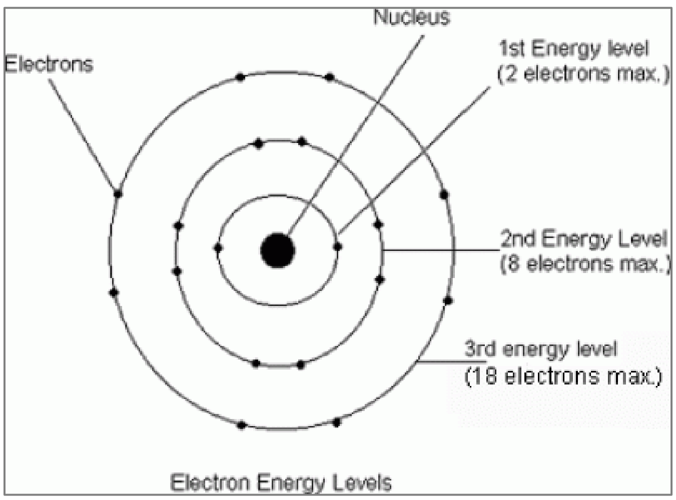
Day 1:

Bohr Models and Energy Levels

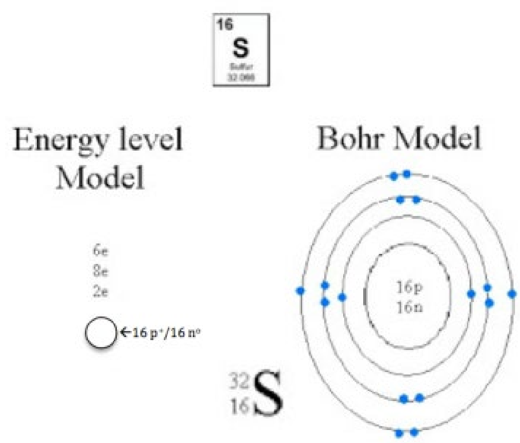
On the last page of this packet a Periodic Table has been provided for your use

The Bohr model of the atom is one example of an atomic model that helps students to understand some useful information about atomic structure, even though it also contains information that we now know is invalid.

The key idea in the Bohr model of the atom is that electrons occupy definite orbits that require the electron to have a specific amount of energy. Orbits closer to the nucleus would require the electrons to have a smaller amount of energy, and orbits farther from the nucleus would require the electrons to have a greater amount of energy. The possible orbits are known as **energy levels**. Bohr worked out the rules for the maximum number of electrons that could be in each energy level in his model (see image below, left). In the normal state, termed the ground state, this would require the atom to have all of its electrons in the lowest energy levels available.



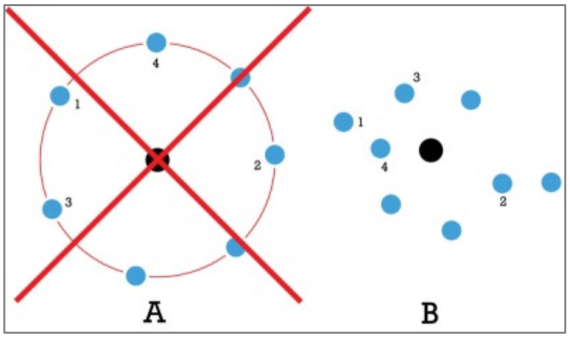
Two models to represent atoms are shown below:



Both the energy level and planetary models have shortcomings, but are both useful to the modern understanding of the atom. For example, in both models the nucleus is too large. Real atoms are mostly empty space. Also, in reality, scientists cannot tell exactly where an electron is at a given moment or where it is going. Showing the electrons in defined energy levels and at defined locations is inaccurate. Scientists can calculate the probability of where to find the electrons, but that is not the same as knowing where an electron is located. See more information below on these shortcomings.

Shortcomings of Bohr's theory

Bohr's system was only successful for atoms that have a single electron, which meant that the Bohr model did not accurately reflect the behaviors of most atoms. Another problem with Bohr's theory was that the Bohr model did not explain why certain energy levels existed. Yet another problem with the Bohr model was the predicted positions of the electrons in the electron cloud. If Bohr's model were correct, the electron in the hydrogen atom in the ground state would always be the same distance from the nucleus. Although the actual path that the electron followed could not be determined, scientists were able to determine the positions of the electron at various times. If the electron circled the nucleus as suggested by the Bohr model, the electron positions would always be the same distance from the nucleus. In reality, the electron is found at many different distances from the nucleus. In the figure below, the left side of the image (labeled as A) shows the positions an electron would occupy as predicted by the Bohr model, while the right side (labeled as B) shows some actual positions of an electron.



The Bohr model was not, however, a complete failure. It provided insights that triggered the next step in the development of the modern concept of the atom, the Quantum Mechanical model. Second, the Bohr model provides a simplistic means for visualizing the placement of the electrons in defined energy levels, a significant consistency with the modern concept of the atom.

Questions:

1. What aspects of Bohr models are useful?

2. List at least 3 ways are Bohr models incorrect. Explain why each is incorrect.

3. If historical models are found to be partly or completely incorrect, why do teachers task students to learn about them?

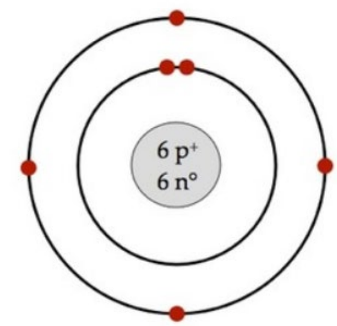
Day 2

Bohr Diagrams

In the previous lesson we reviewed some of the reasons that Bohr models, while flawed, can be useful to developing an understanding of energy levels in atoms, a key concept in understanding the reactivity of atoms of different elements. Now, we will practice drawing Bohr diagrams for elements in periods 1-3, and draw conclusions about an important pattern on the periodic table.

Drawing Bohr Diagrams (Bohr models) Example below.

1. Use your Periodic Table (provided on the last page of this packet) to locate the element you're going to model
2. Determine the number of electrons the element has as a neutral atom
3. Determine the number of energy levels to draw by finding out what period the element is in. The period is the number of energy levels you need to draw.
4. Write the element's symbol to represent the nucleus of your atom, and the number of protons and neutrons in the atom. You can use "p^+" to represent protons, "n^0" to represent neutrons.
5. Draw the correct number of circles around the nucleus to represent energy levels
6. Draw a dot (or an "e^-") to represent each electron in the atom, starting with the lowest energy level and filling each level as you go, until you have the correct number of electrons.
   - Remember the lowest energy level can hold up to 2 electrons.
   - The second energy level can hold up to 8 electrons.
   - The third energy level can hold up to 18 electrons.



Carbon

Example:

- ❖ Carbon has 6 protons, 6 neutrons, 6 electrons
- ❖ It has 2 energy levels (it's in the second period)
- ❖ The first 2 electrons are drawn in the lowest energy level, that level is now full
- ❖ There are 4 additional electrons, they are all in the second energy level.
- ❖ Electrons are placed apart from each other in the diagram because "like charges repel" each other.
- ❖ For elements with more electrons, complete the second energy level (8 electrons) then add any additional electrons to the third energy level.

Now it's your turn to practice. Draw the Bohr models for elements 1 through 18.

1- Hydrogen	• If you're completing the models on a computer, you can draw energy levels and electrons by using the "insert shape" tool in MS Word, or the "insert drawing" tool in Google docs. • If these options don't work for you, you can draw the diagrams by hand on paper and take a photo of them if you're submitting electronically or turn them in on paper to your teacher.						2- Helium
3- Lithium	4- Beryllium	5- Boron	6- Carbon	7- Nitrogen	8- Oxygen	9- Fluorine	10- Neon
11-Sodium	12- Magnesium	13- Aluminum	14- Silicon	15-Phosphorous	16- Sulfur	17- Chlorine	18- Argon

Analysis: Claim, Evidence, Reason

Claim: Write a claim to answer the question: What patterns/ trends do you see regarding the Bohr models you've drawn?

Evidence: What evidence do you see in your Bohr models that supports your claim? (Be sure to state what specific things you see)

Reason: How does your evidence support your claim? What logical conclusion can you draw from your models?

Connections/ Prediction: How do you think the pattern or trend(s) is/are related to other important concepts in Chemistry?

Day 3

Ions

Now that we've reviewed a bit about atomic structure, let's recall what ions are and how ions are formed. Atomic structure plays a critical role in the formation of ions (and chemical reactions).

Some important terms to remember for understanding how ions are formed:

Valence electrons: the outermost electrons in an atom. By counting across the groups in the periodic table, you can figure out how many valence electrons each atom has. For example, calcium has two valence electrons and nitrogen has five valence electrons. Elements in the same group of the periodic table have the same number of valence electrons, which explains why they have similar properties. So, why do we care about valence electrons anyway? It's the valence electrons that are used in forming and breaking chemical bonds. Since this making and breaking of bonds is something that we find interesting, we, in turn, find valence electrons to be interesting, too.

The Octet Rule: The octet rule states that atoms tend to gain or lose electrons so they can have the same number of electrons as the nearest noble gas. In other words, *the atom will gain or lose electrons so that its outer energy level is full*.

Here are some things that you might want to keep in mind about this unbelievably important rule: The nearest noble gas may be either before the element (just as neon is before aluminum) or after the element (as argon is after phosphorus). Remember your Bohr models? The first energy level can hold up to 2 electrons, therefore if the element is closest to Helium on the Periodic Table, it will tend to gain or lose electrons to end up with 2 valence electrons. If it is closest to Neon, it will tend to gain or lose electrons to get to 8 valence electrons.

If the nearest noble gas is *before* the element of interest, *the element will lose valence electrons* until it has the same number of valence electrons as that noble gas. This will give it a *positive charge*. As a result, aluminum will lose three electrons to become like neon, giving it a +3 charge.

If the nearest noble gas is *after* the element of interest, the element will *gain valence electrons* until it has the same number of valence electrons as that noble gas. This will give it a *negative charge*. As a result, phosphorus will gain three electrons to become like argon, giving it a -3 charge. Ignore the d- and f-blocks on the periodic table when counting forwards and backwards. In this way, gallium (Ga) wants to lose three electrons to become like argon – we just ignore all of the transition elements that are between them.

Ions: Atoms that have gained or lost electrons are ions. *Cations* are atoms that have *lost electrons*, and therefore have a *positive charge* (like Al^{+3}). *Anions* are atoms that have *gained electrons*, and therefore have a *negative charge* (like O^{-2}). Ions are important in some of the most common types of chemical reactions.

Questions:

1. How can you determine how many electrons an atom will gain or lose when it becomes an ion?

2. Why do atoms that gain electrons end up with a negative charge while atoms that lose electrons end up with a positive charge?

3. Based on your understanding of valence electrons and the octet rule, do elements that are Noble Gases tend to become ions? Why or why not?

4. Draw Bohr models for ions of elements in the table below. You do not have to draw models for the darkly shaded cells.

1- Hydrogen ion	● If you're completing the models on a computer, you can draw energy levels and electrons by using the "insert shape" tool in MS Word, or the "insert drawing" tool in Google docs. ● If these options don't work for you, you can draw the diagrams by hand on paper and take a photo of them if you're submitting electronically or turn them in on paper to your teacher.						2- Helium
3- Lithium ion	4- Beryllium ion				8- Oxygen ion	9- Fluorine ion	10- Neon
11-Sodium ion	12- Magnesium ion				16- Sulfur ion	17- Chlorine ion	18- Argon

5. Compare your Bohr models for the neutral atoms in Lesson 10 to your models for the ions in the table above. What patterns do you see?

<u>Analysis</u>: Claim, Evidence, Reason

Claim: Based on your models and your reading from the previous review lessons, write a claim to answer the question: How are valence electrons related to the reactivity of elements in different families?

Evidence: What evidence do you see in your Bohr models and from your reading that supports your claim?

Reason: How does your evidence support your claim? What logical conclusion can you draw?

Day 4

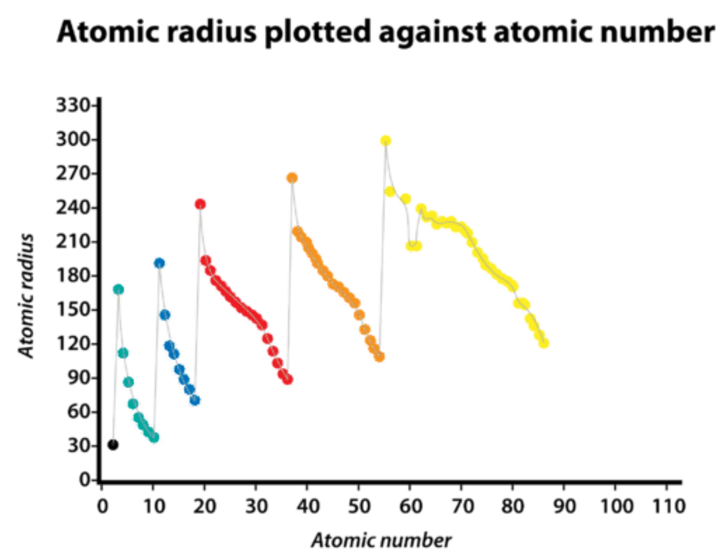
How are Atomic Structure and Periodic Trends related? Part 1, atomic structure

In the next few lesson activities we're going to draw connections between your understanding of the structure of atoms and the periodic trends that have been reviewed so far in Lessons 1-11 in order to provide depth to your understanding of one of the main driving forces of chemical reactions. We'll start with the relative sizes of atoms of different elements.

Recall from review Lesson 1, "The charges in the atom are crucial in understanding how the atom works. An electron has a negative charge, a proton has a positive charge and a neutron has no charge. Electrons and protons have the same magnitude of charge. Like charges repel, so protons repel one another as do electrons. Opposite charges attract which causes the electrons to be attracted to the protons. As the electrons and protons grow farther apart, the forces they exert on each other decrease."

In Lesson 7, you reviewed some important trends on the Periodic Table of Elements, including the trend for the atomic radii of atoms. Recall: "The atomic radius is half of the distance between nuclei in covalently bonded diatomic molecules". In other words, it's a measure of the size of atoms. It may seem that the trend in the size of atoms is a simple concept, but the reason for those sizes (the number, location and charges of their protons, neutrons and electrons) affects how atoms react and bond!

Review two visual representations of the sizes of atoms below.



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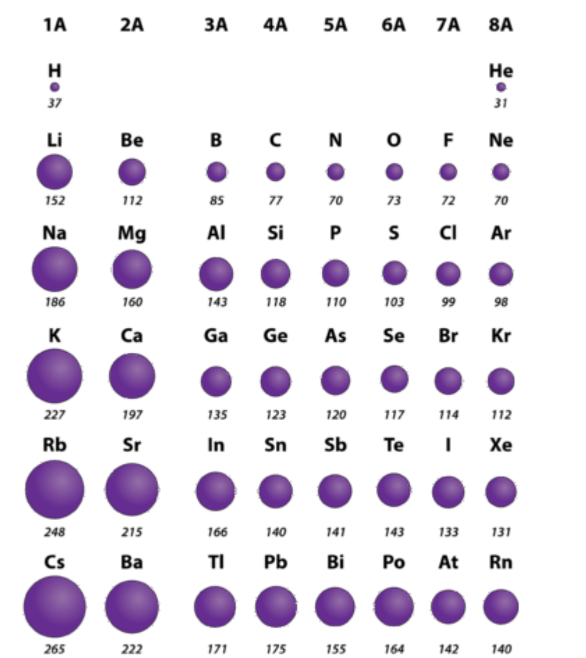


*Atomic radii in pm

Questions/ Connections:

1. Using what you know about subatomic particles, energy levels, and the visual representations above, what causes the atomic radii to decrease from left to right across the Periodic Table?

2. What causes the atomic radii to increase within a group on the Periodic Table?

Day 5

How are Atomic Structure and Periodic Trends related?

Part 2, Graphing and Analyzing Data

Recall, from Lesson 7, "**Ionization Energy** is the energy required to remove an electron from an atom". In this lesson you are going to graph the ionization energies of atoms, and to interpret the trends from the graph.

Directions:

1. Graph the data on grid paper.You may use the graph/grid provided on the following page, or you may graph the data using MS Excel, Desmos, etc.
2. The atomic number of the element will be plotted on the x-axis and the Ionization Energy will be plotted on the y-axis.
3. Be sure to choose an appropriate domain (x-axis) and range (y-axis).
    a. Numbers on the x-axis will go from 0-20, labeled in equal increments
    b. Numbers on the y-axis will go from 0-2500, labeled in equal increments (choose an appropriate scale)
4. Include axis labels for both the x and y axes, including units.
5. Include an appropriate title

Table 1: First Ionization Energies of Elements (the amount of energy to remove the first electron from an atom)

Atomic Number	Ionization Energy (kj/mol)		Atomic Number	Ionization Energy (kj/mol)
1	1312		11	495
2	2372		12	737
3	520		13	577
4	899		14	786
5	800		15	1011
6	1086		16	1000
7	1402		17	1251
8	1313		18	1520
9	1681		19	419
10	2080		20	590

Questions:

1. What patterns/ trends do you see on the graph?

2. How do the patterns/trends relate to the periodic table?

3. How do the patterns/trends relate to your understanding of atomic structure?

Day 6

How are Atomic Structure and Periodic Trends related? Part 3

These items provide an example of science test questions you may encounter on future standardized exams. You may use the periodic table included on the last page of this packet to answer the following items.

Ronda learned about the reactivity of metals in the first column of the periodic table. Her teacher demonstrated that the same amount of sodium (Na) placed in water is much more reactive than lithium (Li), which just fizzes-up a little (images below)

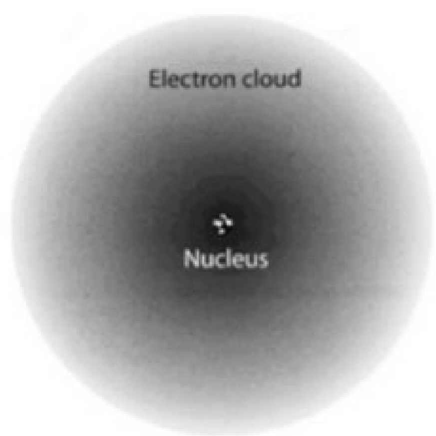
Note: the left picture is sodium (Na) and the right picture is Lithium (Li).

Images from *https://chemdemos.uoregon.edu/*

Ronda wondered why, at the atomic level, the reaction of sodium and water was so much more violent than that of lithium and water. To help her explain this, let's begin with an atomic-level representation. Below left is a representation of sodium.

To the right of this representation **draw** an atomic-level picture of lithium that shows the size of lithium relative to sodium.

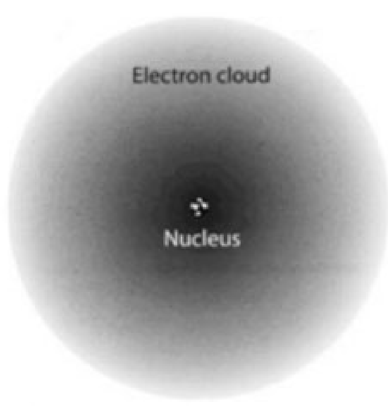
Sodium (Na)	Lithium (Li)
	

Predict whether the energy required to remove an electron from lithium would be more or less than that required to remove an electron from sodium. **Explain** your prediction in terms of the forces and interactions involved in this process. The representation you constructed in part A should be helpful here.

The relative ease with which an electron is transferred away from a metal determines how reactive that metal is with water. Given this, **predict** whether rubidium (Rb) would be more or less reactive in water than sodium. **Explain** your prediction in terms of the forces and interactions involved in removing an electron from an atom.

Ronda placed a piece of magnesium (Mg) in a beaker with water. There was no immediate, visible change. She wondered why the sodium was reactive in water, while magnesium, which is only one place to the right of it in the periodic table, did not seem to react when placed in a beaker of water. To help Ronda think about this, let's again begin with an atomic-level representation.

Below is a depiction of sodium. To the right of this representation, **draw** an atomic-level picture of magnesium that shows the size of magnesium relative to sodium.

Sodium (Na)	**Magnesium**
	

Bearing in mind that metal reactivity with water is determined by the ease of electron removal from that metal, **provide an atomic-level explanation** for why magnesium is less reactive than sodium when placed in water.

Explain why the relative size of the magnesium atom you drew is reasonable. Include the forces and interactions that govern the size of an atom in your answer.

Periodic Table

Group → / Period ↓	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
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
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
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
7	87 Fr	88 Ra	89 Ac **	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
			*	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	

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