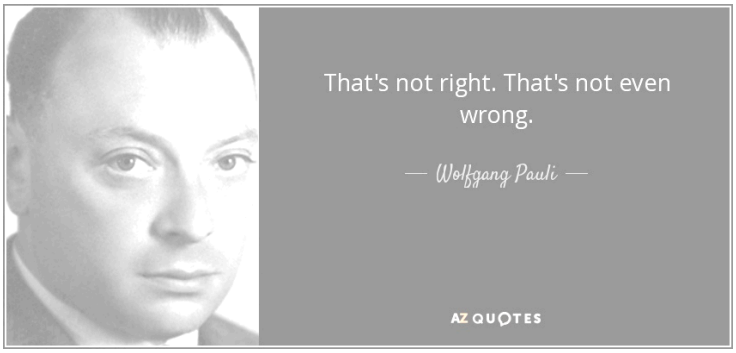
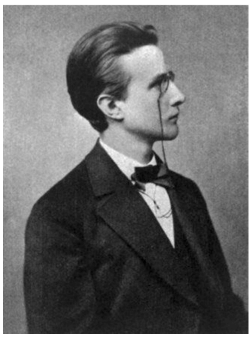
AP Learning Objectives:

- Explain the quantitative relationship between the mass spectrum of an element and the masses of the element's isotopes.(1.2)
- Represent the electron configuration of an element or ions of an element using the Aufbau principle. (1.5)
- Explain the relationship between the photoelectron spectrum of an atom or ion and: a. The electron configuration of the species. b. The interactions between the electrons and the nucleus.(1.6)
- Explain the relationship between trends in atomic properties of elements and electronic structure and periodicity (1.7).
- Explain the relationship between trends in the reactivity of elements and periodicity(1.8).
- Explain the relationship between a region of the electromagnetic spectrum and the types of molecular or electronic transitions associated with that region (3.11)
- Explain the properties of an absorbed or emitted photon in relationship to an electronic transition in an atom or molecule. (3.12)

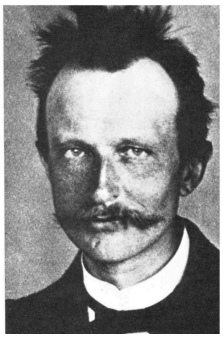


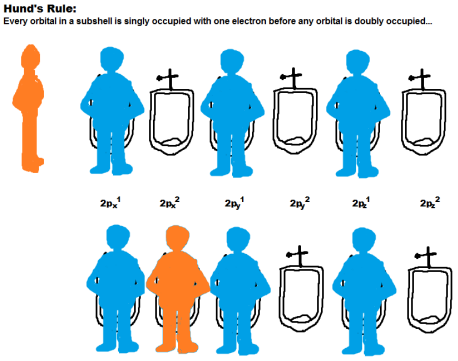
Max Planck:

Before quantum physics:

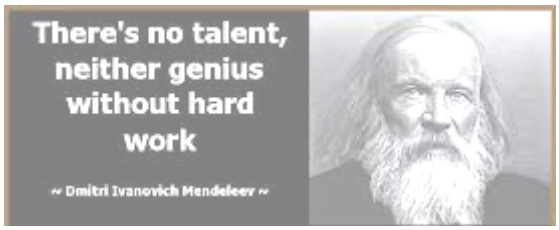


After quantum physics:

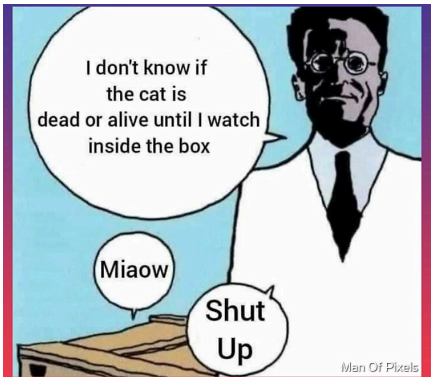












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Regents Review

Directions: Fill in the table below using your AP reference tables. There are isotopes and ions on this table.

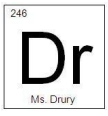
Name (and charge)	Symbol	Protons	Neutrons	Electrons	Atomic #	Mass #	Nuclear charge
Oxygen			8	8	8	16	8+
		7	7		7		
	$^{34}_{16}S$			18		34	
Hydrogen (deuterium)		1	1		1		
Hydrogen (tritium)		1			1	3	
		9				19	
Beryllium ion				2		9	
		12				24	
		12				25	
				92		238	92+
Chloride ion				18		35	
		17				37	
	$^{84}_{36}Kr$			36		84	
Iron (III) ion		26				56	
Silver ion			60	46			

1. Give the name, symbol, charge and location of all subatomic particles in an atom:

2. Define ions with an example from above:

3. Define isotopes with an example from above:

4. Naturally occurring chlorine that is put in pools is 75.53% Cl-34.969 and 24.47% Cl-36.9666. Calculate the atomic mass.



5. Magnesium has three isotopes Mg-24 (78.70%), Mg-25 (10.13%), and Mg-26 (11.7%). The actual mass of Mg-25 is 24.98584 and the actual mass of Mg-26 is 25.98259. What is the actual mass of Mg-24 if the atomic mass of magnesium equals 24.3050 amu?

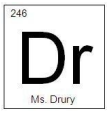
6. The results taken from a mass spectrum of chlorine gas show peaks at m/z 35.00 and m/z 37.00 (The m/z peaks on a mass spectrum identify the different isotopes of an element that are present in the sample).
   (a) Given that the relative abundances of Cl 35.00 and Cl 37.00 are 77.50% and 22.50% respectively, calculate the average relative atomic mass of chlorine atoms to four significant figures.

   (b) Suggest all the possible masses of Cl_2 molecules that are made when two chlorine atoms bond together.

   (c) Which of the molecules you have suggested in (b) will be the most abundant? Explain your answer.

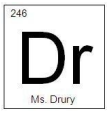
7. Draw the model, give the nickname and scientific experiment where appropriate.

Dalton	**Thompson**	**Rutherford**	**Bohr**
Model Nickname:	*Model Nickname:*	*Model Nickname:*	*Model Nickname:*
Experiment:	*Experiment:*	*Experiment:*	*Experiment:*
Discovery:	*Discovery:*	*Discovery:*	*Discovery:*



Electron Configurations

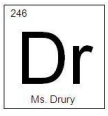
	Full e- Configuration	# Valence e-	Valence orbital notation	Lewis Diagram
H				
He				
Li				
Be				
B				
C				
N				
O				
F				
Ne				
Mg				
Si				
S				
Ar				
K				
Fe				
Zn				
As				
Br				
Sr				
Zr				
Te				
Xe				



Ions and Exceptions

	Atom's e- Configuration (Noble Gases Short Cut Allowed)	# Valence e-	Ion's e- Configuration (Noble Gases Short Cut Allowed)	Ion's Lewis Diagram
Li^{+}				
Cu^{+2}				
Co^{+3}				
Mg^{+2}				
Sc^{+2}				
Na^{+}				
Ag^{+}				
Ca^{+2}				
O^{-2}				
Zn^{+2}				
K^{+}				
Br^{-}				
As^{-3}				
Cu^{+1}				
Au^{+}				
N^{-3}				
Mo^{+}				
S^{-2}				
F^{-}				
As^{+5}				
P^{-3}				
As^{+3}				

1. Explain why arsenic can form three different ions in terms of valence electrons and stability.



2. The electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$ has _____ unpaired electrons and represents element _____.
3. The electron configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$ has _____ unpaired electrons and represents element _____.
4. Which electron configuration(s) are in the excited state? Circle all that apply:

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

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

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

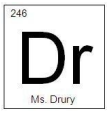
   $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5\ 4s^2\ 3d^{10}\ 4p^2$ $1s^2\ 2s^2\ 2p^6\ 3s^3$

   $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^8$

5. Were any of the above configurations not allowed? Box them and explain their errors.
6. Draw shapes for the s p and d orbitals.

7. How many electrons can fill each orbital? Why? (Give a rule and explanation)

8. How many electrons fill the s orbital? ____
9. How many electrons fill the 4p sublevel? ____
10. How many electrons fill the 3d sublevel? ____
11. How many electrons fill the 5f sublevel? ____
12. How many electrons will fill a p orbital? ____
13. For each of the following sets of orbitals, indicate which orbital is **higher** in energy.
    (a) 1s, 2s _____ (c) 4s, 3dyz _____
    (b) 2p, 3p _____ (d) 3px, 3py, 3pz _____
14. Indicate the block (s, p or d) in which each of the following elements is found.
    (a) Sc _____ (b) Ni _____ (c) P _____ (d) As _____



15. Identify the element from the electron configurations of **atoms** shown below.
    (a) [Ne] $3s^2\ 3p^2$ _____

    (b) [Ar] $4s^2\ 3d^7$ _____

    (c) [Xe] $6s^2$ _____

16. Give the symbol of the atom or ion represented by the following sets of atomic numbers and configurations.

Atomic #	Electronic Configuration	Symbol of Atom or Ion
(a) 8	$1s^2\ 2s^2\ 2p^4$	_____
(b) 11	$1s^2\ 2s^2\ 2p^6$	_____
(c) 14	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$	_____
(d) 22	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$	_____

17. Give the electron configurations for the following transition metal ions.

    (a) Sc^{3+} _____

    (b) Cr^{2+} _____

18. Consider the element Scandium, atomic # 21.
    (a) If the configuration of the element were constructed, into which orbital would the final electron be placed?

    (b) When scandium forms an ion with a charge of +1, from which orbital would the electron be removed?

19. Of the following species (Sc, Ca^{2+}, Cl, S^{2-}, Ti^{3+}), which are isoelectronic?

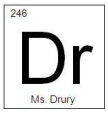
20. Identify two positive **and** two negative ions that are isoelectronic with argon.

21. Explain each scientist's rule. Consider models:
    a. Aubfau:

    b. Hund:

    c. Pauli:

    d. Heisenberg:



Magnetic Properties of Matter Lab

Guiding Question: What is the relationship between the magnetic properties of matter and e- configurations?

Materials: $CuSO_4$, $MnSO_4$, MnO_2, $KMnO_4$, NaCl, Zn, Fe, Magnets.

Prelab:

1. For each of the following, give the electron configuration ($1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$).

O _____

Na _____

Cl _____

S _____

K _____

Mn _____

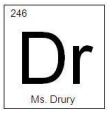
Cu _____

Zn _____

Fe _____

2. Using electron configurations in question 1, write the orbital notations for each of the elements below:

	1s	**2s**	**2p**	**2p**	**2p**	**3s**	**3p**	**3p**	**3p**	**4s**	**3d**	**3d**	**3d**	**3d**	**3d**
O	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
O^{2-}	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Na	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Na^+	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Cl	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Cl^-	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
S	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
S^{6+}	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
K	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
K^+	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Mn	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__





	1s	2s	2p 2p 2p	3s	3p 3p 3p	4s	3d 3d 3d 3d 3d
Mn^{2+}	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Mn^{4+}	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Mn^{7+}	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Cu	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Cu^{+}	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Cu^{2+}	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Zn	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Zn^{+2}	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __
Fe	__	__	__ __ __	__	__ __ __	__	__ __ __ __ __

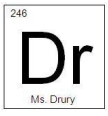
Procedure:

1. Complete the “ions present” column on your data table.
2. Hypothesize whether the reactants listed will be magnetic or not based on your prelab.
3. Determine the magnetic attraction of the substance by moving the magnet slowly across the substance.
4. Record your observations about the color and attraction of each substance.

Data:

Formula	Oxidation states	Hypothesis (Y/N)	Magnetism (Y/N)
NaCl			
$MnSO_4$			
MnO_2			
$KMnO_4$			
$CuSO_4$			
Zn			
Fe			

Claim: Answer the guiding question.





Evidence: Show relevant data and calculations that support your claim. Then explain how the data is interpreted.

Justification: Provide necessary scientific concepts that support your claim and evidence.

Additional Questions:

1. Based on your "CEJ", would CuCl be attracted to the magnet? Why or why not?

2. The element Chromium has three oxidation states: 0, 3+ and 6+.
   a. Draw the valence orbital notations for these ions.

	1s	**2s**	**2p**	**2p**	**2p**	**3s**	**3p**	**3p**	**3p**	**4s**	**3d**	**3d**	**3d**	**3d**	**3d**
Cr^{0}	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Cr^{3+}	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__
Cr^{6+}	__	__	__	__	__	__	__	__	__	__	__	__	__	__	__

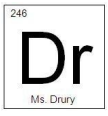
   b. Substances that are attracted to the magnet are *paramagnetic*. Which of these substances do you expect to be paramagnetic? Why?

   c. List the oxidation states in order of increasing magnetism and explain.

3. Hund's rule states that "the most stable arrangement of electrons is one with the maximum number of unpaired electrons, all in the same spin direction." Are more elements para- or di-magnetic? Explain.

4. How would your results differ if the sublevels did not fill according to Hund's rule? (If each sublevel was filled before the next, would the things that you listed as magnetic be magnetic still?) Use Fe to explain.

5. Would the magnetism be changed if you took away electrons from 3d before 4s for cations? Use Cu^{+1} to explain.





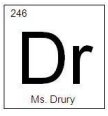
Orbital Filling Rules

The rules that you have been applying in order to determine the electronic configuration of an atom are:

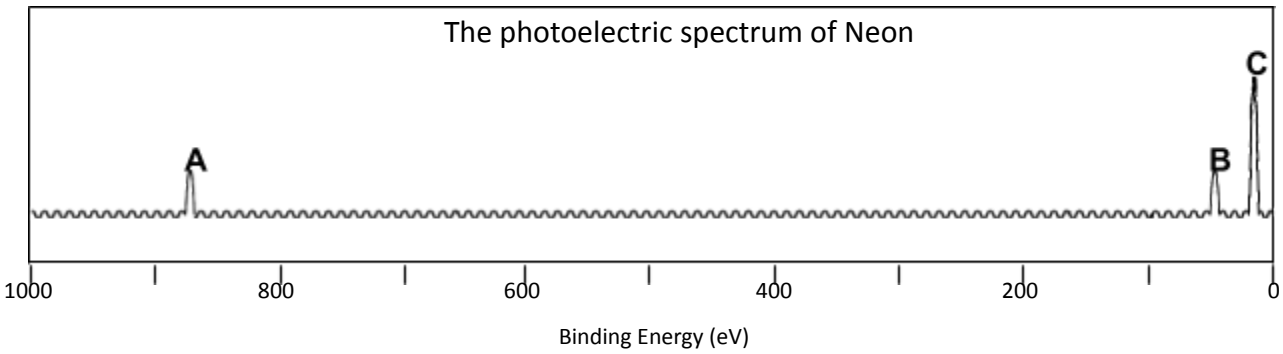
A. Lowest energy orbitals are filled first. **THE AUFBAU PRINCIPLE**.
B. Orbitals can only contain a maximum of two electrons and when two electrons enter the same orbital they must have opposite spins so that each electron has a unique set of quantum numbers. (In the electrons in boxes diagram they must be drawn  **NOT**  **OR** OR). **THE PAULI EXCLUSION PRINCIPLE**.
C. When orbitals of identical energy (degenerate) are available electrons enter these orbitals singly before any spin pairing takes place. **HUNDS RULE**.

Consider each of the elements listed and the INCORRECT electronic configuration associated with each one. In each case identify which of the above rules or principles (**A**, **B**, **C** or **D**) is violated and insert the correct electronic configuration.

Element	Incorrect Configuration	Violation	Correct Configuration
N	**$1s^2\ 2s^2\ 2px^2\ 2py^1$**	**C**	**$1s^2\ 2s^2\ 2px^1\ 2py^1\ 2pz^1$**
Al	$1s^2\ 2s^2\ 2p^6\ 3p^3$		
B	$1s^2\ 2s^3$		
P	$1s^2\ 2s^2\ 2p^6\ 3p^5$		
C	$1s^2\ 2s^1\ 2px^1\ 2py^1\ 2pz^1$		
C	$1s^2\ 2s^2\ 2px^2$		
Mn	[Ar] $4s^1\ 3d^6$		
Ni	[Ar] $4s^2\ 3dxy^8$		
Sc	[Ar] $3d^3$		
B	$1s^2\ 2s^1\ 2px^1\ 2py^1$		
Na	$1s^1\ 2s^2\ 2p^6\ 3s^2$		
S	[Ne] $3s^2\ 3px^2\ 3py^2$		
V	[Ar] $3d^5$		
P	[Ne] $3s^2\ 3px^2\ 3py^1$		
Kr	[Ar] $4s^2\ 3d^{16}$		



Photoelectron Spectra (PES)

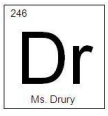


1. The peaks A, B, and C represent the binding energies of electrons in which subshells of neon?
   a. 1s, 2s and 2p
   b. 2p, 2s and 1s
   c. 1s, 1s and 1s
   d. 2s, 2p and 2p

2. Which of the following statements best accounts for the peak A being far to the left of peaks B and C:
   a. The electron configuration of neon is $1s^2\ 2s^2\ 2p^6$
   b. Neon has 8 electrons located in its valence shell
   c. Core electrons of the atom experience a much higher effective nuclear charge that the valence electrons
   d. Peaks B and C show first ionization energies of electrons in neon, whereas peak A shows the second ionization energy or neon.

3. Which of the following statements best accounts for peak C being three times the height of peak B:
   a. The intensity of the photoelectron signal at a given energy is a measure of the number of electrons in that energy level.
   b. Electrons represented by peak B have approximately triple binding energy than those represented by peak C.
   c. In the photoelectron spectrum, as binding energy increases, the relative number of electrons decreases.
   d. The height of the peaks in a photoelectron spectrum does not have any relation to the structure of the atom.

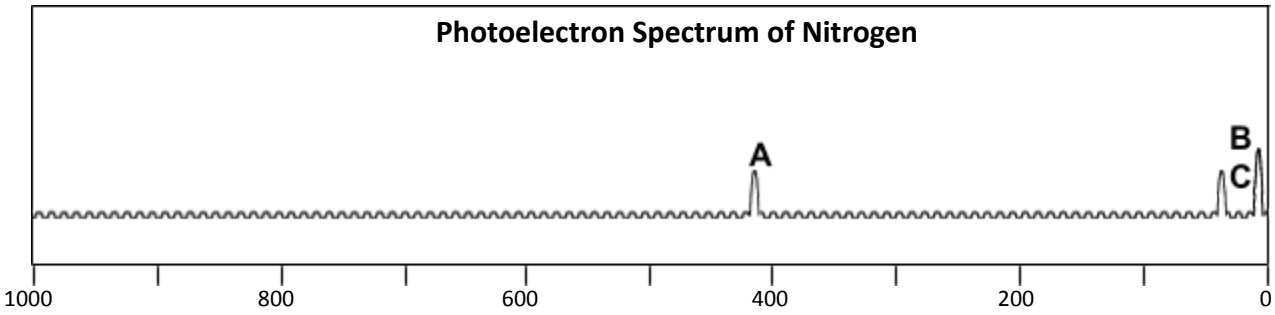
4. Nitrogen shows 3 peaks in its photoelectron spectrum. These 3 peaks correspond to the binding energies of nitrogen's 1s, 2s and 2p electrons. The relative heights of the 1s, 2s, and 2p peaks in the PES spectrum of nitrogen will be:
   a. 1, 1, 1
   b. 2, 2, 3
   c. 1, 1, 2
   d. 2, 2, 5



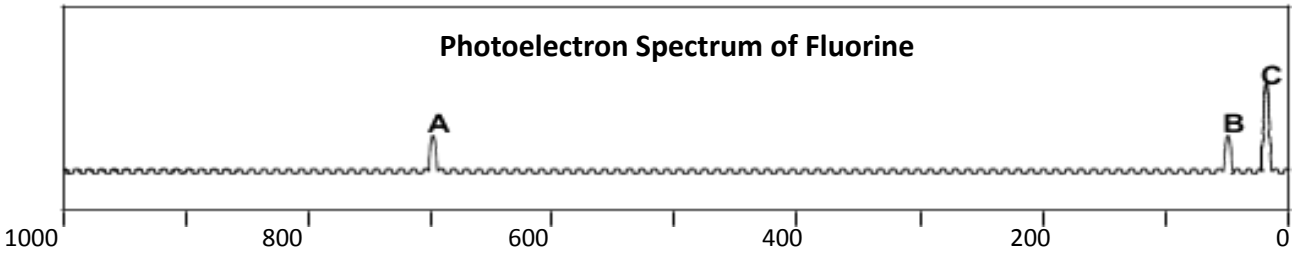
5. If the binding energy of an electron is between 0 - 10 eV, it can be assumed that the electron originated in:

a. an s orbital b. a p orbital c. a core orbital d. a valence orbital

6. Shown below are the photoelectron spectra of nitrogen (atomic number 7) and fluorine (atomic number 9).



Binding Energy (eV)

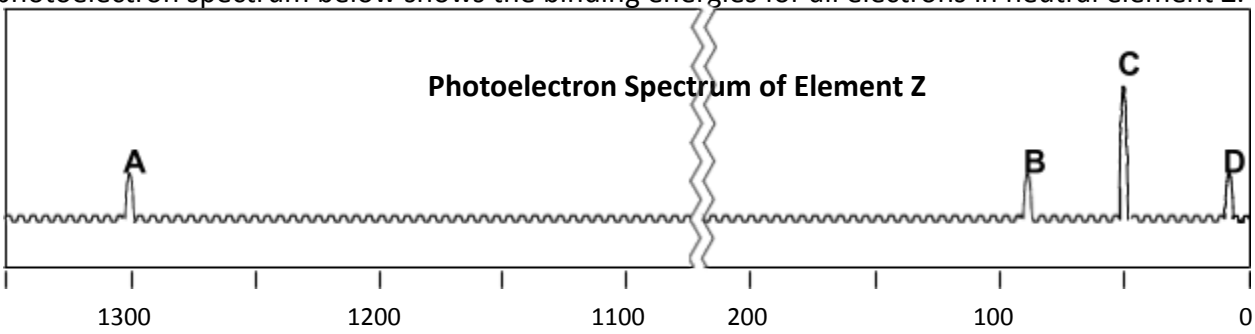


Binding Energy (eV)

Given the data in the two spectra, what do you expect the binding energy of the 1s electrons in oxygen (atomic number 8) to be?

a. 0 - 50 eV b. 50 - 400 Ev c. 400 -700 eV d. 700 - 1000 eV

7. The photoelectron spectrum below shows the binding energies for all electrons in neutral element Z:



Binding Energy (eV)

What is the identity of element Z?

a. boron b. carbon c. neon d. magnesium

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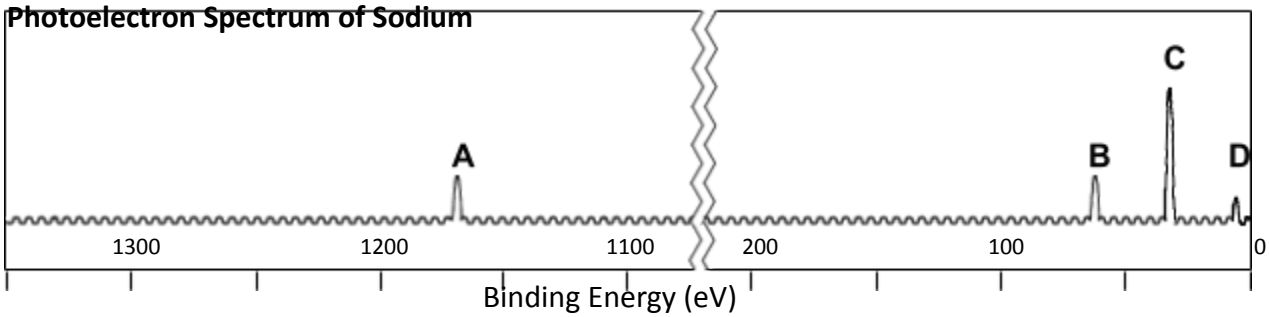
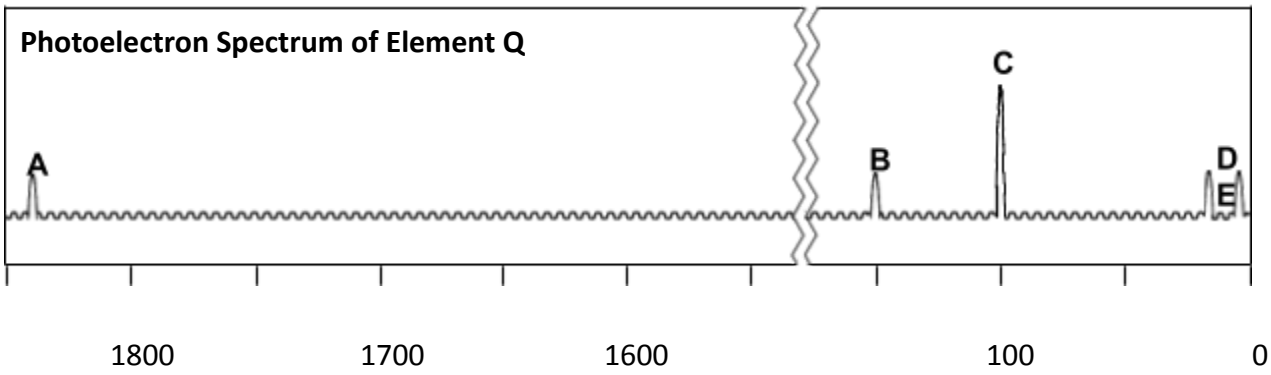


8. Identify the peak (A-D) in the photoelectron spectrum of sodium that represents the binding energy of electrons in the 2s orbital.

8. Identify the peak (A-D) in the photoelectron spectrum that gives the binding energy of sodium's valence electrons.

9. Which peak (A-D) in the photoelectron spectrum of sodium shows electrons closest to the nucleus?

10. Which peak (A-D) in the photoelectron spectrum of sodium represents an orbital containing 6 electrons?



12. The electrons that feel the strongest effective nuclear charge are given by which peak?
a. A b. B c. E
d. all electrons in this spectrum experience the same effective nuclear charge

13. Which peaks in the photoelectron spectrum represent valence electrons of element Q?
a. A only b. B and C c. D and E d. E only

14. Which peaks in the photoelectron spectrum of Q are given by the binding energy of p orbital electrons?
a. C only b. D only c. C and E d. B, C, and D

15. What is the identity of element Q?
a. carbon b. oxygen c. magnesium d. silicon

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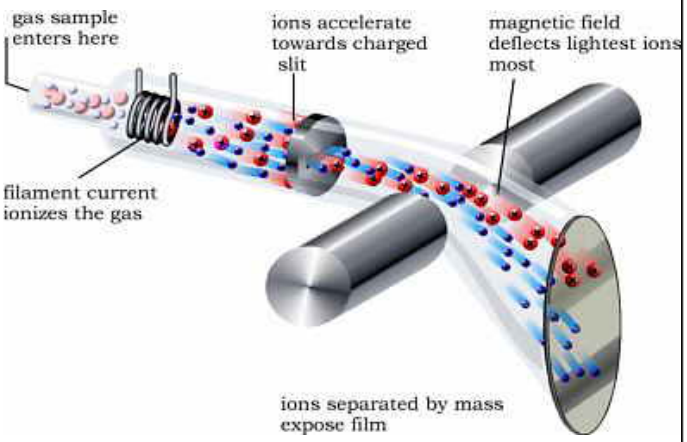
Mass Spectrometry, Percent Abundance & Atomic Mass

Mass spectrometry is a technique used in chemistry to determine average atomic mass values for the elements. It can also be used to identify unknown *compounds*, such as in forensics and archeology.

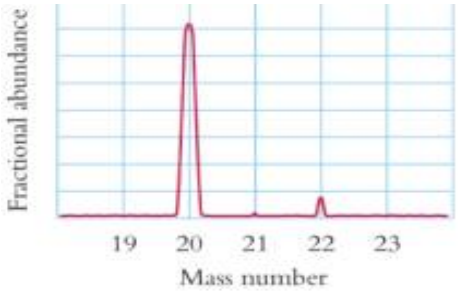
Mass Spectrometer



How a Mass Spectrometer Works:

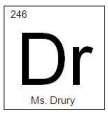
- A gaseous sample is hit with electrons, which produces positive ions
- These positive ions are sent through a magnetic field, which exerts a force on the charged ions
- The ions are deflected by the force of the magnetic field and hit a film on the other side of the magnetic field – this is how they are detected
- The lighter (less massive) particles are affected by the force of the magnetic field more than the heavier particles, and so will be deflected more from the straight-line path.
- By comparing the amount or degree of deflection for the particles in a sample of a given element to the standard deflection angles for a sample of carbon-12, atomic masses* can be determined. *These are *relative* atomic masses – relative to the standard (carbon-12).
- From the *number* of hits at each deflection point, the relative abundances of the isotopes in the sample can be determined.

Mass Spectrometer Reading for Neon

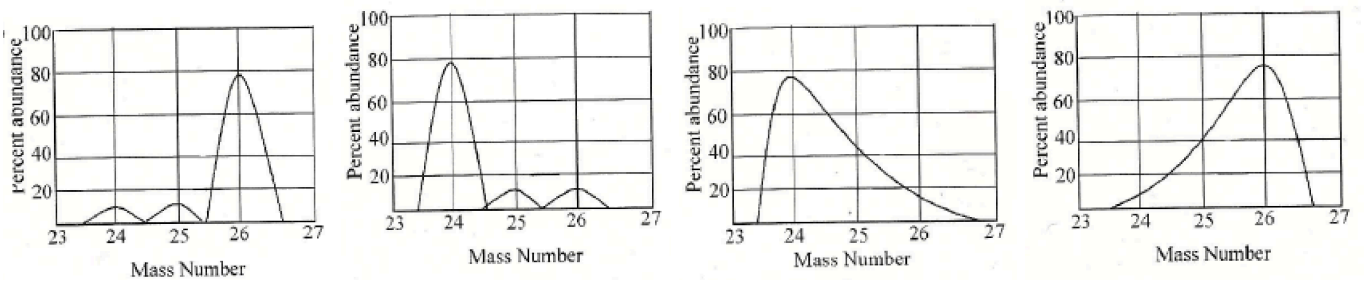


How to Read a Mass Spectrogram:

- A peak indicates that particles with that mass number were detected
- The taller the peak, the more particles there were in the sample with that particular mass
- This spectrogram for neon shows a very tall peak at mass = 20 and a short peak at mass = 22. This means that there were more neon-20 atoms than neon-22 atoms. The neon-20 isotope contributes more to the average atomic mass for neon because there are more atoms of neon-20 than neon-22. Looking up the average atomic mass for neon on the Ref. Tables, we find that it is 20.18 amu, which makes sense because there are many more neon-20 atoms than neon-22 atoms, making the average atomic mass closer to 20 than to 22.
- You may be asked to pick out a correct spectrogram for an element, given its atomic mass, on the AP Exam.



1. Which mass spectrometer reading represents naturally occurring magnesium?



(A) **(B)** **(C)** **(D)**

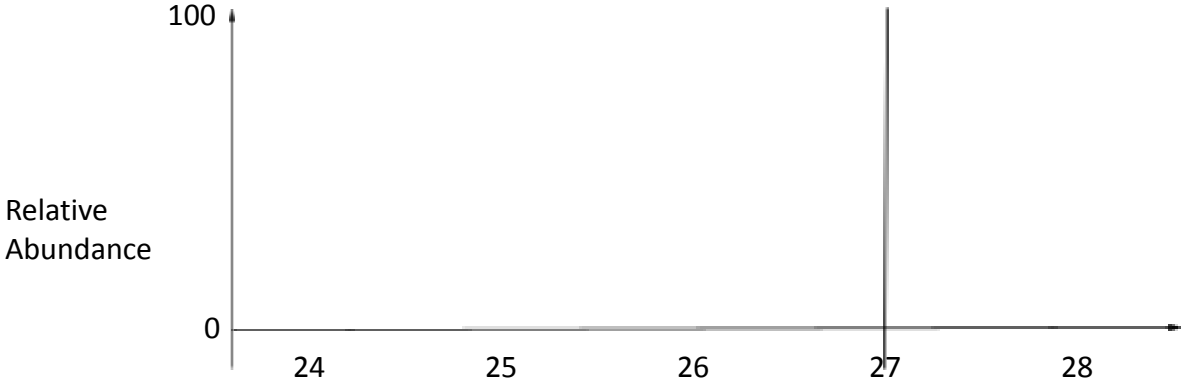
2. An unknown element 'Z' is analyzed in a mass spectrometer and is found to have the following isotopes with the corresponding relative abundances.

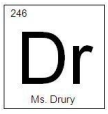
Isotope Z	Relative Abundance
50 Z	4.34
52 Z	83.79
53 Z	9.50
54 Z	2.37

a. Sketch the expected mass spectrum that these data would provide. Label the axes and pay attention to the size of any lines that you draw.

b. Calculate the average atomic mass of Z and identify the element.

3. Consider the following mass spectrum that was generated from the analysis of an element. What does the existence of only a single peak in the spectrum suggest about the element? What is the element's identity?





4. How many peaks would be observed in the mass spectrum of O_2^+, given that there are three, commonly occurring isotopes of O with mass numbers of 16, 17 and 18?

5. Bromine has two isotopes, Br-79 and Br-81. The isotopes occur in a 50:50 (1:1) ratio. Given that the mass spectrum of bromine contains peaks for both bromine atoms *and* diatomic bromine molecules, predict the number of peaks in the spectrum. What would be the relative height of the atomic peaks? What would be the relative height of the molecular peaks? Assume that z = +1 in each case.

6. A sample of carbon atoms were chemically combined with a sample of oxygen atoms to yield the compound carbon dioxide. The sample of oxygen atoms was artificially manufactured to have a 3:1 ratio of O-16 and O-18, Assuming carbon to have only a single isotope C12, predict the mass spectrum that the resulting sample of carbon dioxide would produce in terms of the number of peaks. Which peak would be the smallest? Explain. Assume the mass spectrum only includes peaks for the compound carbon dioxide and that z = +1 in each case.

7. Neon atoms are known to produce a mass spectrum that consists of three peaks at m/z values of 20, 21 and 22, with the relative abundance of the peaks found to be in the ratio 112 : 0.21 : 11.1. Assuming the z value of the species causing the peaks to be +1 in each case, calculate the average atomic mass of neon based on these data.

8.

81
Tl
204.38

There are two stable isotopes of the element thallium (Tl). Information about the mass number and the percent abundance for one of these isotopes is shown in the table below.

Fill in all of the missing information in the table below. Show your scratch work and calculations in the space below the table.

Isotope	Mass Number	Protons	Neutrons	Abundance
^{205}Tl	205			70.5%

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Electromagnetic Radiation (EMR)

Light, which is also called electromagnetic radiation, carries energy and has characteristics of both particles and waves (deBroglie). The absorption and emission of ER is one of the powerful tools used to probe molecular structure and chemical reactions, forms the basis of medicine's magnetic resonance imaging (MRI), and is intrinsic to many analytical techniques used to monitor manufacturing processes and the environment. You will use the emission and absorption of light by matter to obtain the relative energies of the atomic and molecular orbitals and other energy levels, which are associated with vibration, rotation, electron spin and nuclear spin as well as spectra.

Reference Table Formulas:

$E = h\nu$	E = energy	Planck's constant, $h = 6.626 \times 10^{-34}$ Js
$c = \lambda\nu$	ν = frequency	Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
	λ = wavelength	Avogadro's number = 6.022×10^{23} mol^{-1}

Key Questions:

1. What is the equation showing the relationship between energy of a photon and frequency of light? How are they correlated (direct or indirect?) Explain.

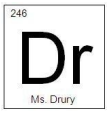
2. What is the relationship between frequency and wavelength? Explain how they are correlated.

Exercises:

3. The laser in a cd uses light with a wavelength of 780nm.

    a. Calculate the frequency of this light.

    b. Calculate the energy of a single photon of this light.



Hydrogen Spectra Lab

Guiding Question: Which element, of the elements tested, will emit a color that is correlated to the movement of an electron from the 6th level to the 2nd?

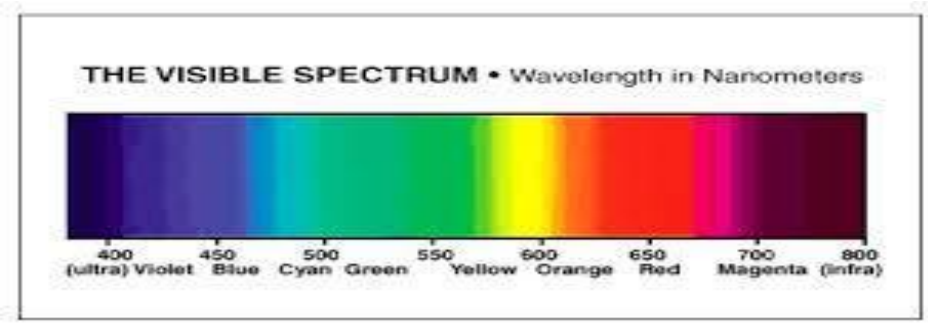
Background Information: In 1885 Balmer and Rydberg observed that when hydrogen gas is excited by an electric current, it emitted light which produced a visible spectrum when resolved through a prism. They then identified the wavelengths of the spectral lines using the following formula

$$E_n = -2.178 \times 10^{-18} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

In 1913, Bohr suggested that electrons of hydrogen moved in circular orbits around the nucleus. The electrons were held in those orbits due to two opposing forces, the attraction of the electrons and the protons as well as the momentum of two similarly charged and moving electrons. He described the electron having the lowest energy moving around the nucleus at the closest distance as the ground state (n=1). If energy was absorbed by the electron, it would jump up to a higher energy level and be in the excited state. Depending on the amount of energy absorbed, electrons could jump up into many different energy levels (n=1,2,3…). Naturally, the electrons would fall back down, releasing energy in the form of light. This could occur in one or more steps. Regardless, energy difference between levels could be calculated using equation 2 and 3:

$$\Delta E = h\upsilon \qquad c = \lambda v$$

This equation, however, is only accurate for the hydrogen atom because it does not take into account the electrostatic interaction between electrons or the shielding effect of inner electrons. But the general idea can be used for all atoms. The electromagnetic spectrum is given below. In this experiment, you will calculate the energy, wavelength, and/or frequency for each of the visible emission lines of hydrogen. The spectroscope used will not have a scale (like the one intended for this lab), so a scale has been made for you in the data section. The scale is based on helium's spectra.

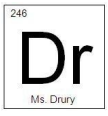


Pre-lab Questions:

1. The element sodium emits 3.37×10^{-19} J/atom of energy.
    a. Calculate the frequency of this light.

    b. Calculate the wavelength of this light.

    c. In what region of the ER spectrum would the sodium emission fall?
    d. What color is the light?



2. a. Calculate the energy associated with an electron falling from level 2 to 1.

b. Calculate the wavelength of light produced in this emission.

c. What part of the electromagnetic spectrum does this emission correlate to?

d. Why will we be assuming all of the visible spectra are a result of electrons falling to level two instead of level one?

Procedure:

1. Observe spectra of hydrogen, Helium and Neon and record your observations in the box below.
2. Using the data table, graph the spectral lines for Helium and then color those lines based on what you see in the spectra and the chart on page 1.

Data Collection:

Helium (nm)

300 400 500 600 700 800

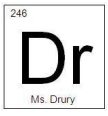
Hydrogen (nm)

300 400 500 600 700 800

Neon (x 10^{-19} J)

6 5 4 3 2

He	H	Ne
Wavelength (nm)	*Wavelength (nm)*	*Energy (J))*
438.8	410.1	3.68×10^{-19}
443.8	434.0	3.39×10^{-19}
447.1	486.1	3.34×10^{-19}
471.3	656.2	3.29×10^{-19}
492.2		3.27×10^{-19}
501.6		3.22×10^{-19}
504.8		3.19×10^{-19}
587.6		3.16×10^{-19}
667.8		3.13×10^{-19}
		3.11×10^{-19}
		3.10×10^{-19}
		3.05×10^{-19}
		3.01×10^{-19}
		2.87×10^{-19}
		2.82×10^{-19}



Calculations: (Show all work with proper sig figs and units!)

1. Using the **hydrogen data**, calculate the frequency of **all of the lines.**

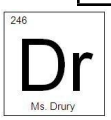
2. Using the **hydrogen data**, calculate the energy of **all of the lines.**

3. Using the **helium** data, calculate the frequency and energy of the line with the **shortest wavelength.**

v of shortest wavelength	E of shortest wavelength

4. Using the **helium** data, calculate the frequency and energy of the line with the **longest wavelength.**

v of longest wavelength	E of longest wavelength



5. Using the **neon data**, calculate the frequency and wavelength of the line with the **most energy**.

ν of most energy	λ of most energy

6. Using the **neon data**, calculate the frequency and wavelength of the line with the **least energy**.

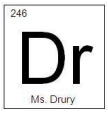
ν of least energy	λ of least energy

Claim: Answer the guiding question:Which element, of the elements tested, will emit a color that is correlated to the movement of an electron from the 6th level to the 2nd?

Evidence: Show relevant data and calculations that support your claim. Then explain how the data is interpreted.

Justification: Provide necessary scientific concepts that support your claim and evidence using terms such as electron, energy level, excited state, ground state, energy, wavelength, and frequency.





Light Equations Practice

1. **Find the frequency of the following:**
   a. A wave with a wavelength = 5.60×10^{-5}m.
   b. A wave with a wavelength of 555 nm.
   c. A wave with 2.31×10^{-20} J of energy.
   d. A wave with 3.56×10^{-30} kJ of energy.

2. **Find the wavelength of the following in nm:**
   a. A wave with a frequency of 40050 s^{-1}.
   b. A wave with a frequency of 2.00×10^{6} Hz.
   c. A wave with 3.89×10^{-25} J of energy.
   d. A wave with 4.57×10^{-31} kJ of energy.

3. **Find the energy of the following:**
   a. A wave with a frequency of 1.00×10^{3} s^{-1}.
   b. A wave with a frequency of 3.40×10^{18} Hertz.
   c. A wave with a wavelength of 2.89×10^{3} m.
   d. A wave with a wavelength of 567 nm.

4. **Convert each of the energies to kJ/mol.**

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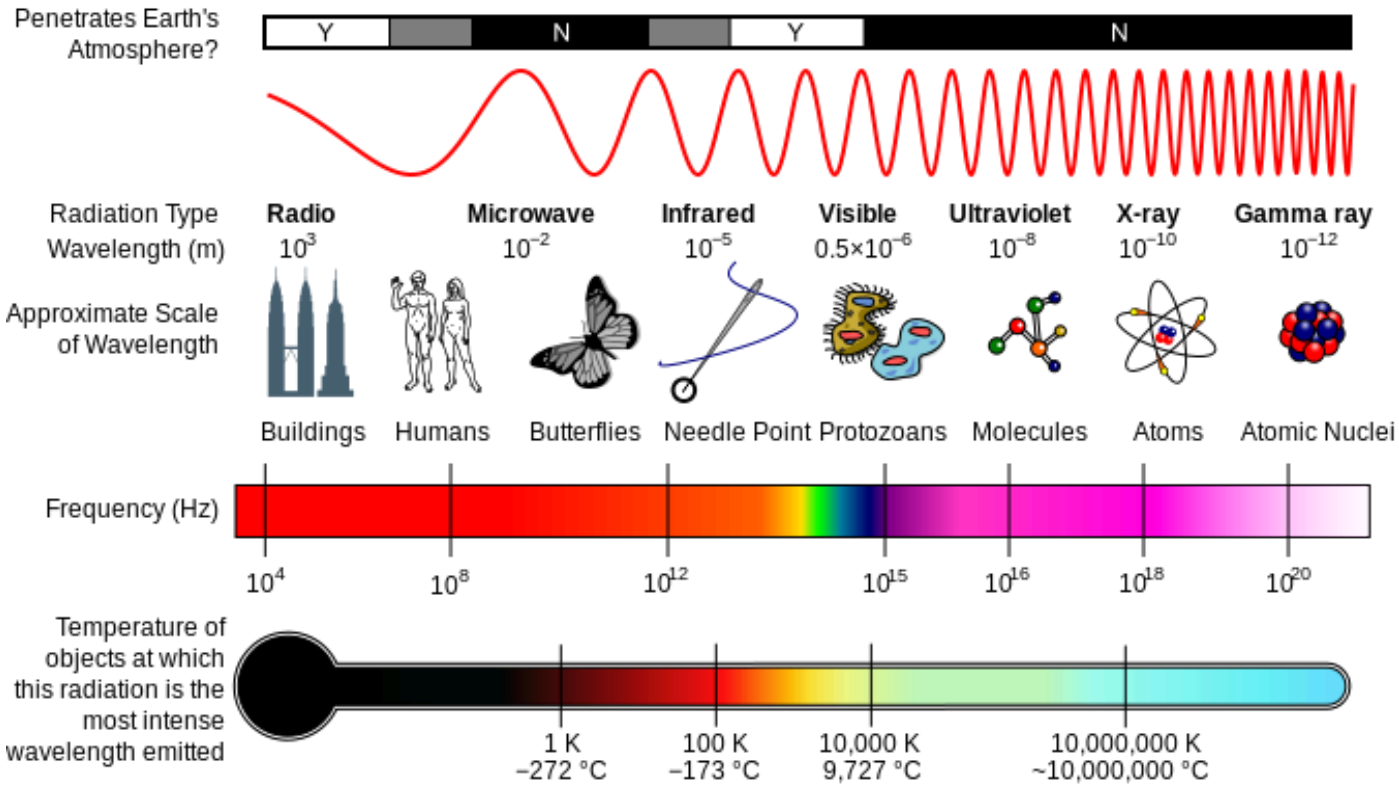
5. Find the type of radiation emitted using the following diagram for questions 3a to 3d.

6. The N-O bond requires approximately 200kJ/mol to break. Which (if any) of these EMR would break the bond?





Periodic Table Review

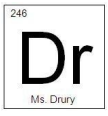
1. Define and name the periodic groups with properties.

Group	Name	Properties
1		
2		
3-12		
17		
18		

2. Define the periodic periods and explain what elements in the same period have in common with each other.

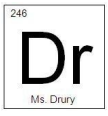
3. Put a check in each box that correctly describes the element given.

	Metal	Metalloid	Nonmetal	Alkali Metal	Alkaline Earth Metal	Transition metal	Halogen	Noble gas	Monatomic	Diatomic
Sb										
Sr										
Rn										
P										
Pt										
Cs										
S										
Fe										
Br										
Ar										



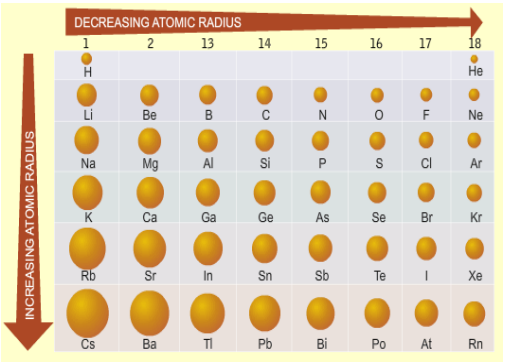
4. Use the property provided to match with the terms.

	Metal	Metalloid	Nonmetal	Property
a.				Located on the left side of the P.T.
b.				Located on the right side of the P.T.
c.				Solids are brittle
d.				Majority of the elements
e.				Gain electrons to form negative ions
f.				Located along the “staircase”
g.				Have luster
h.				Malleable
i.				Lose electrons to form positive ions
j.				Ductile
k.				Excellent conductors of heat & electricity
l.				Poor electrical & heat conductors
m.				Low electronegativity values
n.				Low ionization energy
o.				High ionization energy
p.				High electronegativity values
q.				Ions are larger than their atoms
r.				Ions are smaller than their atoms



Atomic Radius

Trends:



Across a period atomic radius _____

due to _____

_____

Down a group atomic radius _____

due to _____

---

1. Identify and explain the trend in atomic size for the following transitions in the periodic table.
   (a) Moving vertically from Ar to He

   (b) Moving horizontally from Ar to Na

2. In each of the following pairs, pick the larger species. Explain your answer in each case.
   (a) Cu and Cu^{2+}

   (b) F and F-

   (c) Na^{+} and K^{+}

3. Only one of the following statements is correct. Which one?
   (a) All cations are larger than their corresponding atoms
   (b) All anions are smaller than their corresponding atoms
   (c) Atomic size increases on transitioning from left to right across period 2 of the periodic table
   (d) The most common ion of chlorine is smaller than a chlorine atom
   (e) The most common ion of strontium is larger than a strontium atom
   (f) The most common potassium ion is larger than the most common sodium ion
   (g) The ions most commonly formed by group 16 elements are smaller than their corresponding atoms

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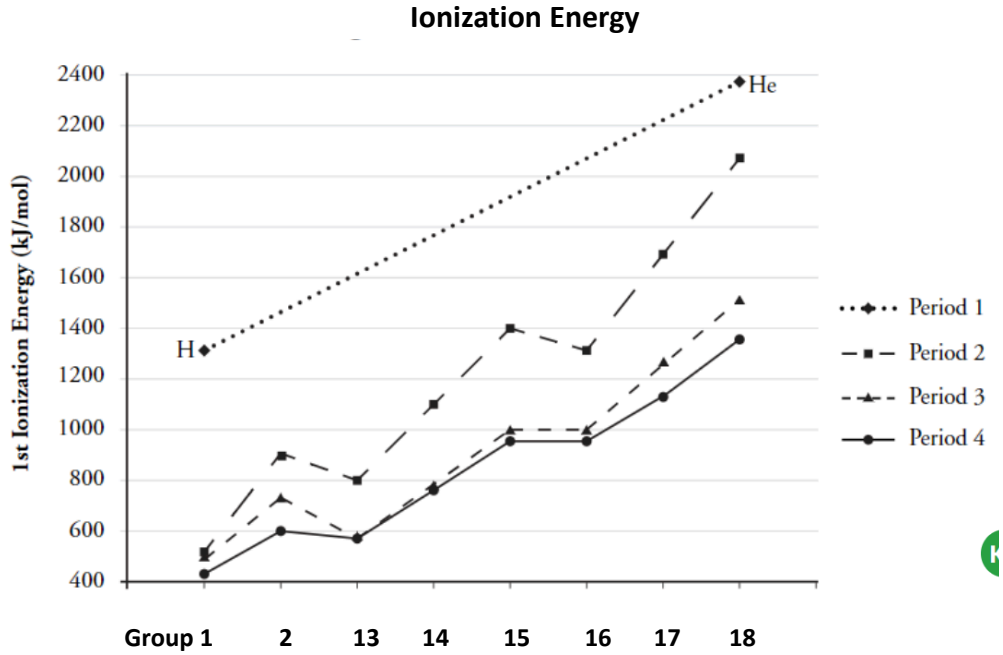


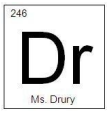
1. How does the graph in Model 3 illustrate the periodic trend for ionization energy as you move across a row of the periodic table? Explain the trend.

2. How does the graph in Model 3 illustrate the group trend for ionization energy as you move across a row of the periodic table? Explain the trend.

3. Are either of the trends you described in the questions above perfect?
   If not, where are the "blips" or discrepancies in the trends? Which elements do they represent? Label them on the graph.
4. Draw electron configuration and notation of the "blip" element to describe the discrepancy in Group 13.

5. Draw electron configuration and notation of the "blip" element to describe the discrepancy in Group 16.

6. Create a hierarchical list of most stable to least stable electron configuration/notations for s and p orbitals.



Atomic Radii and Ionization Energy Trends Practice

1. Using the metal magnesium as an example, write two separate equations to show the first and second ionization energy of magnesium. (Remember state symbols are important as they form part of the definition).

   First Ionization: _____

   Second Ionization: _____

2. Which of the following elements (one from each pair) would you expect to have the highest first ionization energy? Explain your answers.

   Ca or Be: _____

   Na or Ar: _____

3. Consider the table:

IE	1st	2nd	3rd	4th
	578	1817	2745	11580

   (a) In which group does this element appear on the periodic table?

   (b) Predict the formula of the compound that this element forms with fluorine.

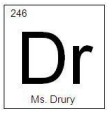
   (c) What is the minimum number of electrons that this element must have?

4. Arrange the following species in order of increasing size. Rb^+, Y^{3+}, Br^-, Kr, Sr^{2+} and Se^{2-}.

5. Are there any atoms for which the second ionization energy is greater than the first? Explain your answer.

6. Is it possible for two different atoms to be isoelectronic? If so, give examples.

7. Is it possible for two different anions to be isoelectronic? If so, give examples.



8. Consider the table below:

IE	1st	2nd	3rd	4th	5th	6th
	737	1450	7732	10540	13360	17995

(a) In which group will X be found? Explain.

(b) Predict the formula of X's bromide.

9. Use Coulomb's Law to explain why rubidium tends only to form a +1 ion.

10. Use Coulomb's Law to explain why elements in the same group react in similar ways. You may use a specific example to justify your answer.

11. Identify any (and all) isoelectronic species in the following list; Fe^{2+}, Sc^{3+}, Ca^{2+}, F^{-}, Co^{2+}, Co^{3+},Sr^{2+}, Cu^{+}, Zn^{2+} and Al^{3+}.

12. Arrange the following atoms into order of increasing first ionization energy. Sr, Cs, S, F and As.

13. Explain each of the following observations.
(a) Sodium has a lower first-ionization energy than lithium.

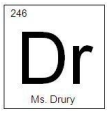
(b) Oxygen has a lower first-ionization energy than nitrogen.

(c) There is a general increase in the first ionization energy from sodium to argon.

(d) Boron has a lower first ionization energy than beryllium.

(e) The first ionization energy of neon (atomic number 10) is significantly higher than that of argon (atomic number 18) but significantly lower than the first ionization energy of helium (atomic number 2), despite all three elements being in the same group.

(f) Helium has the highest first ionization of all the elements shown.



14. Consider the ionization energies of elements X and Y shown below in kJmol-1. X and Y are in the same period of the periodic table and are adjacent to one another in the table.

IE	1st	2nd	3rd	4th	5th	6th	7th	8th	9th
x	1680	3375	6050	8409	11022	15165	17868	92038	106440
y	2080	3950	6122	9370	12180	15239	20000	23068	115375

(a) In which group would one find element X? Explain.

(a) Does element X lie to the right or the left of element Y in the periodic table? Explain.

(b) Which is the first period on the periodic table that these elements could be in? Explain.

(c) Why are the second ionization energies of both elements larger than their respective first ionization energies?

(d) It is found that Y has the largest first ionization energy in the period that it is found. What does this tell us about Y?

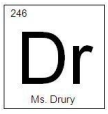
(e) It is found that element Q, which is in the same period as X and Y but lies to the left of element X in the periodic table, only has values for its first four ionization energies. Suggest a reason for this observation.

15. (a) Define first ionization.

(b) Write an equation to show the second ionization energy of calcium.

16. Why does N have a higher first IE than O? Explain using orbital notations.

17. Why does Be have a higher IE than B? Explain using orbital notations.



AP Chemistry: Atomic Structure and PT Multiple Choice

22. $1s^2\ 2s^2 2p^6\ 3s^2 3p^3$

Atoms of an element, X, have the electronic configuration shown above. The compound most likely formed with magnesium, Mg, is…

(A) MgX (B) Mg_2X (C) MgX_2 (D) MgX_3 (E) Mg_3X_2

58. Which of the following represents the ground state electron configuration for the Mn^{3+} ion? (Atomic number Mn = 25)

(A) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^4$ (B) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^5\ 4s^2$ (C) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^2\ 4s^2$

(D) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^8\ 4s^2$ (E) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^3\ 4s^1$

Use the following answers for questions 1 - 3.

(A) F (B) S (C) Mg (D) Ar (E) Mn

1. Forms monatomic ions with 2^- charge in solutions
2. Forms a compound having the formula KXO_4

33. Which of the following conclusions can be drawn from J. J. Thomson's cathode ray experiments?

(A) Atoms contain electrons.

(B) Practically all the mass of an atom is contained in its nucleus.

(C) Atoms contain protons, neutrons, and electrons.

(D) Atoms have a positively charged nucleus surrounded by an electron cloud.

(E) No two electrons in one atom can have the same four quantum numbers.

4. Use these answers for questions 4 - 7.

(A) $1s^2\ 2s^2 2p^5\ 3s^2 3p^5$ (B) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6$ (C) $1s^2\ 2s^2 2p^6 2d^{10}\ 3s^2 3p^6$

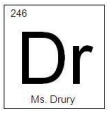
(D) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^5$ (E) $1s^2\ 2s^2 2p^6\ 3s^2 3p^6 3d^3\ 4s^2$

4. An impossible electronic configuration
5. The ground-state configuration for the atoms of a transition element
6. The ground-state configuration of a negative ion of a halogen
7. The ground-state configuration of a common ion of an alkaline earth element

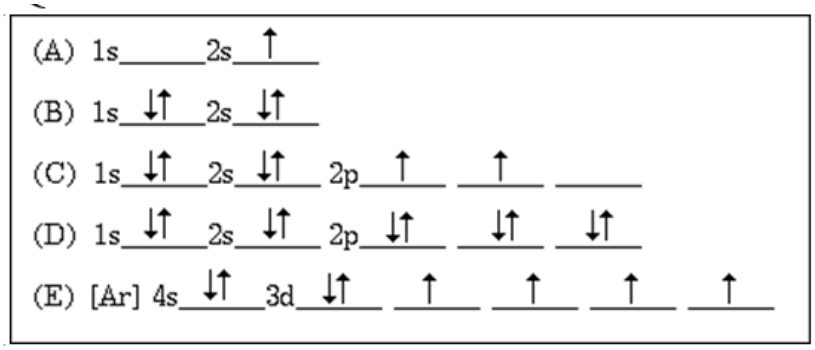
(A) Heisenberg uncertainty principle (B) Pauli exclusion principle (C) Hund's rule

(D) Shielding effect (E) Wave nature of matter

1. Can be used to predict that a gaseous carbon atom in its ground state has electrons with parallel spins
2. Indicates that an atomic orbital can hold no more than two electrons
3. Predicts that it is impossible to determine simultaneously the exact position and the exact velocity of an electron



5. Questions 5-8 refer to atoms for which the occupied atomic orbitals are shown below:

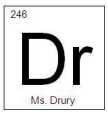


5. Represents an atom that is chemically unreactive
6. Represents an atom in an excited state
7. Represents an atom that has four valence electrons.
8. Represents an atom of a transition metal.

51. Which of the following is a correct interpretation of the results of Rutherford's experiments in which gold atoms were bombarded with alpha particles?
(A) Atoms have equal numbers of positive and negative charges.
(B) Electrons in atoms are arranged in shells.
(C) Neutrons are at the center of an atom.
(D) Neutrons and protons in atoms have nearly equal mass.
(E) The positive charge of an atom is concentrated in a small region.

19. Which of the following represents a pair of isotopes?

		Atomic Number	Mass Number
(A)	I.	6	14
	II.	7	14
(B)	I.	6	7
	II.	14	14
(C)	I.	6	14
	II.	14	28
(D)	I.	7	13
	II.	7	14
(E)	I.	8	16
	II.	16	20



Questions 1-2

Consider atoms of the following elements. Assume that the atoms are in the ground state.

(A) S
(B) Ca
(C) Ga
(D) Sb
(E) Br

1. The atom that contains exactly two unpaired electrons.
2. The atom that contains only one electron in the highest occupied energy sublevel.

17. In which of the following groups are the three species isoelectronic, i.e. have the same number of… electrons?

(A) S^{2-}, K^{+}, Ca^{2+}
(B) Sc, Ti, V^{2+}
(C) O^{2-}, S^{2-}, Cl^{-}
(D) Mg^{2+}, Ca^{2+}, Sr^{2+}
(E) Cs, Ba^{2+}, La^{3+}



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