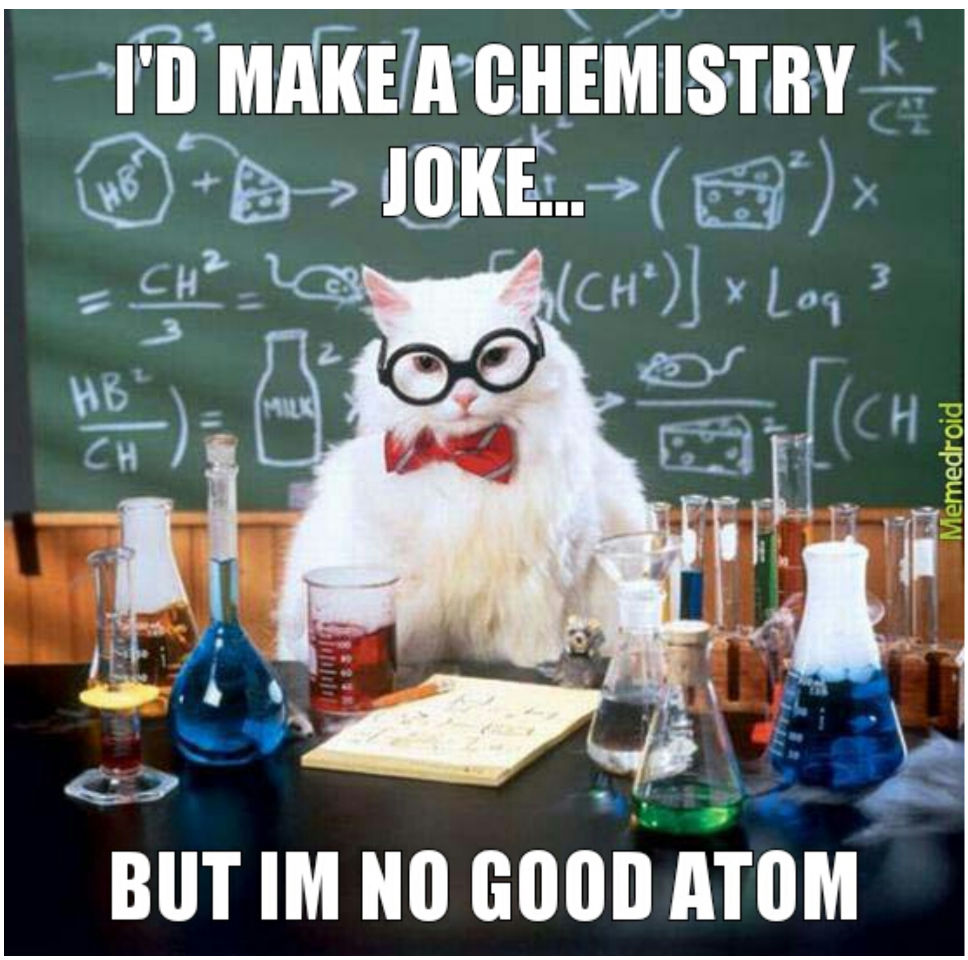
Unit 5: Atomic Structure

Practice Packet



VOCABULARY

For each word, provide a short but specific definition from YOUR OWN BRAIN! No boring textbook definitions. Write something to help you remember the word. Explain the word as if you were explaining it to an elementary school student. Give an example if you can. Don't use the words given in your definition!

Alpha particle: ____

Anion: ____

Atom: ____

Atomic Mass: ____

Atomic Number: ____

Bohr: ____

Bright Line Spectra: ____

Cation: ____

Deflected: ____

Electron Configuration: ____

Electron: ____

Excited State: ____

Gold foil experiment: ____

Ground State: ____

Isotope: ____

Mass Number: ____

Neutron: ____

Nuclear charge: ____

Nucleus: ____

Orbital: ____

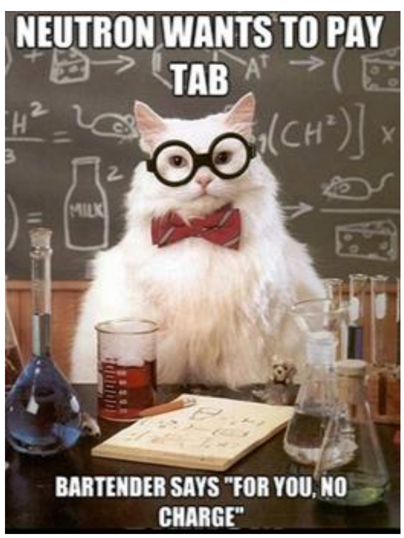
Proton: ____

Rutherford: ____

Thompson: ____

Valence Electrons: ____

Wave Mechanical Model: ____



Introduction: Models of the Atom

Why is the history of the atom so important?

It is fundamental to the understanding of science that science is a process of trial and improvement and represents the best known at the time, not an unerring oracle of truth. Development of an idea and refinement through testing is shown more in the understanding of atomic structure.

John Dalton (1803)

Modern atomic theory (what people think the atom looks like) was born with Dalton when he published his work in 1803. Dalton was the first scientist to propose that matter is composed of tiny particles called atoms. Actually, I'm lying! He really stole this idea from an ancient Greek theorist named Democritus. There was no evidence or proof of atoms back in 460 BC, so people just thought Democritus was crazy and he wasn't recognized for his theory. With all of the advancements made in science over time, Dalton was able to propose the following theory about atoms and people believed him. Some of the points in Dalton's theory still stand today, however some we now know to be untrue.

i. All elements are composed of tiny indivisible particles called atoms that *cannot* be broken down further. *
ii. Atoms of a given element are all similar in their physical and chemical properties. *
iii. Atoms of different elements differ in their physical and chemical properties.
iv. Atoms of different elements combine in simple, whole number ratios to form compounds.
v. In chemical reactions, atoms are combined, separated, or rearranged, but never created, destroyed, or changed. *

******indicates that the postulate still holds true today***

Key Content Questions (Dalton)

1. Why has the structure of the atom changed so drastically over time?

2. Sketch a diagram of Dalton's atom.

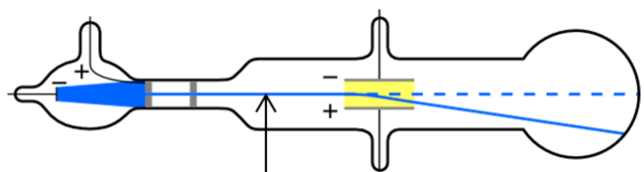
3. What property did Dalton use to differentiate between atoms of different elements?

4. Even though new models were discovered after Dalton created his 5 principles, which of these principles still holds true? (Actually write them out)

5. What is conservation of mass?

J.J. Thomson (1889)

Prior to the work of British physicist J.J. Thomson in 1889, scientists believed that atoms were simply solid spheres with no internal structure. Thomson conducted a number of experiments using cathode rays (beams of radiation produced by atoms that travel from one piece of metal to another inside of a glass tube) and proposed the idea that atoms are actually made up of smaller particles.



Thomson discovered that cathode rays travel in straight lines except when they are bent by electric or magnetic fields. Because the cathode rays bent away (were repelled) from a negatively charged plate, Thomson concluded that these rays are made of negatively charged particles. Thomson then measured the mass of cathode rays, showing they were made of particles that were around 1800 times lighter than the lightest atom, hydrogen. Therefore, the particles were not atoms, but a new particle, the first subatomic particle to be discovered.

Soon after this discovery, Thomson began speculating on the nature of the atom. He knew that atoms had an overall neutral charge, so if these subatomic particles within atoms are negatively charged, the atom must contain a positive counterpart. Thomson suggested the "Plum Pudding Model" (yum!!) to describe an atom. A modern illustration of this idea would be a chocolate chip cookie, with the chips representing negatively charged particles and the dough representing positively charged sphere that they are in.

Key Content Questions (Thomson)

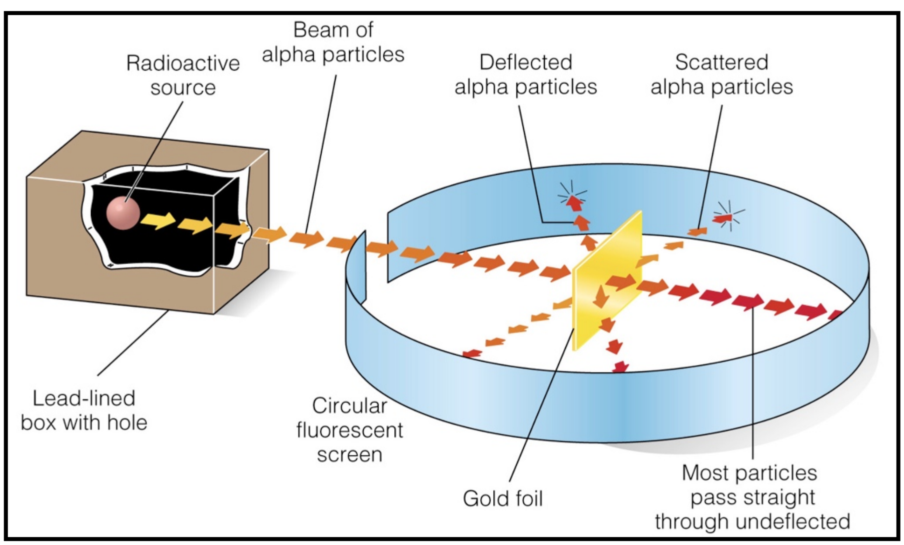
1.What are cathode rays and how are they produced?

2. What did Thomson determine the charge of an electron to be? (positive or negative)

3. How did he determine this charge based on his experimentation?

4. Sketch a diagram of Thomson's atom and label the positive and negative parts.

Rutherford (1911)



Ernest Rutherford conducted the famous Gold Foil Experiment in 1911 which disproved the plum pudding model. Alpha particles, which are positively charged, were fired at the gold atoms located in a very thin layer of gold foil. Rutherford made observations regarding if and how the alpha particles passed through the sheet of gold foil. Look at the diagram to the right. Did all of the alpha particles pass through the foil? What could the inside of the atom look like if some passed through and some bounced back? What could the particles be hitting inside of the atom to make them bounce back (repel)?

The result of the experiment allowed Rutherford to conclude the following:

1) The atom consists of mostly empty space (where electrons orbit). This is why most alpha particles passed straight through.
2) Atoms have a nucleus (it is small, dense, positively charged). This is what caused the bouncing back!

Key Content Questions (Rutherford)

1. BRIEFLY, describe the gold foil experiment in your own words.

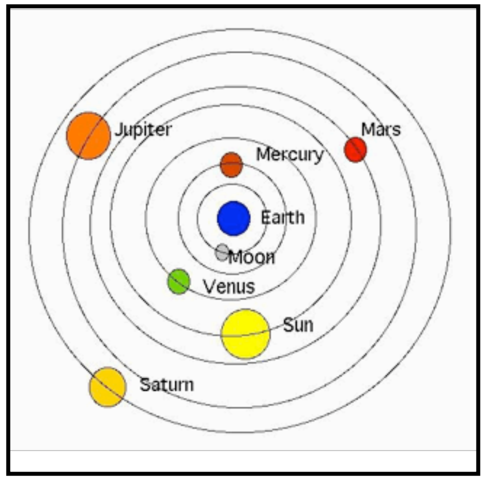
2. What were two observations Rutherford made about the alpha particles during the gold foil experiment?

3. What did these observations lead Rutherford to conclude about the structure of the atom?

4. Why do alpha particles bounce off of the nucleus?

5. Sketch a diagram of Rutherford's atom. Label all important parts.

Niels Bohr (1913)



Until 1913, scientists believed that the atom consisted of a very small, very dense nucleus (center), containing a positive charge. They also believed that atoms were made of mostly empty space aside from small, negatively charged particles (called electrons), far outside of the nucleus.

Niels Bohr proposed some changes regarding the structure of an atom based off of observations and calculations made from studying the spectral lines produced by a hydrogen atom. Spectral lines are different wavelengths (colors) of light produced by an atom when it releases energy. By looking at the number of spectral lines produced by different elements and the different wavelengths of light produced (which are based on how much energy the light particles have), Bohr realized that electrons must orbit the nucleus in "shells", each with a certain energy level. He made the following assumptions based on his experiments:

1. The electron in a hydrogen atom travels around the nucleus in a circular orbit.
2. The further the electron is from the nucleus, the more energy it has and vica-versa.
3. There is a certain amount of space in each orbit for a certain number of electrons. The first electron shell can only hold 2 electrons. The second electron can hold 8 electrons, etc.
4. If an atom absorbs energy, electrons that are close to the nucleus can jump to outer electron shells.
5. When electrons jump back to their original electron shells, light (spectral lines) is emitted (released).

The Bohr model depicts the atom as a nucleus with electrons orbiting around it at specific distances. His model is also known as the “planetary model” due to the fact that electrons orbiting the nucleus in different shells resemble the planets of our solar system orbiting the sun at very specific distances.

Key Concept Questions (Bohr)

1. What are spectral lines?

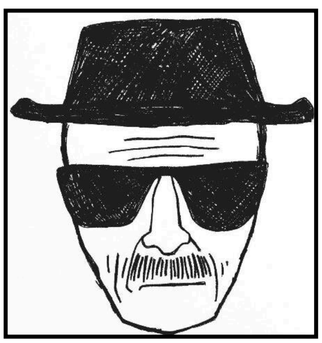
2. How are spectral lines (light) produced by atoms?

3. Why is Bohr’s model of the atom also called the “planetary model”?

4. Sketch a diagram of Bohr's atom. Label all important parts.

Wave Mechanical Model (Electron Cloud Model)

The current atomic model is a result of work done from the 1920's to the present by multiple scientists. We now know of a third subatomic particle discovered by a scientist by the name of Chadwick. These particles are also found in the nucleus and contribute to the mass of an atom, but have no charge.

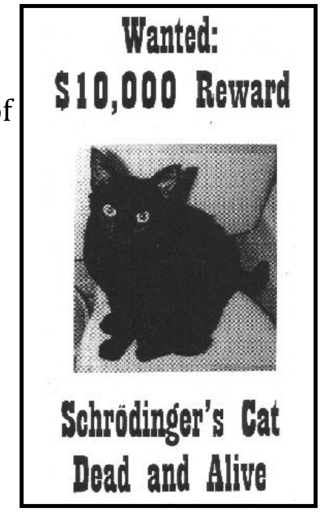


Other ideas were suggested based on the research of multiple scientists regarding the behavior of the electron. Modern physics tells us that the movement of atoms and molecules, including subatomic particles like electrons, cannot be described by the classical laws of physics. Quantum physics was discovered to describe how these tiny things behave. The Uncertainty Principle, discovered by Heisenberg, states that you cannot measure how fast a particle is moving and exactly where it is located simultaneously, with great accuracy. This means that the path of any particle, like an electron, can never be predicted exactly! All quantum physics can do is provide you with probabilities for the location of an electron revolving around an atom (meaning we can determine the general area that the electron is most likely in, but not its exact location). The electron cloud model emerged out of a mathematical equation written by German scientist Schrodinger. The Schrodinger model predicts clouds of probability around the atom where electrons can be found, which we now call orbitals.

Key Concept Questions

1. Which subatomic particle was discovered by Chadwick? What role does this particle play within the atom?

2. What is the "Uncertainty Principle"? How did this principle change our ideas on the behavior of the electron?

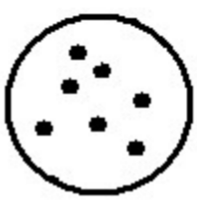


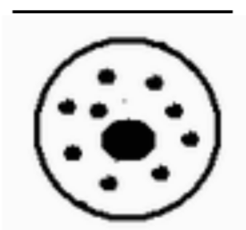
3. Which other scientists, aside from Chadwick, contributed to our current theory of the atom?

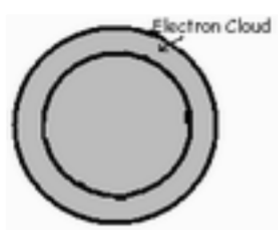
4. Sketch a diagram of the "Electron Cloud Model" of the atom. Label all important parts.

Extension Question:

Label the models (Dalton, Thomson, Rutherford, Bohr, or Quantum Mechanical model).







LESSON 5.1: ATOMIC THEORY

Objective:

- Describe how the modern model of the atom has evolved over a long period of time through the work of many scientists
- Relate experimental evidence to models of the atom
- Describe in detail Rutherford's Experiment and the conclusions he made

Observation of Thompson's plum pudding model:

Go to https://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_en.html and click on plum pudding model. Click on the blue button on the gun on the left side labeled "alpha particles." Observe what is occurring and answer the questions below.

1. According to what you observed about Thompson's plum pudding model, what was happening with the alpha particles?

2. *Is this model accurate from what you now know about the structure of the atom?*

Observation of Rutherford's Model:

Go to https://phet.colorado.edu/sims/html/rutherford-scattering/latest/rutherford-scattering_en.html and click on Rutherford Atom. Then click on the blue button on the gun on the left side labeled "alpha particles." Observe what is occurring and answer the questions below.

1. As the positively charged alpha particles are shot at the gold foil what happens to most of the particles based upon the simulation?

2. Do all particles pass right through the atoms undeflected?

3. What can you conclude about the charge of the nucleus from the simulation? What about the size?

Development of the Model of the Atom

Using your notes, draw a model for each of the following:

Dalton | Thompson | Rutherford | Bohr

REGENTS PRACTICE

1.) J.J. Thomson's Cathode Ray Tube experiment led to the discovery of

1. the positively charged subatomic particle called the electron
2. the positively charged subatomic particle called the proton
3. the positively charged subatomic particle called the electron
4. the negatively charged subatomic particle called the electron

2.) According to the Bohr Model,

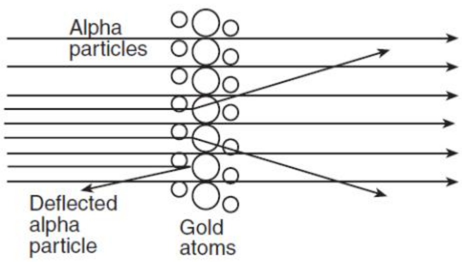
1. electrons are found in areas of high probability called orbitals
2. electrons travel around the nucleus in circular paths called orbits
3. electrons are found in areas of high probability called orbits
4. electrons travel around the nucleus in random paths called orbitals

3.) According to the Wave-Mechanical Model,

1. electrons are found in areas of high probability called orbitals
2. electrons travel around the nucleus in circular paths called orbits
3. electrons are found in areas of high probability called orbits
4. electrons travel around the nucleus in random paths called orbitals

4.) In Thomson's cathode-ray experiment, what evidence led him to believe that the ray consisted of particles, and why did he conclude that the ray was negatively charged?

5.) *One model of the atom states that atoms are tiny particles composed of a uniform mixture of positive and negative charges. Scientists conducted an experiment where alpha particles were aimed at a thin layer of gold atoms. Most of the alpha particles passed directly through the gold atoms. A few alpha particles were deflected from their straight-line paths. An illustration of the experiment is shown below.*



a. Most of the alpha particles passed directly through the gold atoms undisturbed. What does this evidence suggest about the structure of the gold atoms?

b. A few of the alpha particles were deflected. What does this evidence suggest about the structure of the gold atoms?

c. How should the original model be revised based on the results of this experiment?

6.) *In 1897, J. J. Thomson demonstrated in an experiment that cathode rays were deflected by an electric field. This suggested that cathode rays were composed of negatively charged particles found in all atoms. Thomson concluded that the atom was a positively charged sphere of almost uniform density in which negatively charged particles were embedded. The total negative charge in the atom was balanced by the positive charge, making the atom electrically neutral. In the early 1900s, Ernest Rutherford bombarded a very thin sheet of gold foil with alpha particles. After interpreting the results of the gold foil experiment, Rutherford proposed a more sophisticated model of the atom.*

a. State *one* conclusion from Rutherford's experiment that contradicts one conclusion made by Thomson.

b. State *one* aspect of the modern model of the atom that agrees with a conclusion made by Thomson.

7.) In terms of electrons, what can we conclude about the current model of the atom?

ASSESS YOURSELF ON THIS LESSON:

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INQUIRY ACTIVITY: BUILDING AN ATOM

Go to https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html and click "atom."

Build an Atom

*Before running the simulation, check off "stable" and click the + sign next to mass number and net charge. Then build 3 different examples of atoms that have a **stable** nucleus and fill in the table below.*

	Number of particles in your nucleus:	Draw your nucleus	What element is it?
1.	Protons: 1 Neutrons: 1		
2.	Protons: 2 Neutrons: 2		
3.	Protons: 3 Neutrons: 4		

1. Which **particle (or particles)** determines the name of the **element** you built?

Test your idea by identifying the element for the 3 cases.

	Particles	Element Symbol?	Mass Number	What Particle determines the Identity of the Element?	Name the element
1.	Protons: 6 Neutrons: 6 Electrons: 6			Proton Neutron Electron	
2.	Protons: 7 Neutrons: 6 Electrons: 6			Proton Neutron Electron	
3.	Protons: 6 Neutrons: 7 Electrons: 7			Proton Neutron Electron	

1. What subatomic part affects the **mass** of your atom?

2. What is a rule for determining the mass?

3. Using all of your rules, figure out what changes for each of these actions to an atom. You can test your ideas with the simulation. If you have new ideas, rewrite your rules.

Start the simulation with 3 protons, 4 neutrons and 3 electrons

Action	What Changes?	How Does it affect the atom?
Add a Proton	Element	
	Charge	
	Mass	

Action	What Changes?	How does it affect the atom?
Remove a Neutron	Element	
	Charge	
	Mass	

Action	What Changes?	How does it affect the atom?
Remove an Electron	Element	
	Charge	
	Mass	

Action	What Changes?	How does it affect the atom?
Add a Electron	Element	
	Charge	
	Mass	

LESSON 5.2: SUBATOMIC PARTICLES

Objective:

- Identify the subatomic particles of an atom (proton, neutron, and electron)
- Determine the number of protons, neutrons

<u>DO NOW:</u> Complete the chart below

Subatomic Particles:

Subatomic Particle	Charge	Mass
Proton		
Electron		
Neutron		

Fill in the table below using the periodic table and table S in your reference table.

Name	Symbol	Protons	Neutrons	Electrons	Atomic #	Mass # This is p + n
Hydrogen			0	1		1
	He	2				4
Carbon	C		6			12
Nitrogen			7	7	7	
Oxygen	O	8	8		8	
Aluminum			14	13		27
Iron		26	30			
	Co	27				59
Lithium			4			
Beryllium	Be			4	4	9
Boron		5				11
	Ne		10			
Sodium		11	12			
Fluorine		9	10			

The term nuclear charge represents the number and sign of the charge inside the nucleus. Protons which are ________________ charged and neutrons which are ________________ charged, are in the nucleus.

This means the nucleus is always ____________ charged. The # of nucleons is the number of particles in the nucleus _____________________ plus ___________________________.

Complete the table below (all are neutral):

Atom	Protons	Electrons	Neutrons	Mass #	Atomic #	Nuclear charge	# of Nucleons	Nuclear diagram	Element's Symbol
A	44			101		+44		57 n 44 p	
B		84	125						
C				55					Mn
D					89		227		
E			31			+28			
F		92					238		
G			81						Ba

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REGENTS PRACTICE

1. The number of protons in the nucleus of $^{32}_{15}P$ is:

   (1) 15 (2) 17 (3) 32 (4) 47

2. What is the total number of electrons in an atom with an atomic number of 13 and a mass number of 27?

   (1) 13 (2) 14 (3) 27 (4) 40

3. The mass number of an atom is equal to the total number of its

   (1) electrons only (2) protons only (3) electrons and protons (4) protons and neutrons

4. What is the total number of neutrons in an atom of $^{39}_{19}K$?

   (1) 1 (2) 20 (3) 39 (4) 58

5. The atomic number of an atom is always equal to the total number of

   (1) neutrons in the nucleus (2) protons in the nucleus
   (3) neutrons plus protons in the atom (4) protons plus electrons in the atom

ASSESS YOURSELF ON THIS LESSON:

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LESSON 5.3: ISOTOPES AND AVERAGE ATOMIC MASS

Objective:

- Differentiate between atomic number, mass number, and (average) atomic mass
- Calculate the (average) atomic mass for all isotopes of an element
- Calculate the number of neutrons in an isotope

In terms of subatomic particles, **isotopes** are atoms of the same element, which have a different number of ____________________ but the same number of _________________. Thus the ____________________ number is the same but the _______________ number is different. Isotope names can be followed by the mass number. **Example:** Carbon -14 and Carbon-12. Isotope symbols include the element's symbol with the mass number on the top left and the atomic number on the bottom left. **Example:** $^{14}_{6}C$ and $^{12}_{6}C$

Name	Isotope Symbol	Atomic Mass	Atomic Number	Neutrons	Protons	Electrons	Nuclear Charge
Lithium-6							
Lithium-7							
Boron-10							
Boron-11							
Sodium-22							
Sodium-24							
Aluminum-26							
Aluminum-27							
Iron-55							
Iron-56							
Zinc-65							
Zinc-66							

For each pair above, look up the element's atomic mass from the reference table and decide which isotope is more abundant. Circle the more abundant isotope's name.

Calculating average atomic mass

1. What is your quarter grade if your categories have the following weights?

	Your Grades:
Test (*60% of the average of all tests scores*)	**60**
Labs (*15% of the average of all lab scores*)	**90**
Quiz (*25% of the average of all hw/cw*)	**95**

Find the average atomic mass for each of the following: (show your work)

2. Silver: 55.0% Ag-107, 45.0% Ag-109

3. Indium: 40.0% In-113, 60.0% In-115

4. Rhendium: 30.0% Re-185, 70% Re-187

5. Copper: 75.0% Cu-63, 25.0% Cu-65

6. Chlorine: 75.5% Cl-35, the rest Cl-37

7. Oxygen: 99.76% O-16, 0.046% O-17, and 0.20% O-18

REGENTS PRACTICE

1. Which two notations represent isotopes of the same element?

   A) $^{14}_{7}N$ and $^{18}_{7}N$ B) $^{20}_{7}N$ and $^{20}_{10}Ne$

   C) $^{14}_{7}N$ and $^{17}_{10}Ne$ D) $^{19}_{7}N$ and $^{16}_{10}Ne$

2. Each diagram below represents the nucleus of a different atom.

   Which diagrams represent nuclei of the same element?

   A) *D* and *E*, only B) *D*, *E*, and *Q*

   C) *Q* and *R*, only D) *Q*, *R*, and *E*

3. Which two notations represent different isotopes of the same element?

   A) $^{6}_{4}Be$ and $^{9}_{4}Be$ B) $^{7}_{3}Li$ and $^{7}_{3}Li$ C) $^{14}_{7}N$ and $^{14}_{6}C$ D) $^{32}_{15}P$ and $^{32}_{16}S$

4. Base your answers to the following questions on the data table below, which shows three isotopes of neon.

Isotope	Atomic Mass (atomic mass units)	Percent Natural Abundance
^{20}Ne	19.99	90.9%
^{21}Ne	20.99	0.3%
^{22}Ne	21.99	8.8%

   a. In terms of *atomic particles,* state one difference between these three isotopes of neon.

   b. Based on the atomic masses and the natural abundances shown in the data table, show a correct numerical setup for calculating the average atomic mass of neon.

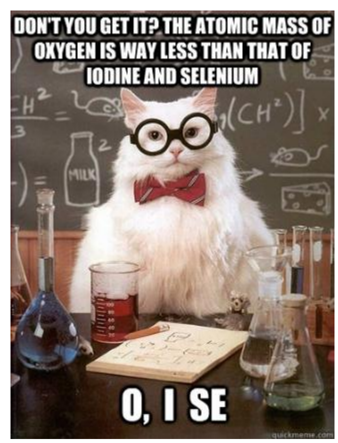
   c. Based on natural abundances, the average atomic mass of neon is closest to which whole number?

5. The atomic mass of element A is 63.6 atomic mass units. The only naturally occurring isotopes of element A are A-63 and A-65. The percent abundances in a naturally occurring sample of element A are closest to
(1) 31% A-63 and 69% A-65 (2) 50% A-63 and 50% A-65
(3) 69% A-63 and 31% A-65 (4) 100% A-63 and 0% A-65

6. What is the most abundant or common isotope of Vanadium (V)?

a) ^{50}V b) ^{52}V c) ^{51}V d) ^{49}V

7. A 100.00-gram sample of naturally occurring boron contains 19.78 grams of boron-10 (atomic mass = 10.01 atomic mass units) and 80.22 grams of boron-11 (atomic mass = 11.01 atomic mass units). Which numerical setup can be used to determine the atomic mass of naturally occurring boron?



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LESSON 5.4: IONS

Objective:

- Determine the number of protons, neutrons, and electrons in an ion

Cations are _____________ charged because they _______________ _______________ when bonding.
(positively/negatively) ***(gain/lose)*** ***(electrons/protons)***

Anions are _____________ charged because they ________________ ________________ when bonding.
(positively/negatively) ***(gain/lose)*** ***(electrons/protons)***

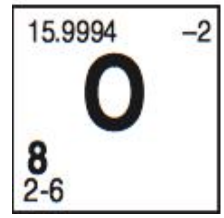
- If a particle has 9 protons and 10 electrons, what is the symbol and charge? ____________
- If a particle has 11 protons and 10 electrons, what is the symbol and charge? ____________
- If a particle has 12 protons and 10 electrons, what is the symbol and charge? ____________

Give a mathematical rule for determining the charge of a particle: ______________________________

1. *Complete the following table*

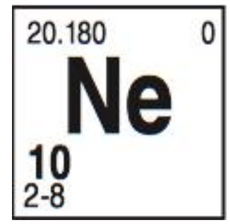
Name	Symbol	Protons	Neutrons	Electrons	Atomic #	Mass #
Copper ion	Cu^{+2}	29	35			64
Barium ion	Ba^{+2}				56	137
	K^{+}					39
Gold ion	Au^{+3}		118		79	
Chloride ion			18	18	17	
Fluoride ion	F^{-}		10			
Sulfide ion	S^{-2}		16			
	O^{-2}					16
Aluminum ion		13		10		27

Since all atoms strive to become stable (8 electrons in their valence shell) when they gain or lose electrons to form ions they become stable. The resulting electron configuration of the ion is the SAME as that of a noble gas. For example, look at the electron configuration of oxygen.



The unstable atom of oxygen has an electron configuration of **2-6**.

To become stable the atom gains **2** valence electrons resulting in 8 valence electrons and an electron configuration of **2-8** which is the same configuration as the noble gas
Neon.



*****remember when an atom forms an ion it **doesn't become a noble gas** it just **has the same configuration as one**. The proton and neutrons do not change from atom to ion and therefore will be different than the noble gas.

2. *Complete the following Table*

Ion	Charge	# of protons	# of electrons	Electron configuration	Electron Configuration same as…	Ion larger or smaller than atom?
Al^{3+}	+3	13	10	2-8	Ne (2-8)	smaller
P^{3-}						
Sr^{2+}						
Na^{+}	+1					
	-2	8				
F^{-}						

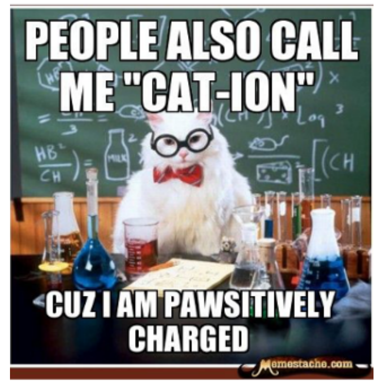
3. Explain, in terms of subatomic particles, why an oxygen atom is electrically neutral.

4. Explain, in terms of subatomic particles, why an oxide ion, O^{2-}, has a negative charge.

5. Compare the number of protons to the number of electrons in a positive ion.

6. Compare the number of protons to the number of electrons in a negative ion.

7. Explain, in terms of subatomic particles, why a chlorine ion is larger than a chlorine atom



REGENTS PRACTICE

1. What is the net charge of an ion that has 11 protons, 10 electrons, and 12 neutrons?

   A) 1+ B) 2+ C) 1– D) 2–

2. An ion that consists of 7 protons, 9 neutrons, and 10 electrons has a net charge of

   A) 2– B) 2+ C) 3+ D) 3–

3. What is the overall charge of an ion that has 12 protons, 10 electrons, and 14 neutrons?

   A) 2– B) 2+ C) 4– D) 4+

4. An ion that consists of 7 protons, 6 neutrons, and 10 electrons has a net charge of

   A) 4– B) 3– C) 3+ D) 4+

5. An oxide ion (O^{2-}) formed from an oxygen-18 atom contains exactly

   A) 8 protons, 8 neutrons, 10 electrons
   B) 8 protons, 10 neutrons, 8 electrons
   C) 8 protons, 10 neutrons, 10 electrons
   D) 10 protons, 8 neutrons, 8 electrons

6. A fluoride ion (F^-) has the same electron configuration as

   A) Na B) Na^+ C) Cl D) Cl^-

7. The correct electron configuration of the O^{2-} ion is

   A) 2-4 B) 2-5 C) 2-7 D) 2-8

8. How many electrons are in an Fe^{2+} ion

   A) 24 B) 26 C) 28 D) 56

ASSESS YOURSELF ON THIS LESSON:

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LESSON 5.5: BOHR DIAGRAMS

Objective:

- ***Construct Bohr diagrams for atoms and ions***

Bohr diagrams show the number of protons and neutrons in the nucleus and the number of electrons in their energy levels. The electron configuration shows how many electrons are in each level in the ground state, or under normal conditions.

Draw the Bohr diagrams of the following atoms: *The non-shaded boxes are positive ions & the shaded boxes are negative ions.*

H-1	**K-40**	**Li-7**	**Be-9**
F-19	**Cl-35**	**P-31**	**S-28**

Draw the Bohr Diagrams for the following ions: *The non-shaded boxes are positive ions & the shaded boxes are negative ions.*

H^{+1}	K^{+1}	Li^{+1}	Be^{+2}
F^{-1}	Cl^{-1}	P^{-3}	S^{-2}

1. Draw the Bohr diagram for B and Mg^{+3}

2. When an atom forms a positive ion, does the size get larger or smaller? Explain.

3. Draw the Bohr diagram for S and S^{-2}

4. When an atom forms a negative ion, does the size get larger or smaller? Explain.

REGENTS PRACTICE

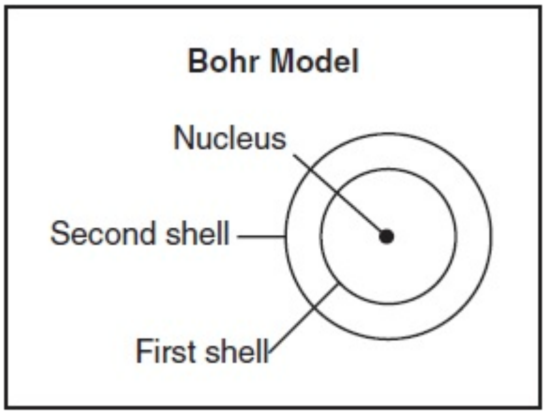
1. When a fluorine atom becomes an ion, it will
   A) gain an electron and decrease in size
   B) gain an electron and increase in size
   C) lose an electron and decrease in size
   D) lose an electron and increase in size

2. A chloride ion *differs* from a chlorine atom in that the chloride ion has
   A) more protons
   B) fewer protons
   C) a larger radius
   D) a smaller radius

3. When an atom of bromine forms a bromide ion, the radius
   A) decreases
   B) increases
   C) remains the same

4. Base your answer to the following question on the information below and on your knowledge of chemistry.

The Bohr model of the atom was developed in the early part of the twentieth century. A diagram of the Bohr model for one atom, in the ground state, of a specific element, is shown below. The nucleus of this atom contains 4 protons and 5 neutrons.



State the number of electrons in *each* shell in this atom in the ground state.

Number of electrons in first shell:

Number of electrons in second shell:

ASSESS YOURSELF ON THIS LESSON:

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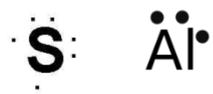
LESSON 5.6: LEWIS DOT DIAGRAMS

Objective:

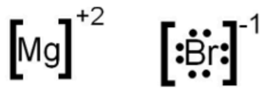
- ***Construct Lewis dot diagrams for atoms and ions***

Lewis diagrams show only the atom's symbol and dots representing the valence electrons. The most valence electrons an atom can have is ____ so the most dots you will draw is ____. Please make dots very visible!

Examples of atoms:



Name	Protons	Neutrons	Electrons	Electron Configuration	Valence Electrons	Lewis Diagram
Rubidium-85						
Barium-138						
Strontium-88						
Bromine-80						
Oxygen-16						
Nitrogen-14						

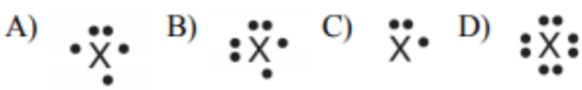
Examples of Ions: 

Name	Protons	Neutrons	Electrons	Electron Configuration	Valence Electrons	Lewis Diagram
Rb^{+1}						
Ba^{+2}						
Sr^{+2}						
Br^{-1}						
O^{-2}						
N^{-3}						

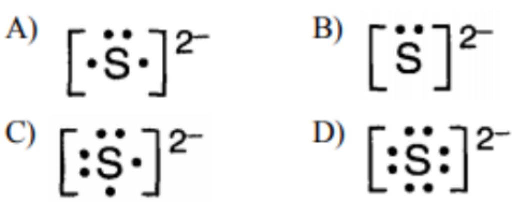
REGENTS PRACTICE

1. Which Lewis electron-dot diagram represents a nitrogen atom in the ground state?

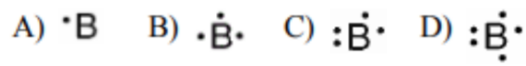
2. An atom in the ground state contains a total of 5 electrons, 5 protons, and 5 neutrons. Which Lewis electron-dot diagram represents this atom?



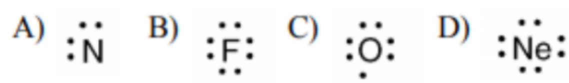
3. Which Lewis electron-dot diagram is correct for a S^{2-} ion?



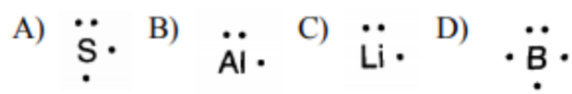
4. Which Lewis electron-dot diagram represents a boron atom in the ground state?



5. Which Lewis electron-dot structure is drawn correctly for the atom it represents?



6. Which electron-dot symbol correctly represents an atom of its given element?



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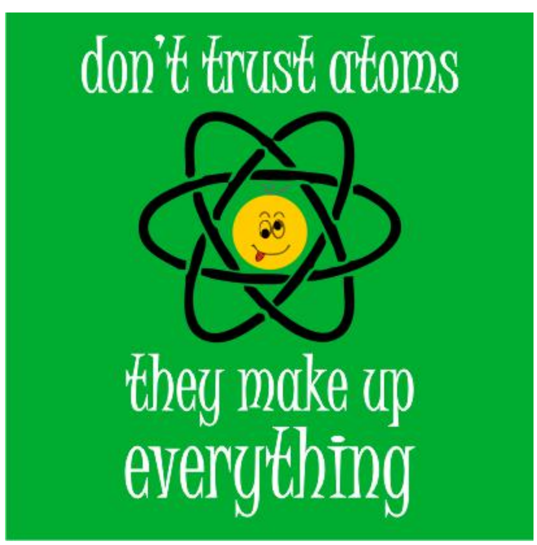


LESSON 5.7 GROUND VS. EXCITED STATE AND BRIGHT LINE SPECTRUM

Objective:

- ***Differentiate between excited and ground state***
- ***Explain how light is produced***
- ***Identify substances based upon their bright line spectra***

1. Complete the table below. The first one is done for you.

	Electron configuration	Total # of e-'s	Total # of e- shells	e- shell with the highest energy e-'s	Excited- or ground-state	Atom's symbol
Atom A	2-8-4	14	3	3rd	ground	Si
Atom B	2-3-1					
Atom C	2-8-7-1					
Atom D	2-8-18-6					
Atom E	2-8-18-17-5					
Atom F	1-7					

2. Complete the table below. The first one is done for you.

Element symbol	Element name	Ground-state electron configuration	# of valence electrons	Give an example of an excited-state electron configuration
O	oxygen	2-6	6	2-5-1
Mg				
He				
K				
N				
P				
F				
Sr				
Al				
Br				
Cu				

3. Draw a Bohr model for Sodium in the ground state. Indicate with an arrow what direction an electron moves when energy is absorbed.

4. On the same model above, draw an arrow to indicate what direction an electron moves when energy is released. What type of energy is released?

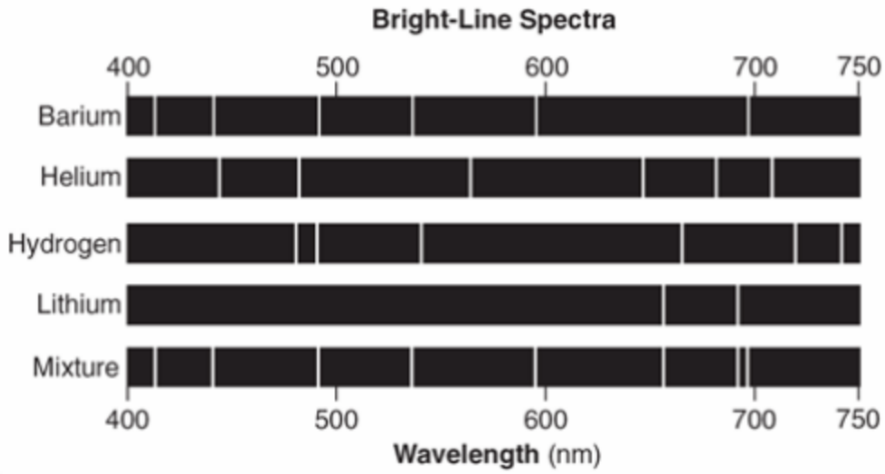
5. Base your answers to questions 4 and 5 on the diagram below, which shows bright-line spectra of selected elements.

6. Identify the *two* elements in the unknown spectrum. _______________ __________________

7. Explain how a bright-line spectrum is produced, in terms of *electrons* and *energy states.*

REGENTS PRACTICE

1. The diagram below represents the bright-line spectra of four elements and a bright-line spectrum produced by a mixture of two of these elements.



Which two elements are in this mixture?

A) barium and hydrogen
B) barium and lithium
C) helium and hydrogen
D) helium and lithium

2. Which electron transition in an excited atom results in a release of energy?

A) first shell to the third shell
B) second shell to the fourth shell
C) third shell to the fourth shell
D) fourth shell to the second shell

3. A specific amount of energy is emitted when excited electrons in an atom in a sample of an element return to the ground state. This emitted energy can be used to determine the

A) mass of the sample
B) volume of the sample
C) identity of the element
D) number of moles of the element

4. An electron in a sodium atom moves from the third shell to the fourth shell. This change is a result of the atom

A) absorbing energy
B) releasing energy
C) gaining an electron
D) losing an electron

5. During a flame test, ions of a specific metal are heated in the flame of a gas burner. A characteristic color of light is emitted by these ions in the flame when the electrons

A) gain energy as they return to lower energy levels
B) gain energy as they move to higher energy levels
C) emit energy as they return to lower energy levels
D) emit energy as they move to higher energy levels

ASSESS YOURSELF ON THIS LESSON:

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