of energy to remove the second electron. Why?

2. What are the two factors involved when thinking about how high an ionization energy will be?

3. Does ionization energy apply to **atoms** or to **electrons** (think carefully about this one)?

4. Determine how many protons Helium, Aluminum, Silicon, and Chlorine have (remember, use your periodic table).

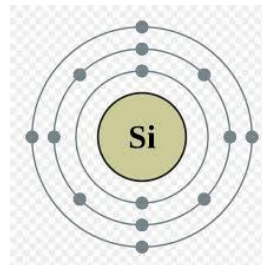
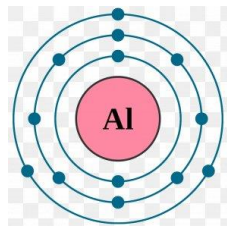
He: _____ protons

Al: _____ protons

Si: _____ protons

Cl: _____ protons

5. Observe Aluminum and Silicon’s Bohr models below.



You would expect aluminum and silicon’s **first** ionization energy (the energy it takes to remove the **outermost** electron) to be:

- a. Al: 780 kJ/mol; Si: 578 kJ/mol
- b. Al: 578 kJ/mol; Si: 780 kJ/mol
- c. Al: 578 kJ/mol; Si: 578 kJ/mol

There are two factors to consider.

For one, silicon has more electrons. Therefore, its outermost electron might be farther from the nucleus. However, the picture tells us (and we’ll learn more about the picture soon) that silicon’s extra electron is *not* any further from the nucleus than aluminum’s. Therefore, distance is not going to have an effect on ionization energy.

Second: silicon has more **protons**. Therefore, its electrons are more strongly pulled towards the center. It will probably be harder to remove its outermost electron. We expect Silicon to have a higher first ionization energy.

The second answer is correct.

Tuesday, March 24

Chemistry Unit: Electrons and orbitals
Lesson 2: “Inner” & “Outer” Electrons

Unit Overview

As you move from left to right across a row of the periodic table, the first ionization energy tends to **increase**.

As you move down a column of the periodic table, the first ionization energy tends to **remain the same or decrease**.

These two observations are consistent with “octet rule” theory and help confirm that noble gases are very “stable”.

Chemistry 10: Electrons and “orbitals”

March 23-27

Lesson 2 Socratic Question: Keep these questions in mind as you study this lesson!

What are “valence” electrons and what do they have to do with bonding?

Objective: Be able to do this by the end of this lesson.

Be able to articulate to an adult the fact that as an element comes closer to having 8 valence electrons, the harder it is to remove an electron.

Introduction to Lesson 2

A table of elements with the first ionization energies (in kJ/mol) listed under each element is below.

Period	First Ionization Energies of Some Elements (kJ/mol)																	
	Group 1	Group 2											Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1	H 1310																	He 2370
2	Li 520	Be 900											B 800	C 1090	N 1400	O 1310	F 1680	Ne 2080
3	Na 490	Mg 730											Al 580	Si 780	P 1060	S 1000	Cl 1250	Ar 1520
4	K 420	Ca 590	Sc 630	Ti 660	V 650	Cr 660	Mn 710	Fe 760	Co 760	Ni 730	Cu 740	Zn 910	Ga 580	Ge 780	As 960	Se 950	Br 1140	Kr 1350
5	Rb 400	Sr 550	Y 620	Zr 660	Nb 670	Mo 680	Tc 700	Ru 710	Rh 720	Pd 800	Ag 730	Cd 870	In 560	Sn 700	Sb 830	Te 870	I 1010	Xe 1170
6	Cs 380	Ba 500	La 540	Hf 700	Ta 760	W 770	Re 760	Os 840	Ir 890	Pt 870	Au 890	Hg 1000	Tl 590	Pb 710	Bi 800	Po 810	At ...	Rn 1030
7	Fr ...	Ra 510																

Take a look at this table of elements and look to notice **trends**. Answer each of these questions and use them as a guide.

Look down each column—do you notice a significant change? Describe: _____

Look across each row—do you notice any significant change or lack of change?

Describe: _____

Where are the **highest** first ionization energies located? _____

Which elements have the **lowest** first ionization energies? _____

In **general**, does it become harder to remove an electron the further across the periodic table you go, or does it become easier? _____

Read and Annotate

As we move **across** the periodic table, elements have more protons. This would tend to mean that as you move from left to right (e.g. from K to Kr), the first ionization energy will go up (i.e. it is much harder to remove an electron from Kr or Se when compared to K or Sc). We see that this is true in most cases, with a few exceptions (Gallium has an unexpectedly low ionization energy,

and Zn has an unexpectedly high ionization energy for its place in the 4th row of the periodic table).

As we move **down** a column of the periodic table (e.g. for the first column), we notice that the ionization energies **don't** get higher—in fact, they tend to get lower. These numbers were determined in a lab, by shining light of higher frequencies onto samples of these elements and measuring the energy required to cause electrons to flow from a sample of the element. This is (almost totally) unexpected. Remember that, in general, more protons in the nucleus means the nucleus exerts a stronger pull on the electrons, and thus they're harder to pull off. We see the **opposite** effect in a column of the periodic table.

Why do you think we see ionization energies remain the same or slightly decrease as we move down a column? To use an example, why does it take 490 kJ/mol of energy to remove Na's first electron while it takes a similar amount of energy (420 kJ/mol) to remove K's first electron?

Write your answer here:

Read and Annotate:

We explain this phenomenon by first noting **trends**. We notice that, when elements get closer to having the same number of electrons as a noble gas, the ionization energy goes up, and the noble gases have the highest ionization energies of all the elements. This is consistent with what we've intuited before—that the noble gases are very **stable**, and it's hard to break down something that is stable. On the other hand, on the other side of the periodic table, the alkali metals (Li, Na, K, Rb, Cs) have the lowest ionization energies of all the elements. This is consistent with what we've intuited before—that the alkali metals are **almost** very stable, and to become stable they need to **lose one electron** in order to become like a noble gas (you'll remember that they thus take on a 1+ charge). The same rule applies for the alkali earth metals in the second column (Be, Mg, Ca, Sr, Ba) to a lesser extent—they have the second-lowest ionization energies, and they tend to lose 2 electrons in order to become like a noble gas. It's a little harder to remove those two electrons because their nuclei have more protons than the alkali metals to their left, so their ionization energies are a little higher, but they still have very low ionization energies (see table).

So, one way of saying what we've said is that atoms have “outer”, or “valence”, electrons. These electrons are much easier to remove than the “inner” or “core” electrons. An element will “want” to have eight valence electrons.

This provides some backup for the “octet rule”, which states that **an element will “want” to have 8 valence electrons, which is how many electrons the nearest noble gas has. It will lose or gain electrons in order to do this.** Elements with only a few valence electrons will tend to lose them and become like the nearest noble gas, while elements with close to 8 valence electrons will tend to gain electrons to become like the nearest noble gas.

Memorize this key fact about ionization energy/the periodic table: As you move from left to right across a row of the periodic table, the first ionization energy tends to **increase**.

Memorize this key fact about ionization energy/the periodic table: As you move down a column of the periodic table, the first ionization energy tends to **remain the same or decrease**.

Practice

Do some review. Write the number of valence electrons each element has below, then write how many electrons it “wants” to lose or gain in order to become like the nearest noble gas. Then write the charge that results.

Remember how to determine the number of valence electrons an element has by counting from left to right and ignoring columns 3-12. So, counting from left to right, we see that Boron has 3 valence electrons, Carbon and Silicon have 4 valence electrons, and so on. The nearest noble gas to Boron (atomic number 5) is Helium (atomic number 2).

Na

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

Ca

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

O

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

F

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

Cl

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

Kr

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

N

of valence electrons: _____

of electrons it wants to gain/lose (circle one): _____

Charge once it has gained/lost electrons: _____

Closing: Check your understanding of the lesson by articulating an answer to the following questions.

Use a periodic table to answer these questions (don't use the periodic table with the ionization energies on it).

Which has a higher ionization energy:

H or He?

Li or B?

K or Fe?

Pd or Rb?

Mg or Be?

Pb or C?

Chemistry LibreTexts has a decent page on this topic, which can be found at

- https://chem.libretexts.org/Courses/Oregon_Institute_of_Technology/OIT%3A_CHE_202_-_General_Chemistry_II/Unit_3%3A_Periodic_Patterns/3.3%3A_Trends_in_Ionization_Energy
- or by searching “Trends in Ionization Energy Chemistry LibreTexts” online.

Wednesday, March 25

Chemistry Unit: Electrons and orbitals

Lesson 3: Atomic radius and electron orbits

Unit Overview:

As you move down a column in the periodic table, the number of protons drastically increases, causing the atomic radius to increase. Since distance has an inverse and exponential effect on electrostatic attraction, this means the nucleus for these “bigger” atoms exert a very weak pull on their “outer” electrons. As such, in general the ionization energy of elements decreases as you move down any given column.

There are seven distinct energy levels which correspond to electron orbit paths around the nucleus. Electrons also are thought to take certain path shapes, including a spherical, “dumbbell”, four-leaf clover, and “double-pacifier” shape, among others. These are named “s, p, d, and f.”

Lesson 3 Socratic Question: Keep these questions in mind as you study this lesson!

How do electrons (valence and inner/core electrons) interact with the nucleus? Do they all clump around the nucleus, or do they fall into structured orbits and patterns around the nucleus?

Chemistry 10: Electrons and “orbitals”

March 23-27

Objective: Be able to do this by the end of this lesson.

Identify different orbital shapes and begin to name the four different “types” of orbital by shape.

Introduction to Lesson 3

You’ll remember from yesterday that, across a row, it becomes harder to remove an electron from an element because as you move across a row the number of protons increases. However, move down to the next row, and you’re dealing with a whole new set of lower numbers.

Take a look at the diagram from yesterday. Each element in the second highlighted row has a lower ionization energy, across the board, than the one above it.

Period

Group

First Ionization Energies of Some Elements (kJ/mol)

1	2	13	14	15	16	17	18
1	H 1310						He 2370
2	Li 520	Be 900					
3	Na 490	Mg 730					
4	K 420	Ca 590	Sc 630	Ti 660	V 650	Cr 660	Mn 710
5	Rb 400	Sr 550	Y 620	Zr 660	Nb 670	Mo 680	Tc 700
6	Cs 380	Ba 500	La 540	Hf 700	Ta 760	W 770	Re 760
7	Fr ...	Ra 510					

We’ll remember that this flies in the face of previous theory. Before, the theory was that **as the number of protons increases, the ionization energy increases**. However (to take an easy example), Rubidium (Rb) has more protons than Potassium (K). Yet Rubidium’s electrons are easier to kick off. Why is that?

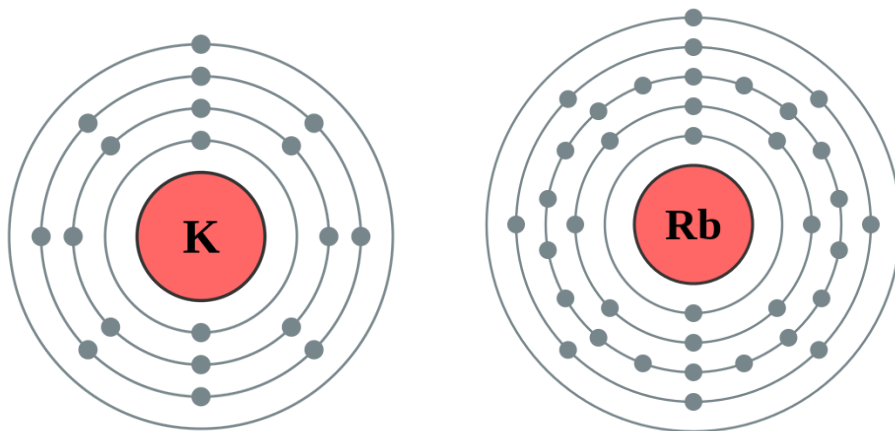
Write your answer here:

Today, we explain by saying that: as the number of protons and electrons increases, the **size** of the element’s atom gets larger. This means that Rb has not only more protons, but more electrons, and those electrons are more spread-out than potassium’s. Another way of saying this (because we view atoms as spherical) is by saying that **bigger atoms have a bigger atomic radius**. Rubidium is a bigger atom.

Because the valence (outer) electrons of Rubidium are so much farther from its nucleus, the positive charges of its protons exert a weaker pull on its outer electrons. Therefore, it’s easier to remove one

of Rubidium’s outer electrons when compared to one of Potassium’s outer electrons, which are much **closer** to its nucleus.

Let’s take that one step further. Not only do we think that electrons orbit at a distance from the nucleus, but we also think that they fall into predetermined distances, or levels, from the nucleus. Observe the Bohr diagrams below (named for Niels Bohr, a Danish physicist who first theorized these levels):



Notice Rubidium in this diagram has 5 “levels”, or layers, that its electrons fall neatly into. Potassium similarly has 4 “layers” or levels that contain its electrons.

How do we justify this diagram? Why can’t we just say that the electrons orbit the nucleus at **different** distances, and that those distances can vary? Why do we have to say electrons will fall into “level 1”, “level 2”, “level 3”, and so on?

In part because of the ionization energy data we’ve looked at earlier. Take a look at potassium, for instance, in the table below.

Table 3.3.1: Successive Ionization Energies for Selected Elements (kJ/mol)							
Element	IE ₁	IE ₂	IE ₃	IE ₄	IE ₅	IE ₆	IE ₇
K	418.8	3051.8	4419.6	5876.9	7975.5	9590.6	11343
Ca	589.8	1145.4	4912.4	6490.6	8153.0	10495.7	12272.9
Sc	633.1	1235.0	2388.7	7090.6	8842.9	10679.0	13315.0
Ga	578.8	1979.4	2964.6	6180	8298.7	10873.9	13594.8
Ge	762.2	1537.5	3302.1	4410.6	9021.4	?	?
As	944.5	1793.6	2735.5	4836.8	6042.9	12311.5	?

source: https://chem.libretexts.org/Courses/Oregon_Institute_of_Technology/OIT%3A_CHE_202_-_General_Chemistry_II/Unit_3%3A_Periodic_Patterns/3.3%3A_Trends_in_Ionization_Energy

One way of reading this table is to look at the ionization energies listed for each element. “ IE_1 ” indicates the amount of energy required to remove the element’s first electron, “ IE_2 ” indicates the energy required to remove the element’s second electron, and so on.

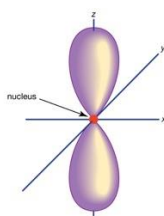
You’ll notice that, as expected, the more electrons you take from Potassium, the harder it gets to remove the next one. However, these numbers are true for **all potassium atoms**. They aren’t just taken from a single sample. These numbers are true for Potassium **wherever in the world you find it**.

The same goes for the numbers applying to Calcium, Scandium, Gallium; if we could provide data for the other elements, you’d see the same trend. Ionization energies for a particular electron on a particular element are always the same, no matter where in the universe you find that element.

Because of this, we say that Potassium and all elements must have **distinct** (we’ll also say “discrete” (which is not the same as “discreet”))—energy levels (as opposed to vague or relative). You can **measure** the energy of each electron level—they aren’t just relative “this one comes first, this one comes second, etc” levels. They have **known** energy requirements associated with them that you can use to determine what it takes to “kick off” the electrons in that layer.

So, what are these layers?

Well, based on the ionization energy data, it seems that there are 7 “layers”. Additionally, electrons don’t orbit the nucleus in neat ellipses or circles like planets around the sun. They seem to take non-elliptical orbit paths, like the one below:



In this orbital path, the electrons take a kind of 3-D-infinity sign, or dumbbell shape. This orbit path type is called a “p orbital”.

Electrons in each layer can be found orbiting in “s”, “p”, “d”, or “f” shapes.¹ The first layer can only hold 2 electrons and only the “s” shape is found in it. The second layer, however, can hold 8 electrons, and has “s” and “p” shapes. The third layer can hold 18 electrons and has “s”, “p”, and

¹ If you’re wondering why the shapes are called s, p, d, and f, the answer is a bit involved. The short answer is that in determining these shapes, physical chemists (read: chemists who do physics or physicists who do chemistry) would burn different elements and observe their line spectra (the types of visible light they released). This told them information about the electrons’ orbit paths. They named the orbitals s, p, d, and f for “sharp”, “principal”, “diffuse”, and “fundamental”. As I haven’t yet understood how they did these experiments, I also can’t right now why these names were given to these shapes.

“d” shapes. I’d like to iterate that this changes nothing about your previous knowledge of the electron. Everything you’ve learned is still true and relevant. This is just a bit deeper.

Closing: Check your understanding of the lesson by articulating an answer to the following questions.

1. How many distinct energy levels (“layers”) do electrons fall into? _____
 2. What are the four known orbit shapes that electrons can take within these layers? (simply name them) _____
 3. Using the term “atomic radius”, explain why ionization energy decreases as you move down a column in the periodic table.
-
-

Thursday, March 25

Chemistry Unit: Electrons and orbitals

Lesson 4: Writing “electron configurations”

Unit Overview

Electrons orbit the nucleus at specific distances which correspond to energy “levels” and move in a variety of orbital shapes, named “s”, “p”, “d”, and “f”.

- The biggest elements have electrons on all 7 energy levels, and the smallest elements only have enough electrons to have the first energy level.
- The “s” shape can hold 2, the “p” can hold 6, the “d” can hold 10, and the “f” can hold 14 electrons.

The first energy “level” can only hold 2 electrons, which take the shape of an “s” orbital.

The second energy “level” can only hold 8 electrons, which take two shapes: the “s” orbital (2 electrons) and the “p” orbital (6 electrons)

The third energy “level” can also only hold 8 electrons, which take two shapes: the “s” orbital (2 electrons) and the “p” orbital (6 electrons)

Lesson 4 Socratic Question: Keep these questions in mind as you study this lesson!

(same as yesterday)

How do electrons (valence and inner/core electrons) interact with the nucleus? Do they all clump around the nucleus, or do they fall into structured orbits and patterns around the nucleus?

Objective: Be able to do this by the end of this lesson.

Be able to identify, based on the number of electrons possessed, the specific energy level(s) and orbital(s) filled by an element in the first 3 rows of the periodic table.

Introduction to Lesson 3

We’ll just be looking at the first three rows of the periodic table today.

Chemistry 10: Electrons and “orbitals”

March 23-27

1

IA

1A

1

H

Hydrogen

1.008

2

IIA

2A

3

Li

Lithium

6.941

4

Be

Beryllium

9.012

11

Na

Sodium

22.990

12

Mg

Magnesium

24.305

19

K

Potassium

39.098

20

Ca

Calcium

40.078

37

Rb

Rubidium

85.468

38

Sr

Strontium

87.62

55

Cs

Cesium

132.905

87

Fr

Francium

223.028

2

He

Helium

4.003

10

Ne

Neon

20.180

18

Ar

Argon

39.948

36

Kr

Krypton

84.798

54

Xe

Xenon

131.294

86

Rn

Radon

222.018

118

Uuo

Ununoctium

unknown

13

B

Boron

10.811

14

C

Carbon

12.011

15

N

Nitrogen

14.007

16

O

Oxygen

15.999

17

F

Fluorine

18.998

53

I

Iodine

126.904

81

Tl

Thallium

204.383

82

Pb

Lead

207.2

83

Bi

Bismuth

208.980

84

Po

Polonium

[209]

85

At

Astatine

[210]

86

Rn

Radon

[222]

113

Uut

Ununtrium

unknown

114

Fl

Flerovium

[289]

115

Uup

Ununpentium

unknown

116

Lv

Livermorium

[293]

117

Uus

Ununseptium

unknown

118

Uuo

Ununoctium

unknown

21

Sc

Scandium

44.956

22

Ti

Titanium

47.867

23

V

Vanadium

50.942

24

Cr

Chromium

51.996

25

Mn

Manganese

54.938

26

Fe

Iron

55.845

27

Co

Cobalt

58.933

28

Ni

Nickel

58.693

29

Cu

Copper

63.546

30

Zn

Zinc

65.38

31

Ga

Gallium

69.723

32

Ge

Germanium

72.631

33

As

Arsenic

74.922

34

Se

Selenium

78.972

35

Br

Bromine

79.904

41

Nb

Niobium

92.906

42

Mo

Molybdenum

95.95

43

Tc

Technetium

98.907

44

Ru

Ruthenium

101.07

45

Rh

Rhodium

102.906

46

Pd

Palladium

106.42

47

Ag

Silver

107.868

48

Cd

Cadmium

112.411

49

In

Indium

114.818

50

Sn

Tin

118.711

51

Sb

Antimony

121.760

52

Te

Tellurium

127.6

72

Hf

Hafnium

178.49

73

Ta

Tantalum

180.948

74

W

Tungsten

183.84

75

Re

Rhenium

186.207

76

Os

Osmium

190.23

77

Ir

Iridium

192.217

78

Pt

Platinum

195.085

79

Au

Gold

196.967

80

Hg

Mercury

200.592

104

Rf

Rutherfordium

[261]

105

Db

Dubnium

[262]

106

Sg

Seaborgium

[266]

107

Bh

Bohrium

[264]

108

Hs

Hassium

[269]

109

Mt

Meitnerium

[268]

110

Ds

Darmstadtium

[271]

111

Rg

Roentgenium

[272]

112

Cn

Copernicium

[277]

57-71

Lanthanide Series

89-103

Actinide Series

Alkali Metal

Alkaline Earth

Transition Metal

Basic Metal

Semimetal

Nonmetal

Halogen

Noble Gas

Lanthanide

Actinide

Periodic Table of the Elements

Atomic Number

Symbol

Name

Atomic Mass

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Recall the four types of orbital “shape” from yesterday.

The “s” orbital shape can sustain 2 electrons.

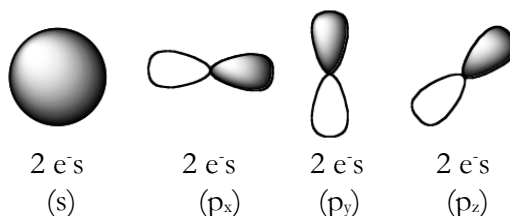
The “p” orbital shape can sustain 6 electrons.

The “d” orbital shape can hold 10 electrons.

The “f” orbital shape can hold 14 electrons.

Therefore, if an element has 8 electrons (in this case that would be Oxygen), it should have one “s” orbital and one “p” orbital. Two of its electrons are orbiting in a “s” and the remaining six of its electrons are orbiting in a “p” shape.

What does that look like? Let’s draw a picture.



Notice that 3 “p” shapes are drawn. They all look the same, but they can be oriented 3 ways (along the x-axis horizontally, along a y-axis vertically, or along a z-axis coming out of the page). This is due entirely to their shape. The “s” orbital by a similar rule cannot be oriented any other way, because

Chemistry 10: Electrons and “orbitals”

March 23-27

it's a sphere, and this is actually why the “s” orbital can only hold 2 electrons, while the “p” orbital can hold 6 (2 in each orientation).

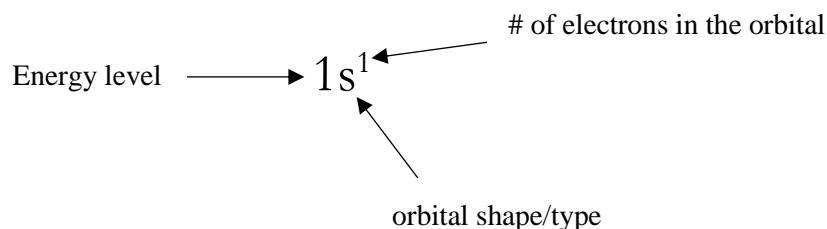
Memorize the important fact that electrons orbit the nucleus at specific distances which correspond to energy “levels” and move in a variety of orbital shapes, named “s”, “p”, “d”, and “f”.

- The biggest elements have electrons on all 7 energy levels, and the smallest elements only have enough electrons to have the first energy level.
- The “s” shape can hold 2, the “p” can hold 6, the “d” can hold 10, and the “f” can hold 14 electrons.

Electron Configurations

Let's now practice writing electron configurations. This is a helpful way of visualizing this in a text-based way that will relate directly to the pictures you've seen above.

The electron configuration for Hydrogen is



There are 7 energy levels. Take a look at the diagram below:

Electron Configurations in the Periodic Table									
1 H 1s								2 He 1s	
3 Li 2s	4 Be		5 B	6 C	7 N	8 O	9 F	10 Ne	
11 Na 3s	12 Mg		13 Al	14 Si	15 P	16 S	17 Cl	18 Ar	

The elements in the first row are in the 1st “level”. The elements in the second row are in the second electron orbit “level”. The elements in the third row are in the 3rd orbit “level”.

Try writing the electron configuration for Helium:

Helium: _____
electron configuration

Check your answer: The electron configuration for Helium is $1s^2$. Check the labeled diagram on the previous page if you're confused.

Practice Problems

Now try writing the electron configurations for these different elements in the first three rows of the periodic table.

Lithium: _____
 electron configuration

Carbon: _____
 electron configuration

Fluorine: _____
 electron configuration

Neon: _____
 electron configuration

Magnesium: _____
 electron configuration

Argon: _____
 electron configuration

Answers are written below.

Lithium: $1s^22s^1$

Carbon: $1s^22s^22p^2$

Fluorine: $1s^22s^22p^5$

Neon: $1s^22s^22p^6$

I think you're probably noticing that this becomes onerous, especially for big elements. There's a shortcut we'll look at tomorrow you can (and probably should) use when writing electron configurations for elements in the 2nd row or further down. I'll write both possible answers for the last two.

Magnesium: $1s^22s^22p^63s^2$ or $[\text{Ne}]3s^2$

Argon: $1s^22s^22p^63s^23p^6$ or $[\text{Ne}] 3s^23p^6$

Closing: Check your understanding of the lesson by articulating an answer to the following questions.

Write the electron configuration of the following elements.

Beryllium: _____
 electron configuration

Chemistry 10: Electrons and “orbitals”

March 23-27

Aluminum: _____
electron configuration

Silicon: _____
electron configuration

Carbon: _____
electron configuration

Challenge:

N^{3-} (nitrogen with three extra electrons!): _____
electron configuration

Na^+ : _____
electron configuration

Cl: _____
electron configuration

Ar: _____
electron configuration

Friday, March 27

Chemistry Unit: Electrons and orbitals

Lesson 2: Writing electron configurations (cont'd)

Unit Overview

Orbital Shapes

- The s orbital takes a spherical shape, while the p orbital takes a sort of “dumbbell” shape. The p orbital can be arranged in space 3 different ways. Since each arrangement can hold 2 electrons, the p orbital can hold 6 total electrons.
- The d orbital can take 2 different shapes which can be arranged in space in a total of 5 different ways. The first shape generally resembles a four-leaf clover, while the second shape has been said to resemble a “double-sided pacifier”² Since any spatial orbital arrangement can hold 2 electrons, this helps explain why the d orbital is said to hold 10 electrons.
- The f orbital can take 2 different shapes which can be arranged in space in a total of 7 different ways. The first shape is hard to describe, but the second shape can be described as “two pacifiers, back to back”. Since any spatial orbital arrangement can hold 2 electrons, this helps explain why the f orbital is said to hold 14 electrons.

Electron Configurations

- When writing electron configurations, they can be shortened by substituting a noble gas in brackets [] to replace the part that corresponds to that noble gas’ configuration

² J. Castillo, 2019.

- When writing an electron configuration for an **ion**, you must first consider the electron configuration of the **uncharged species**, then **add** or **subtract** electrons from the electron configuration to write the ion's electron configuration.

Lesson 5 Socratic Question: Keep these questions in mind as you study this lesson!

(same as yesterday)

How do electrons (valence and inner/core electrons) interact with the nucleus? Do they all clump around the nucleus, or do they fall into structured orbits and patterns around the nucleus?

Objective: Be able to do this by the end of this lesson.

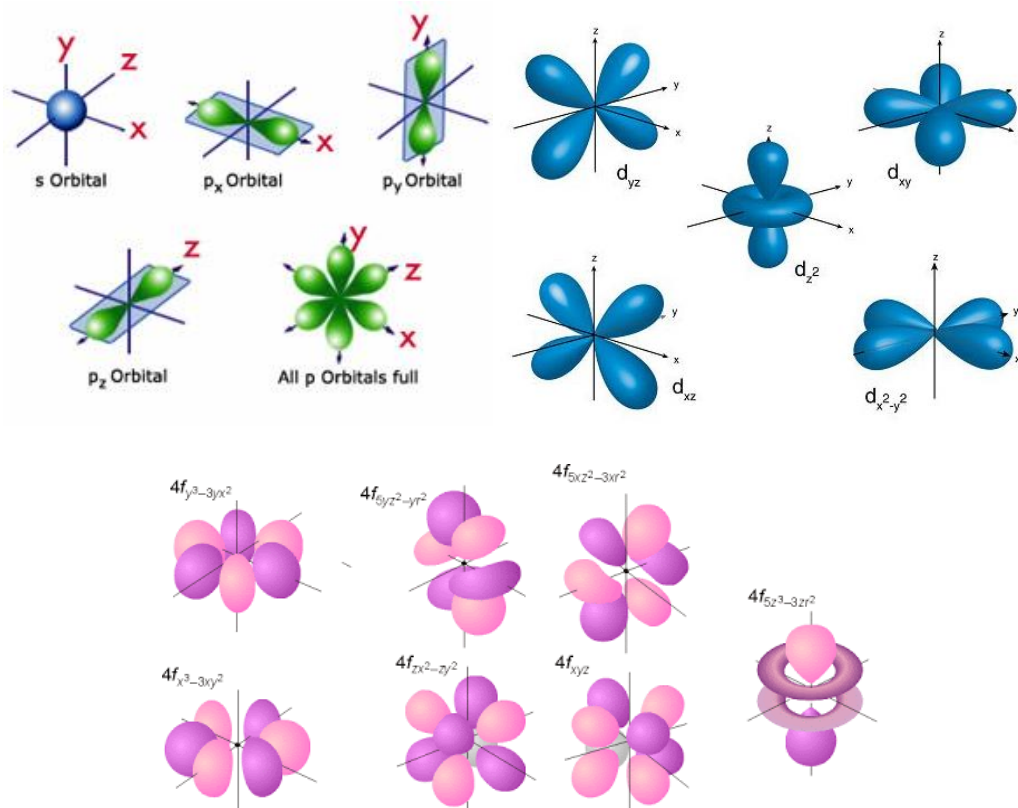
Identify different orbital shapes and begin to name the four different “types” of orbital by shape.

Be able to write an electron configuration for any element or ion, using noble-gas notation to abbreviate the configurations of big elements.

Introduction to Lesson 5

Today's lesson will be a bit simpler.

Here, take a look at all the orbital shapes:



Take some time to describe each type of orbital shape. Notice that the d and f orbital shapes include one very unusual-looking variation. Make sure to describe that as well as the other, “typical” d and f orbital shapes.

March 23-27

f shape description: _____

Today we'll just be practicing writing electron configurations for different elements. This means you'll have to know the order in which electron "orbitals" fill up. The energy levels were determined experimentally, so it's not exactly intuitive. Take a look at this periodic table, in which the orbital levels and shape-types have been labeled:

Electron Configurations in the Periodic Table																	
1 H 1s																	2 He 1s
3 Li 2s	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na 3s	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K 4s	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
37 Rb 5s	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
55 Cs 6s	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
87 Fr 7s	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	113	114				
<div> <div>58 Ce</div> <div>59 Pr</div> <div>60 Nd</div> <div>61 Pm</div> <div>62 Sm</div> <div>63 Eu</div> <div>64 Gd</div> <div>65 Tb</div> <div>66 Dy</div> <div>67 Ho</div> <div>68 Er</div> <div>69 Tm</div> <div>70 Yb</div> <div>71 Lu</div> </div>																	
<div> <div>90 Th</div> <div>91 Pa</div> <div>92 U</div> <div>93 Np</div> <div>94 Pu</div> <div>95 Am</div> <div>96 Cm</div> <div>97 Bk</div> <div>98 Cf</div> <div>99 Es</div> <div>100 Fm</div> <div>101 Md</div> <div>102 No</div> <div>103 Lr</div> </div>																	

by: Sarah Faizi

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Notice that if you go by increasing atomic number (increasing number of protons and electrons), moving left to right and top to bottom: first the 1s orbital fills up, then the 2s, then the 2p (B thru Ne), then the 3s, 3p; then 4s, 3d, 4p; then 5s, 4d, 5p; then the 6s, 4f, 5d, and 6p; then the 7s, 5f, and 6d. Notice how the row does not correspond to the energy level for the d and the f orbital

shapes. The first “d” shape level is on the same energy level as 3s, but it is in the same row as 4s. You’ll see a similar trend with the “f” orbitals. The first “f” shape level is on the same level as 4s and 4p, but it’s in the same row as 6s. This is confusing. We know that the “d” orbitals are on the 3rd, 4th, 5th, and 6th energy levels because of experimental data. They exist at the same distance from the nucleus as electrons in the 3rd, 4th, 5th, and 6th energy levels; however, when an element gets more electrons, the 3d orbital fills after the 4s orbital. There is an element of “this is simply how this works” at play here, but while this is a bit confusing, we have faith that the universe is ordered and knowable. Though we don’t understand *why* the 3d level fills up after the 4s and not after the 3s, we will be constantly looking for reasons, trusting that they exist.

After that long prelude, try your hand at these practice problems. Today will be a more leisurely day.

Practice

We’ll add in two caveats. The first is that of noble-gas configurations. Using noble-gas configurations to abbreviate is a really practical way of writing electron configurations. So, for example, from yesterday’s conclusion:

If you write the electron configuration of He, it would be $1s^2$. When writing the electron configuration for Li, you *could* write it as $1s^2 2s^1$, because Lithium has three total electrons (two in the 1s orbital, one in the 2s orbital), but you could *also* write it as: $[\text{He}] 2s^1$. Since Helium’s electron configuration is $1s^2$, we substitute [He] for “ $1s^2$ ”. This saves a lot of space for the larger elements. For instance:

When writing the electron configuration for neon, you’d write it as: $1s^2 2s^2 2p^6$ (or $[\text{He}] 2s^2 2p^6$). When writing the electron configuration for Magnesium, you’d write it as $[\text{Ne}] 3s^2$.

This saves space. Otherwise you’d have written $1s^2 2s^2 2p^6 3s^2$. Obviously this is even more helpful in the 4th, 5th, 6th, etc rows of the periodic table.

Practice problems:

Write the electron configuration (using noble-gas names in brackets to abbreviate) for each element.

Fe:	_____
	electron configuration
Rubidium:	_____
	electron configuration
Pd:	_____
	electron configuration
Osmium:	_____
	electron configuration
Bismuth:	_____
	electron configuration
Calcium:	_____
	electron configuration

Chemistry 10: Electrons and “orbitals”

March 23-27



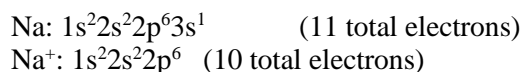
K:	_____
	electron configuration
Cs:	_____
	electron configuration
Californium:	_____
	electron configuration
Ge:	_____
	electron configuration
I:	_____
	electron configuration
Uranium:	_____
	electron configuration

The second caveat is this: When writing an electron configuration for an **ion**, you must first consider the electron configuration of the **uncharged species**, then **add** or **subtract** electrons from the electron configuration to write the ion's electron configuration.

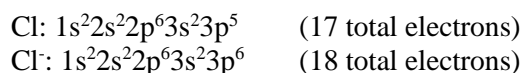
For example:

Sodium (Na)'s electron configuration is $[\text{Ne}]3s^1$. Sodium likes to become Na^+ . Therefore, the electron configuration of Na^+ is simply $[\text{Ne}]$.

In other words:



It works the opposite way for anions. From yesterday's challenge:



(I could also have written these as $[\text{Ne}]3s^23p^5$ and $[\text{Ne}]3s^23p^6$, respectively)

Okay, now try some electron configurations for ions.

Li^+ :	_____
	electron configuration
Ca^{2+} :	_____
	electron configuration
N^{3-} :	_____
	electron configuration
Fe^{2+} :	_____
	electron configuration
Ag^+ :	_____
	electron configuration
Au^+ :	_____

O²⁻: electron configuration

H⁺: electron configuration

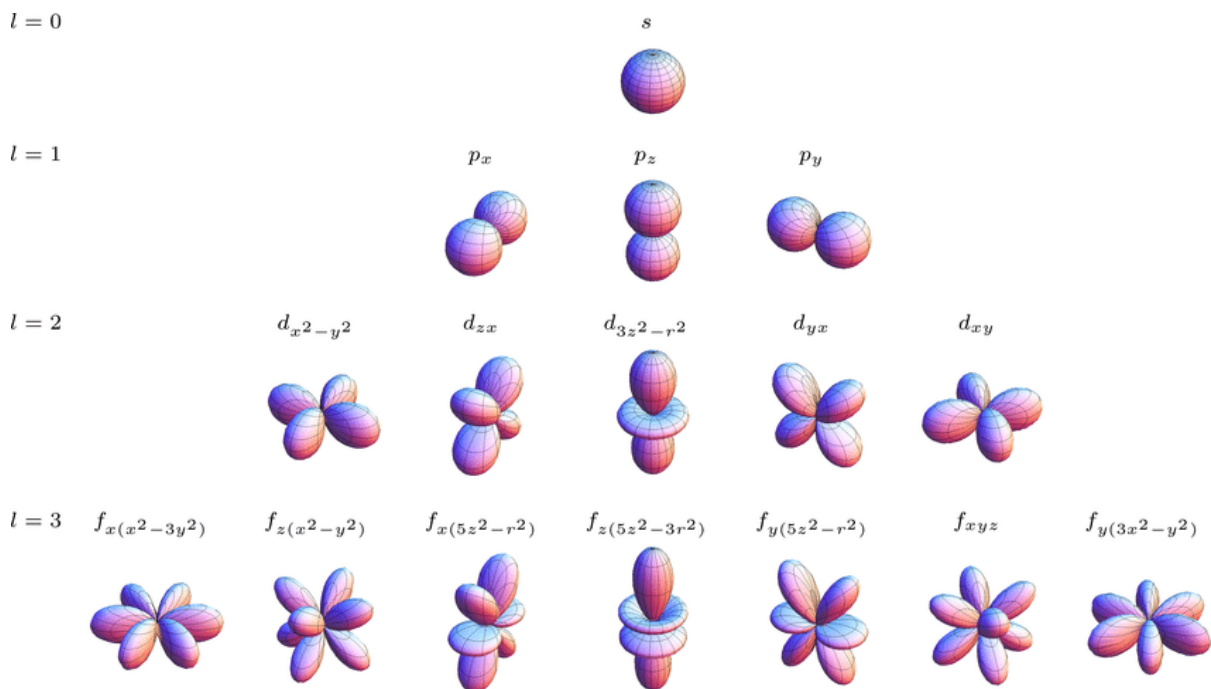
Okay, that last one was mean. Try this one:

H: electron configuration

Moving Forward

As you review your work before returning this packet, here are a couple of resources you might find handy:

1. A diagram of all the orbital types.



2. A handy way of remembering the order in which orbitals fill up (without using a periodic table!). I have attempted to describe this diagram in text, but have failed. Take a close look at it and see how it works. If you like it, use it. Commit it to memory.

