THE ATOM

Unit 3
(Chapters 4 & 5)

VOCABULARY
ATOMIC STRUCTURE

1. atom: the smallest amount of an element that can enter into a chemical reaction. Or, the smallest amount of an element that retains the properties of the element.
2. ion: a charged particle that results when an atom gains or loses one or more electrons. Losing electrons produces an ion with a positive charge; gaining electrons results in an ion with a negative charge.
3. atomic mass: the (weighted) average mass of all the isotopes of a particular element
4. atomic mass unit: $\frac{1}{12}$ the mass of a carbon-12 atom
5. atomic number: the number of protons atoms of a particular element have
6. mass number: total number of protons and neutrons in an atom. To find it, take the element's atomic mass (from periodic table) and round it to the nearest whole number. In isotopic notation (for example ${}^{14}_{6}C$) the upper number is the mass number and the lower is the atomic number.
7. isotope: two atoms are isotopes of the same element if they contain the same number of protons but different numbers of neutrons.
8. isotopic notation: way of writing a particular isotope, showing the element, the atomic number and the mass number. Ex: ${}^{14}_{6}\mathrm{C}$ is carbon-14, with 6 protons, and a mass number of 14.
9. atomic mass unit: the mass assigned to a proton or a neutron is 1 atomic mass unit, or 1amu. (Equivalent to 1.67×10^{-24} grams)
10. electron: one of three major types of particles that make up atoms. Has a -1 charge and a very small mass compared to protons and neutrons. (It takes 1836 of them to equal the mass of one proton or neutron.) Located outside the atoms' nucleus.
11. proton: one of three major types of particles that make up atoms. Has a +1 charge, mass of 1 amu. Present in an atoms' nucleus
12. neutron: one of three major types of particles that make up atoms. Has no charge, mass of 1 amu. Present in an atoms' nucleus.
13. subatomic particles: particles that are smaller than atoms and found within atoms. Include protons, neutrons, electrons, and many other smaller particles.
14. nucleus: extremely dense center of an atom; contains the atom's protons and neutrons. Positively charged due to presence of protons in it.
15. nucleons: particles found in the nucleus of an atom; protons and neutrons are nucleons.
16. nuclear charge: the charge on the nucleus of an atom due to the presence of positively charged protons within the nucleus. The nuclear charge is equal to the number of protons in the nucleus. So, the nuclear charge is equal to the atomic number.
17. orbital: a region outside the nucleus of an atom where an electron with a particular amount of energy is likely to be located.
18. valence electrons: electrons in the outermost occupied energy level (farthest from the nucleus)
19. energy level: regions of space surrounding the nucleus; each is composed of 1 or more orbitals. The farther from the nucleus the energy level is, the more energy the electrons in it must have.
20. ground state: the state of an atom when all electrons are occupying orbitals with the lowest possible energy.
21. excited state: the state of an atom when at least one electron has absorbed energy and moved to a higher energy orbital. An unstable condition.
22. electron configuration: a way of showing the distribution of electrons in the atom's energy levels and orbitals. Ex: $1s^2 2s^2 2p^3$ (modern) or 2-5 (Bohr)
23. line spectrum: also known as an emission spectrum, a line spectrum shows a few individual colored lines, separated by blank spaces.

24. continuous spectrum: a spectrum that includes all wavelengths within a particular range. If the spectrum is in the visible range, all the colors will be visible, blending into each other so there are no "blanks" in the spectrum. (remember ROY G BIV; red orange yellow green blue indigo violet)
25. wave-mechanical model: the modern model of the atom, in which the nucleus is a dense positively charged center of the atom, surrounded by electrons found in orbitals of particular energy. Electrons are viewed as having properties of both particles and waves.

Atomic Structure:

Enduring Understanding: Metals and non-metals exhibit their properties because of electron arrangements

Essential Questions: Why do things behave the way they do? How are chemistry and fluorescent lights related? How has the view of the atom evolved over time? What are the different parts of an atom? How is an atom arranged?

Goals: I Can:

1. State that the modern model of the atom has evolved over a long period of time trough the work of many scientists including Thomson, Bohr and Rutherford.
2. Use models to describe the structure of an atom.
3. Relate experimental evidence to models of the atom.
4. List the particles present in atoms, their charges, locations in the atom, and their relative masses. State the overall charge of the nucleus and that of an atom.
5. State that the number of protons equals the number of electrons in a neutral atom of any element.
6. State that the mass of each proton and each neutron is approximately equal to one atomic mass unit. An electron is much less massive than a proton or a neutron.
7. Distinguish an ion from an atom. Explain how an atom becomes an ion.
8. Determine the number of protons or electrons in an atom or ion when given one of these values.
9. Calculate the mass number of an atom, the number of neutrons or the number of protons, given the other two values.
10. State that atoms of an element that contain the same number of protons, but a different number of neutrons are called isotopes of that element.
11. Given an atomic mass, determine the most abundant isotope of the element from among those given.
12. The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.
13. Calculate the average atomic mass of an element, given the masses and ratios of its naturally occurring isotopes.
14. In the wave mechanical model (electron cloud model) the electrons are in orbitals which are defined as the regions of the most probable electron location (ground state).
15. State that each electron in an atom has its own distinct amount of energy.
16. Define "valence electron." State that (in general) the number of valence electrons affects the chemical properties of an element.
17. Draw a Lewis-dot (electron dot) structure of an atom and ion.
18. Distinguish between valence and non-valence electrons, given an electron configuration (eg: 2-8-2 vs 2-7-3).
19. When an electron in an atom gains a specific amount of energy, the electron is at a higher (excited) energy state.
20. When an electron returns from a higher energy state to a lower energy state, a specific amount of energy is emitted. This energy can be used to identify and element.
21. Identify an element by comparing its bright line spectrum to given spectra.
22. Distinguish between ground state and excited state electron configurations, eg: 2-8-2 vs 2-7-3.

Calendar for unit 3 Regents Chemistry: Red (2 & 3A) and Yellow (9 &8A) classes

23	***24***	***25***	***26***	***27***
B	**C**	**D**	**E**	**S**
Review/Make up Day	*Test Unit 2* *HW: History of the atom introduction*	Topic 3.1 History of the Atom **HW:** Review scientists	Topic 3.1a History Continued	Topic 3.3 Isotopes & average atomic mass HW: HW: practice (sent via remind)
	Chemistry work period		Topic 3.2 POGIL and notes on atomic structure HW: practice (sent via remind)	
30	***31***	***1***	***2***	***3***
A	**B**	**C**	**D**	**E**
Lab 3.1 (isotope)	Lab 3.2 (modeling)	Topic 3.6 Electronic structure evolution HW: practice (sent via remind)	Continued practice with modeling and Emission Tube practice HW: review for quiz assignment	**Quiz- 3.1**
Topics 3.4 & 3.5 Electronic structure and Lewis dot diagrams HW: practice (sent via remind)		Chemistry work period		Lab 3.3 Flame tests
6	***7***	***8***	***9***	***10***
S	**A**	**B**	**C**	
Checkpoint Test #1 Matter, Energy, & Atomic Structure *Checkpoint Test previous unit test grade can replace grade if it is higher	Topic 3.7 The modern model HW/ CW: applying the modern model to electrons	Unit #3 Atomic Structure Review HW- Study for Test	**Unit 3 Test**	**NO SCHOOL VETERANS DAY- end of marking period.**
	Chemistry work period		Chemistry work period	

Calendar for unit 3 Regents Chemistry: Blue (5 & 4B) and Green (7 & 8B) Classes:

23	***24***	***25***	***26***	***27***
B	**C**	**D**	**E**	**S**
Review/Make up Day	*Test Unit 2* *HW: History of the atom introduction*	Topic 3.1 History of the Atom **HW:** Review scientists	Topic 3.1a History Continued HW: practice (sent via remind)	Topic 3.2 POGIL and notes on atomic structure
Chemistry work period		Chemistry work period		Topic 3.3 Isotopes & average atomic mass HW: practice (sent via remind)
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	Lab 3.2 (modeling)		Chemistry work period	
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Lab 3.3 Flame tests		Chemistry work period		

History of the Atom: **How has our view of the atom changed over time?**

Define atom:____

____

Early Models of the Atom

Democritus- Greek Philosopher –about 450 BC

- Democritus believed atoms were ____________________, ________________________,

  ____

2 Principles of Chemical Behavior- Late 1700's

- Antoine Lavoisier – Law of Conservation of Matter (or mass)

  ____

- Joseph Lois Proust (1799)- Law of Constant Composition(also known as law of definite proportions)

  ____
  ____

John Dalton- English school teacher (1808): Dalton's atomic theory ("Father" of atomic theory")

1. ____

   ____
2. ____

   ____
3. ____

   ____

   ____

Some important definitions before we go futher…

Define proton____

____

Define electron____

____

Define neutron____

____

Thought question: Since all atoms have negatively charged electrons, shouldn't every sample of matter have a negative charge? Explain.

Atomic Models: Now that the existence of sub-atomic particles has been proven, how are the particles arranged?

J. J. Thomson: Plum Pudding Model (1897) :____

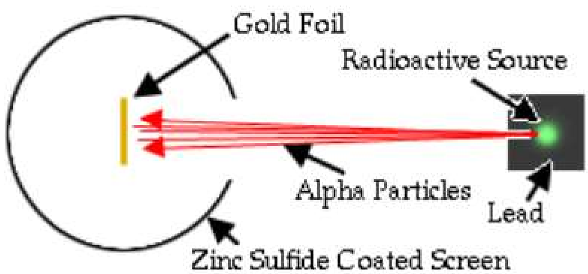
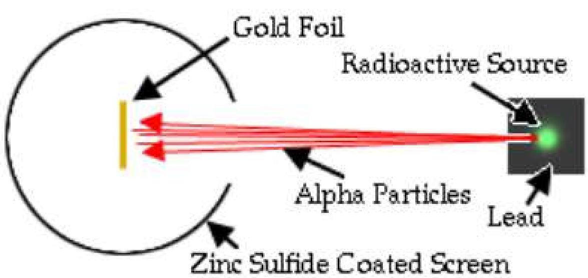
____

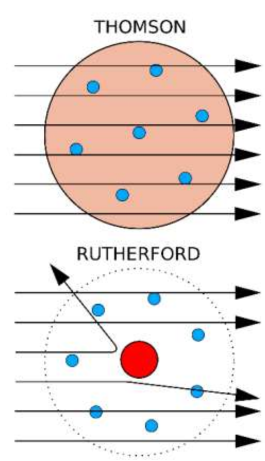
____

Earnest Rutherford: Nuclear Model- New Zeland Scientist (about 1913)

- Alpha Scattering Experiment (_____________________________ experiment)- High speed alpha particles (aka ____) were shot at gold foil- most of the particles passed through the foil, some bounced off the foil (were deflected).
- The gold foil experiment led to Rutherford's Model of the Atom

Two major conclusions from the gold foil experiment

Experimental Result (Observation)	Explanation of Result (Conclusion)
	
	



Compare Rutherford's expected outcome of the gold foil experiment with the actual outcome using these pictures.

How did Rutherford's model of the atom differ from Thomson's? (you may use the pictures to help)

What is the charge, positive or negative, of the nucleus of every atom?

Fundamental particles of an Atom: What are electrons, protons, and neutrons?

POGIL

These particles are *tiny* Most of the atom is ________________ ________________.

All chemical reactions occur because of what ________________ do. Other subatomic particles are *not* involved in chemical reactions. Often during chemical reactions, an atom may gain or lose 1 or more electrons; it is not longer an atom, it is an ion.

Atomic mass Unit- (u)

Unit of mass used to describe ________________

Based on the mass of a carbon- 12 atom (________________)

Atomic number- (See POGIL packet)__________________________

Mass Number- total mass of an atom

This number represents the total number of _____________ and _____________ in the atom.

Examples: Determining the number of p, n, e^- in an atom and a quick look at the periodic table.

Atomic number- found on the periodic table (_____________________________)

Atomic Mass Number- found on the periodic table (_________________________)

Protons- equal to the ________________ ________________

Nuclear charge= ______________________________

Neutrons- equal to the mass- number of protons

Neutrons= mass- #protons

Electrons- equal to the protons – charge

In a neutral atom, #p= $\#e^-$

Atoms=

Ions=

If you need extra practice, use this at home:

Element	Atomic number	Mass number	Charge	Protons	Neutrons	Electrons
Sr						
Al^{+3}						
S^{2-}						

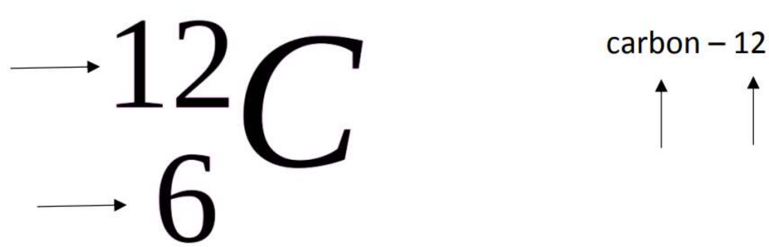
ISOTOPES & AVERAGE ATOMIC MASS

How can we discuss the mass of a large sample of atoms of an element?

Isotopes- atoms of the same element that have a different ___________ ___________.

If you see a symbol such as $^{49}_{22}Ti$, it represents an _______________ of a particular element, titanium in this case. This lower number is the ___________ ____________ (the # of _______________) and the upper number is the ____________ ____________. If you subtract the two numbers you get the number of _______________ in the atom.

How do we know how many neutrons are in an atom? Chemists use different symbols to represent the mass number and atomic number of a given atom. Here's a couple different ways to represent an atom of carbon:



Ex. Chlorine has two naturally occurring isotopes:

^{35}Cl (Cl-35) ^{37}Cl (Cl-37)

What is the atomic mass of the element chlorine? GUESS

Average Atomic Mass- the average atomic mass is a ____________ ____________

$$AAM = \frac{(mass1)(\%\ abundance\ 1) + (mass2)(\%\ abundance\ 2) + \ldots}{100}$$

Ex. There are two isotopes of chlorine. The first isotope is chlorine -35 and found 75.77% in nature. The other isotope, chlorine- 37 is found 24.23% in nature. What is the average atomic mass of chlorine? Show your numerical set-up and answer.

Try Now Please

A sample of element X contains three isotopes in the following ratios

^{45}X 65% ^{47}X 20% ^{48}X 15%

Calculate the average atomic mass of element X

Topic 3.4

Electronic Structure: How are electrons arranged within an atom?

How has the model developed since the time of Rutherford?

Electrons are very important—why?

Most of our knowledge about the electron comes from ____. (more on this on later)

Bohr Model of the atom: Niels Bohr- (Danish Physicist ~ 1913)

- e^- orbited around the nucleus (___________________________________)
- e^- existed in specific energy levels (each ____________________ represents a specified _________________ ___________________)
- If an e^- absorbed exactly the correct amount of energy (called _______________) the electron would rise to a higher energy level
- If an e^- released energy, it would fall to a lower energy level & release energy as a _____________ of light.

Ground State:

Excited state:

Each energy level can hold a specific # of electrons:

1^{st}- 2
2^{nd}- 8
3^{rd}- 18
4^{th}- 32

Bohr diagrams: this is NOT a picture of an atom, but a way to draw them

Li	Na
#p	#p
#n	#n
#e	#e

Bohr's Model was flawed as he only used the hydrogen atom.

Topic 3.5

LEWIS DOT DIAGRAMS

How can we use valence electrons to create a simplified model of the atom?

We will be working with Lewis (electron) dot diagrams more closely when we start to get into Unit 5 (Bonding). For now, we'll just introduce how to draw Lewis Dot Structures of atoms.

Example: Draw the Lewis Dot Diagram for carbon.

#	Instructions	Example
1.	Write the element's symbol.	
2.	Retrieve the electron configuration from the Periodic Table. The last number in the configuration is the number of **valence electrons.**	
3.	Use either an "x" or a dot to represent your electrons. Place the number of valence electrons around the element symbol. (order matters)	

OCTET RULE!

A completely filled valence electron shell has ________ electrons.
Exceptions:

Let's practice…

Topic 3.6

Electronic Structure and what we know NOW versus how it's evolved.

History of the models reviewed

Modern Model (1926) Schrodinger

- Electrons are described in terms of regions of _______________ _______________ ________________
-

Also referred to as :

1.
2.

The higher the principle energy level, the farther from the nucleus the electrons are and the _______________ ____. They show a "________________________" of the wavelengths of light (= individual colors) that an element emits.

Draw a model here showing the energy levels:

Electron Spectra can be used to identify each element because they are _____________ for each element. More information in supplementary materials/ worksheets.

Summary of Models of the Atom

Dalton:

Thomson:

Rutherford:

Bohr:

←Wave Mechanical:

Topic 3.7

What does the modern model tell us? (your notes may be here or in the separate POGIL packet)

The number of the energy level tells you how many sublevels that energy level has. If n=1, there is just one sublevel. If n=2, there are 2 sublevels, etc.

The first principle energy level (n=1) has ___ sublevel, an ___ sublevel.
The second principle energy level (n=2) has ___ sublevels, ___ and ____ .
The third principle energy level (n=3) has ___ sublevels, ____, ____, and ____ .
The fourth principle energy level (n=4) has ___ sublevels, ____, _____, ____ and ____ .

Sublevels are further divided up into orbitals

Sub- Level	# of orbitals
"s"	1
"p"	3-of equal energy
"d"	5-of equal energy
"f"	7 -of equal energy

All the orbitals in a given sublevel contain electrons with equal amounts of energy.

Each orbital can hold a maximum of _____ electrons; to occupy the same orbital, these two electrons must have opposite ____________.

Since s orbitals occur singly, any s level can only hold _____ electrons.

Since p orbitals occur in triplets, any p sublevel can hold _____ electrons.

Since d orbitals occur in groups of 5, any d sublevel can hold _____ electrons.

Since f orbitals occur in groups of 7, any f sublevel can hold _____electrons

What is an electron configuration? It's a way of showing where the electrons are in an atom. The electrons configurations $1s^2 2s^2 2p^6$ tells you several things.

$1s^2$ says that in the 1st energy level there are 2 electrons in an s sublevel

$2s^2$ says that in the 2nd energy level there are 2 electrons in an s sublevel

$2p^6$ says that in the 2nd energy level there are _____ electrons in a _____ sublevel

So $1s^2 2s^2 2p^6$ tells you that overall, there are ______ electrons in the atom. The first _____ energy levels have electrons in them. You can easily find this element on the period table. Count the total number of electrons. The number of protons will be the same as this number (as long as the atom does not have a charge), so that gives you the atomic number of the element.

Identifying the family on the periodic table that the atom belongs to:
Elements in column 1 all have _____ valence electron.
Elements in column 2 all have _____ valence electrons.
Elements in columns 3-12 (transition elements) are somewhat irregular, so we won't make a statement about their valence electrons.
Elements in column 13 all have _____ valence electrons.
Elements in column 14 all have _____ valence electrons.
Elements in column 15 all have _____ valence electrons.
Elements in column 16 all have _____ valence electrons.
Elements in column 17 all have _____ valence electrons.
Elements in column 18 all have _____ valence electrons.

Since the atom had 7 valence electrons, it must be an atom of an element from column ______.
Valence electrons are extremely important because:

"Building" an atom: Rules for determining where the electrons in an atom will be if the atom is in it's ground state.

1. Electrons will fill orbitals with the least energy first. The filling order is:

1s			
2s	2p		
3s	3p	3d	
4s	4p	4d	4f
5s	5p	5d	5f
6s	6p	6d	6f
7s	7p		

2. Each orbital can only hold two electrons, they must have opposite spins to occupy the same orbital.

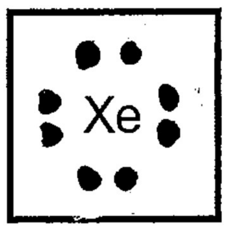
3. When filling a sublevel with multiple orbitals of equal energy, each orbital must get one electron before the others of equal energy get two. Example: the 3d sublevel

4. The outermost energy level of an atom contains only s and p electrons, and never more than 8. This means no more than _______ valence electrons are possible.

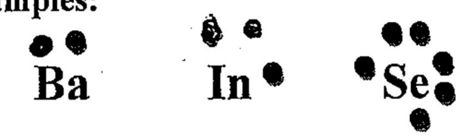
Lewis Structures

Name: ______________________

- Lewis structures, or dot diagrams, are a simplified way to show how the valence electrons are arranged in the outer shell. This is where the chemical reactions take place. Atoms will either share or give away these electrons to form bonds.
- Using your periodic table, determine the number of valence electrons for each element.
- Draw a dot to represent each valence electron around the element symbol.
- [redacted]



Examples:



H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca						

Parts of an ATOM

Name________________

An atom is made up of protons and neutrons which are in the nucleus and electrons which are in the electron cloud surrounding the atom.

The atomic number equals the number of protons. The electrons in a neutral atom equal the number of protons. The mass number equals the sum of the protons and neutrons.

The charge indicates the number of electrons that have been lost or gained. A positive charge indicates the number of electrons (which are negatively charged) lost.
A negative charge indicates the number of electrons gained.
This structure can be written as part of a chemical symbol.

Example:

Mass number → $^{12}_{6}C^{+4}$ ← *Charge*

↑ *Atomic number*

This carbon ion would have 6 protons, 6 neutrons, and 2 electrons.

Complete the following chart.

Element/ Ion	Atomic Number	Mass Number	Charge	Protons	Neutrons	Electrons
1 $^{24}_{12}Mg$						
2 $^{39}_{19}K$						
3 $^{23}_{11}Na^{+1}$						
4 $^{19}_{9}F^{-1}$						
5 $^{27}_{13}Al^{+3}$						
6 $^{1}_{1}H$						
7 $^{24}Mg^{+2}$						
8 Ag						
9 S^{-2}						
10 $^{2}_{1}H$						
11 $^{35}Cl^{-1}$						
12 Be^{+2}						

The Structure of the Atom

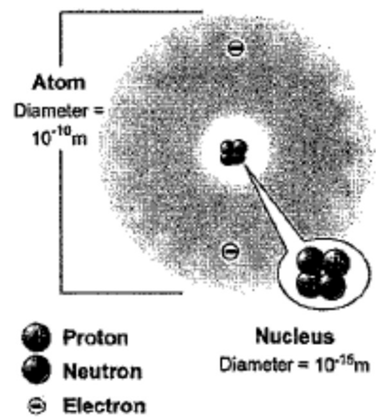
9.1

READ

Atoms are made of three tiny subatomic particles: protons, neutrons, and electrons. The protons and neutrons are grouped together in the nucleus, which is at the center of the atom. The chart below compares electrons, protons, and neutrons in terms of charge and mass.

Occurrence	Charge	Mass (g)
found outside nucleus	9.109×10^{-28}	
found in all nuclei	1.673×10^{-24}	
found in almost all nuclei (exception: most H nuclei)	1.673×10^{-24}	

Size and Structure of the Atom



The **atomic number** of an element is the number of protons in the nucleus of every atom of that element.

Isotopes are atoms of the same element that have different numbers of neutrons. The number of protons in isotopes of an element is the same.

The **mass number** of an isotope tells you the number of protons plus the number of neutrons.

Mass number = number of protons + number of neutrons

The **atomic mass** of an element is based on the mass numbers of the elements isotopes. For example, a standard table of elements lists an atomic mass of 6.94 for the element lithium. That does NOT mean there are 3 protons and 3.94 neutrons in a lithium atom! On average, 94 percent of lithium atoms are lithium-7 and 6 percent are lithium-6 (Figure 9.10). The average atomic mass of lithium is 6.94 because of the weighted average of the mixture of isotope.

EXAMPLES

Carbon has three isotopes: carbon-12, carbon-13, and carbon-14. The atomic number of carbon is 6. The atomic mass of carbon, 12.0111 amu.

a. How many protons are in the nucleus of a carbon atom?	**Answer:** 6 protons; the atomic number indicates how many protons are in the nucleus of an atom.
b. How many neutrons are in the nucleus of a carbon-13 atom?	**Answer:** 7 protons the mass number - the atomic number = the number of neutrons $13 - 6 = 7$
c. Which of the carbon isotopes is most abundant in nature?	**Answer:** At atomic mass (12.0111 amu) is based on the mass number and abundance of each of the carbon isotopes. Because the mass number is close to the whole number "12," we can assume that the most abundant isotope of carbon is carbon-12 (whose mass number is 12).

PRACTICE

Use a periodic table of the elements to answer these questions.

1. The following graphics represent the nuclei of atoms. Using a periodic table of elements, fill in the table.

	What the nucleus looks like	What is this element?	How many electrons does the neutral atom have?	What is the mass number?
a.				
b.				
c.				
d.				
e.				

2. Look at a periodic table. The atomic mass of hydrogen is 1.00794. Why is this number not rounded off to 1?

3. How many protons and neutrons are in the nucleus of each isotope?
   a. hydrogen-2 (atomic number = 1)
   b. scandium-45 (atomic number = 21)
   c. aluminum-27 (atomic number = 13)
   d. uranium-235 (atomic number = 92)
   e. carbon-12 (atomic number = 6)

4. Although electrons have mass, they are not considered in determining the mass number or atomic mass of an atom. Why?

5. A hydrogen atom has one proton, two neutrons, and no electrons. Is this atom neutrally charged? Explain your answer.

6. An atom of sodium-23 (atomic number = 11) has a positive charge of +1. Given this information, how many electrons does it have? How many protons and neutrons does this atom have?

Worksheet 7-1

ATOMS, COMPOUNDS, % COMPOSITION

1. Complete the following table:

atomic number	mass number	number of protons	number of neutrons	number of electrons	symbol of element
7			7		
		9	10		
	39			19	
	59	27			

2. Isotopes of an element contain the same number of __________ and __________, but different numbers of __________.

3. Write chemical symbols for the two isotopes of carbon. One contains $6p^+$, $6n^0$, and $6e^-$, and the other contains $6p^+$, $8n^0$, and $6e^-$.

4. Use the periodic table to determine the number of neutrons in ^{12}C, ^{15}N, and ^{226}Ra.

5. According to the periodic table, how many neutrons do each of the following have:
   a. Ag b. Cu c. He

6. What is the atomic number, mass number, and atomic mass of each of the following:
   a. Ne b. Xe c. Ga

CALCULATING AVERAGE ATOMIC MASS

Most elements have at least two isotopes, some have several. Scientists use a mass spectrometer to find out how many isotopes an element has, the % abundances of each isotope, and the mass of each isotope. From this data, the average atomic mass can be calculated. An example of how to do this follows:

DATA for the element Nickel:

Isotope	Mass of isotope	% Abundance of isotope
$^{58}_{28}Ni$	57.9353	67.88%
$^{60}_{28}Ni$	58.9332	26.23
$^{61}_{28}Ni$	60.9310	1.19
$^{62}_{28}Ni$	61.9283	3.66
$^{64}_{28}Ni$	63.9280	1.08

Calculating the average atomic mass:

Multiply each of the isotope masses by its % abundance. Add these numbers together. Divide by 100:

$$\frac{57.9353 * 67.88 + (58.9332 * 26.23) + (60.9310 * 1.19) + (61.9283 * 3.66) + (63.9280 * 1.08)}{100}$$

Now find the average mass for the three isotopes of silicon. You must show the set-up to find the answer:

$^{28}_{14}Si$	27.9769	92.21%
$^{29}_{14}Si$	28.9765	4.70
$^{30}_{14}Si$	29.9738	3.09

Isotopes and Average Atomic Mass

Elements come in a variety of isotopes, meaning they are made up of atoms with the same atomic number, but different atomic masses. These atoms differ in the number of neutrons.

The average atomic mass is the weighted average of all the isotopes of an element.

Example: A sample of cesium is 75% ^{133}Cs, 20% ^{133}Cs and 5% ^{134}Cs. What is its average atomic mass?

Answer: $\frac{(75)(133)+(20)(133)+(5)(134)}{100}$

Total= 132.85 amu (average atomic mass)

Determine the average atomic mass of the following mixtures of isotopes.

1. 80% ^{127}I, 17% ^{126}I, 3% ^{128}I
2. 50% ^{197}Au, 50% ^{198}Au
3. 15% ^{55}Fe, 85% ^{56}Fe
4. 99% ^{1}H, 0.8% ^{2}H, 0.2% ^{3}H
5. 95% ^{14}N, 3% ^{15}N, 2% ^{16}N

Draw Bohr models of the following atoms.

1. $^{1}_{1}H$

2. $^{4}_{2}He$

3. $^{7}_{3}Li$

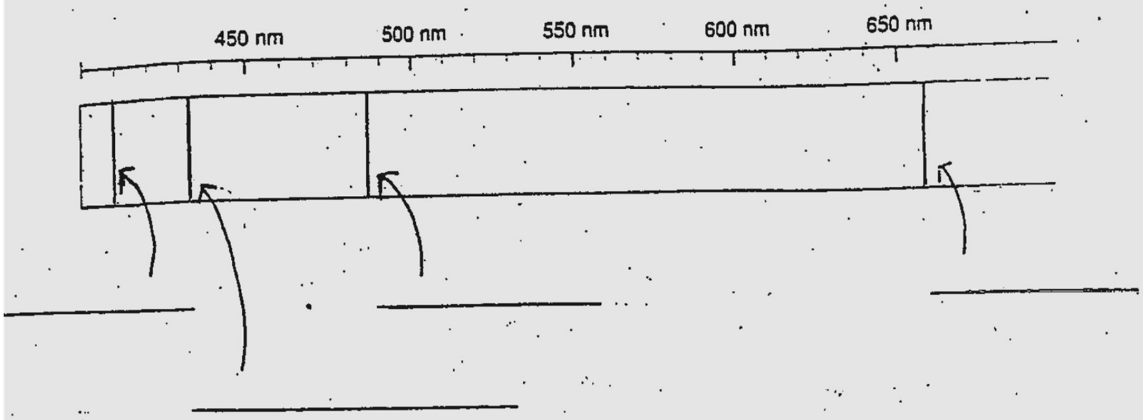
4. $^{23}_{11}Na$

5. $^{35}_{17}Cl$

6. $^{64}_{29}Cu$

2a.

b. Use the Reference Table of Colors and Their Wavelength Ranges to label each line shown below with the color that is being represented.



2b. Use the reference table of line spectra (BOTTOM OF THIS PAGE) to identify/name which gas would produce the bright line spectra below (base it on your above answer) ______________

ABLE OF COLORS AND THEIR WAVELENGTH RANGES OF LIGHT

olor	Range of Wavelengths (nm)
iolet	400 - 420
lue	420 - 490
reen	490 - 570
ellow	570 - 590
range	590 - 650
ed	650 - 700

TABLE OF REFERENCE LINE SPECTRA AT VISIBLE WAVELENGTHS (

Calcium	Helium	Hydrogen	Iron	Oxygen (band)	Mercury	Sodium	Neon
393	447	410	431	686 to 688	405	589	470
397	469	434	438		436	590	535
423	502	486	467		546		585
	588	656	496		579		640
	668		527		615		694
					691		718

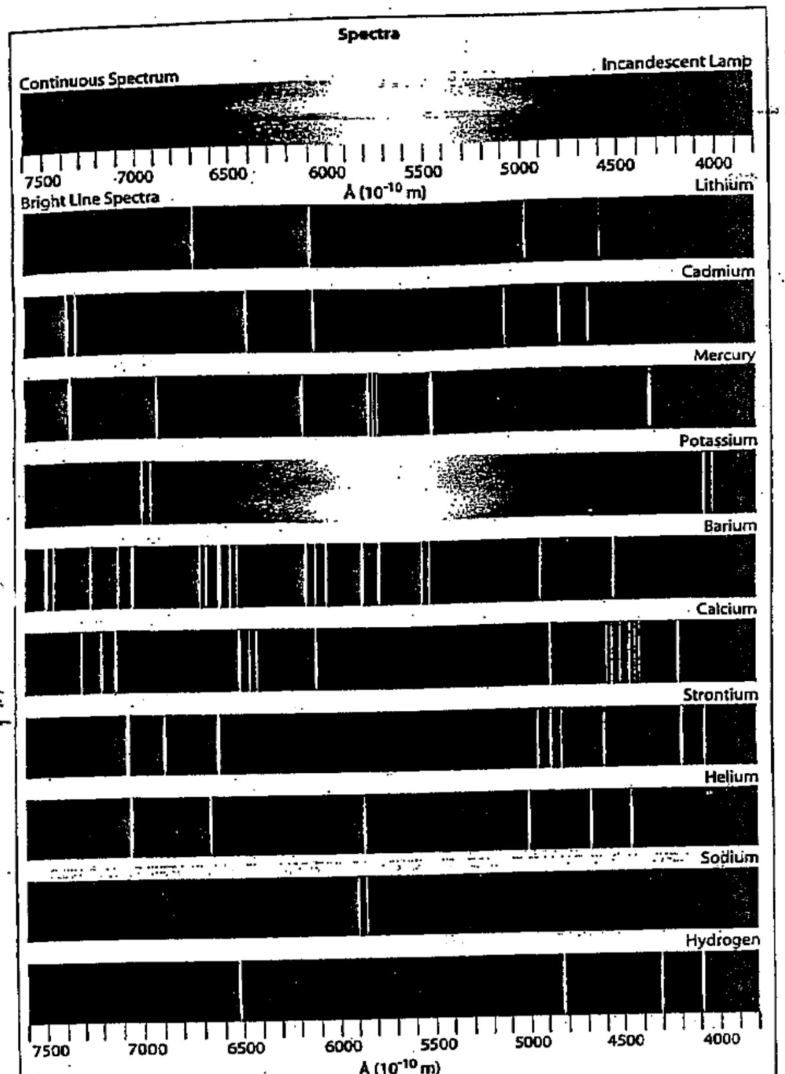


Figure 1-7. Emission spectra

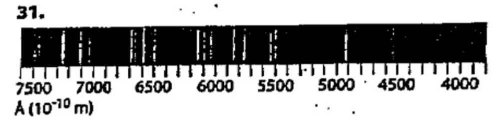
28. The characteristic bright-line spectrum of an atom is produced when (1) nuclei undergo fission (2) nuclei undergo fusion (3) electrons move from higher to lower energy levels (4) electrons move from lower to higher energy levels

29. When the electrons of an excited atom fall back to lower energy levels, the emission of energy produ (1) beta particles (2) alpha particles (3) spectral lines (4) gamma radiation

30. Which nuclide contains the greatest number of neutrons? (1) Cl-37 (2) K-39 (3) Ar-40 (4) Ca-41

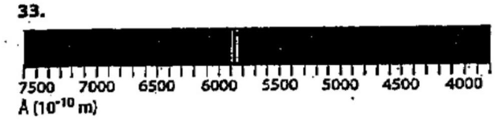
Use Figure 1-7 on page 8 to identify the following elements.

31.



32.

33.



Electron Configuration Practice Worksheet

In the space below, write the full (unabbreviated) electron configurations of the following elements: (note this is the MODERN configuration, not the Bohr configuration found on your reference tables)

1) sodium____
2) iron ____
3) bromine ____
4) argon ____
5) carbon ____
6) cobalt ____
7) nitrogen ____
8) neon ____
9) calcium ____
10) magnesium ____

Determine what elements are denoted by the following electron configurations: (note that if there is an element in [brackets] start at that element and continue with the configuration).

11) $1s^2 2s^2 2p^6 3s^2 3p^4$ ____________________
12) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$ ____________________
13) [Kr] $5s^2 4d^{10} 5p^3$ ____________________
14) [Xe] $6s^2 4f^{14} 5d^6$ ____________________
15) [Rn] $7s^2 5f^{11}$ ____________________

Determine which of the following electron configurations are not valid: State which rule has been violated.

#	Configuration	Valid (Yes/ No)	If no which rule was violate?
16.	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^5$		
17.	$1s^2 2s^2 2p^6 3s^3 3d^5$		
18.	[Ra] $7s^2 5f^8$		
19.	[Kr] $5s^2 4d^{10} 5p^5$		
20	[Xe]		