Chapters

3 and 4

Atoms

&

Arrangement of electrons

Name ______________________

Period_____

Name: ______________________ Period: _______ Date ______________

Worksheet: Atoms, Isotopes, and Ions

Atoms

- The number of protons in an atom determines the identity of the atom.
  Atomic Number = number of Protons
- In a neutral atom, the number of positive protons equals the number of negative electrons.
  Number of Protons = number of Electrons
- Protons and neutrons both have a mass of 1 amu. The mass of the electron is negligible compared to the mass of the proton and neutron. Thus the mass number, or the mass of the atom, is the sum of the number of protons and neutrons.
  Mass number = number of Protons + number of Neutrons

Name	*Symbol*	*Atomic #*	*Mass #*	*# Protons*	*# Neutrons*	*# Electrons*
Selenium					46	
			222	86		
					118	79
		11			12	

Isotopes

- The number of neutrons in any specific type of atom can vary. Atoms of the same element with different numbers of neutrons are called isotopes.
- Isotopes are distinguished from each other by including the mass number with the name or symbol.

Name	*Symbol*	*Atomic #*	*Mass #*	*# Protons*	*# Neutrons*	*# Electrons*
	$^{235}_{92}U$					
	$^{238}_{92}U$					
Carbon-12						
Carbon-13						

Name: ______________________________ Period: ______ Date ____________

Ions

- As we have seen, in a neutral atom, the number of protons and the number of electrons is equal.
- Atoms can gain or lose electrons to become ions. Ions are charged atoms resulting from the difference in number of positive protons and negative electrons.
- A cation is a positive ion. A cation results when an atom loses electrons. Number of Protons is GREATER than the Number of Electrons
  Cations keep the name that they have on the periodic table. For example: when lithium loses one electron we call it a lithium ion.
- An anion is a negative ion. An anion results when an atom gains electrons. Number of Electrons is GREATER than the number of Protons
  Anions lose the last couple of letters from the name on the periodic table and replace it with an "-ide" ending. For example: When chlorine gains one electron we call it the chloride ion.
  Ions are distinguished from atoms by including the ion charge as a superscript in the symbol.

Name	*Symbol*	*Atomic #*	*Mass #*	*# Protons*	*# Neutrons*	*# Electrons*	*Cation or Anion?*
	Al^{3+}				14		
Iron ion			56			24	
				15	15	18	
	F^{1-}		19				

NAME:______________________________

Reminders:
1. In a neutral atom the number of protons equals the number of electrons.
2. An atom can NEVER gain or lose protons
3. The number of protons equals the atomic number

Ion Practice Set

1. What is an ion?

2. What does the number next to the ions signify?

Complete the following table, using the periodic table in the back of your book.

	ELEMENT NAME	ION SYMBOL	NUMBER OF PROTONS	NUMBER OF ELECTRONS	NUMBER OF ELECTRONS LOST OR GAINED
ex	Fluoride	F^-	9	10	gained one
1			53	54	
2			16		gained two
3	potassium				lost one
4		Ca^{+2}			
5			35	36	
6		Sr^{+2}			
7		H^+			
8			8		gained two
9			12		lost two
10	aluminum			10	
11			34	36	
12		H^-			
13	lithium				lost one
14		Rb^+			
15			17	18	

Name ______________________________ Date______________________

Molar Conversion Homework

Write down the representative particle for the following:

Zn ________________________ N_2 ___________________________

CO_2 ______________________ NaCl __________________________

1. Determine the number of grams in one mole of carbon dioxide gas.

2. How many grams are in 0.40 moles of carbon dioxide gas?

3. How many grams are in 5.0 moles of carbon dioxide gas?

4. How many molecules are in one mole of carbon dioxide gas?

5. How many atoms are in 3.5 moles of carbon dioxide gas?

6. How many grams are in one mole of zinc?

7. How many grams are in 0.40 moles of zinc?

8. How many grams are in 4.0 moles of zinc?

9. How many atoms are in 0.50 moles of zinc?

10. Determine the number of grams in one mole of $CaCO_3$.

11. How many grams are in 10.0 moles of $CaCO_3$

12. How many grams are in 0.025 moles of $CaCO_3$?

13. How many moles are in 50.0 grams of $CaCO_3$?

14. How many formula units are in 0.60 moles of $CaCO_3$?

15. What is the molar mass of $Ca_3(PO_4)_2$?

16. How many atoms are in one formula unit of $Ca_3(PO_4)_2$?

17. How many atoms are in 44.0 moles of $Ca_3(PO_4)_2$?

18. How many grams are in 6.54 moles of $Ca_3(PO_4)_2$?

19. How many moles are in 4.5×10^{30}formula units of $Ca_3(PO_4)_2$?

20. What is the formula mass of NaCl?

21. How many formula units are in one formula mass of NaCl?

22. How many formula units are in one molar mass of NaCl?

23. What is the molar mass of NaCl?

24. How many formula units are in 35.0g of NaCl?

Weighted Average Masses – Atomic Masses

Most elements have more than one isotope that occurs naturally. The different isotopes have a different number of neutrons and therefore have different masses. The atomic masses on the periodic table are the weighted averages of the different isotopes. To calculate a weighted average you 1) multiply each % abundance by the mass of the corresponding isotope, 2) then add up the values to get the weighted average.

Example:	1) multiply % by mass	2) add together answers
Copper-63 :	69.1% x 62.93 g/mol = 43.48	43.48 + 20.06 = 63.54 g/mol
Copper-65 :	30.9% x 64.93 g/mol = 20.06	

Calculate the weighted average masses or molar masses of the following elements.

1. lithium-6 6.0151 g/mol 7.5 %
   lithium-7 7.0160 g/mol 92.5 %

2. boron-10 10.0129 g/mol 19.6 %
   boron-11 11.0093 g/mol 80.4 %

3. oxygen-16 15.9949 g/mol 99.76 %
   oxygen-17 16.9991 g/mol 0.04 %
   oxygen-18 17.9992 g/mol 0.20 %

4. magnesium-24 23.9850 g/mol 78.99 %
   magnesium-25 24.9858 g/mol 10.00 %
   magnesium-26 25.9826 g/mol 11.01 %

5. silicon-28 27.9769 g/mol 92.23 %
   silicon-29 28.9765 g/mol 4.67 %
   silicon-30 29.9738 g/mol 3.10 %

Name ____________________ Date __________ Class ________

Practice Problems

1. How many protons and electrons are present in a vanadium atom?

2. How many protons and electrons are present in a nitrogen atom?

3. How many protons and electrons are present in an argon atom?

4. How many protons and electrons are present in a potassium atom?

5. How many protons and electrons are present in a platinum atom?

6. What is the name of the element that has atoms that contain 5 protons?

7. What is the name of the element that has atoms that contain 17 protons?

8. What is the name of the element that has atoms that contain 25 protons?

9. What is the name of the element that has atoms that contain 82 protons?

10. What is the name of the element that has atoms that contain 92 protons?

11. Write the chemical symbol for the ion with 12 protons and 10 electrons.

12. Write the chemical symbol for the ion with 74 protons and 68 electrons.

13. Write the chemical symbol for the ion with 95 protons and 89 electrons.

14. Write the chemical symbol for the ion with 33 protons and 36 electrons.

Name ______________________ Date __________ Class __________

3–3 *Practice Problems (continued)*

15. Write the chemical symbol for the ion with 29 protons and 27 electrons.

16. How many protons, neutrons, and electrons are present in the $^{59}_{28}Ni^{2+}$ ion?

17. How many protons, neutrons, and electrons are present in the $^{91}_{40}Zr^{4+}$ ion?

18. How many protons, neutrons, and electrons are present in the $^{140}_{58}Ce^{3+}$ ion?

19. How many protons, neutrons, and electrons are present in the $^{79}_{34}Se^{2-}$ ion?

20. How many protons, neutrons, and electrons are present in the $^{45}_{21}Sc^{3+}$ ion?

21. How many protons, neutrons, and electrons are present in the $^{13}_{6}C^{4-}$ ion?

22. Write the complete chemical symbol for the ion with 84 protons, 125 neutrons, and 80 electrons.

23. Write the complete chemical symbol for the ion with 27 protons, 32 neutrons, and 25 electrons.

24. Write the complete chemical symbol for the ion with 73 protons, 108 neutrons, and 68 electrons.

25. Write the complete chemical symbol for the ion with 31 protons, 39 neutrons, and 28 electrons.

ATOMIC HISTORY SUMMARY

Resource packet

Democritus 4th century B.C.
- Everything comprised of *Atoms*

Alchemists
- Tried to change metals into Gold

Antoine Lavoisier 1770
- *Law of Conservation of Matter* - Matter is neither created nor destroyed but can be change from one form to another

Joseph Proust 1799
- *Law of Definite Proportions(Constant Composition)* - Proportion of elements by mass in all pure compounds is always the same

John Dalton 1803
- *Law of Multiple Proportions* - Elements can combine in different mass ratios -to form different pure compounds
- Proposed *1st Atomic Theory*
  - All elements are composed of indivisible and indestructible atoms
  - All atoms of same element are exactly alike
  - Atoms of different elements are different in composition, mass, and characteristics
  - Compounds are formed by joining atoms of one or more element
  - Atoms join in definite proportions but the atoms are not altered
- *Dalton's Model of Atom - Solid Sphere*
  - Remained unblemished for 100 years
- Problems in Dalton's Theory
  - Atoms not indivisible
    - Not the smallest particle
  - Atomic masses of atoms of same element differ slightly
    - Average mass always same
  - Atoms of elements can have different masses - ISOTOPES - atoms of the same element with different number of neutrons

William Crookes 1770
- Worked with electrons and the Cathode Ray Tube

Henry Becquerel 1890
- Worked with and *discovered radiation*

J.J. Thompson 1898
- *Discovered electron* using Cathode Ray Tube
- Determined mass of electron to be 1/1837 the mass of a Hydrogen atom
- Hydrogen atom - smallest and lightest element
- *Thompson's Model of Atom*
  - Atom is a mass of positive charged dough with thousands of electrons like raisins in the dough – known as the "*Raisin Bun Model*" or "Plum Pudding"
- Problems in Thompson's Model
  - Too many electrons to the atom (no atom has more than 100 or so electrons)
  - Not all in one mass (electrons are outside the nucleus)

Ernest Rutherford 1911

- *Atom is mostly empty space* with nucleus in center
- Nucleus has a positive charge
- Gold Foil Experiment
    - Most alpha particles passed through foil → atoms must be made of empty space
- *Rutherford's Model of Atom*
    - Positive nucleus in center
    - Negative electrons outside nucleus
    - Rest is empty space
- Problems in Rutherford's Model
    - Positive and negative would attract each other and electron would spiral into nucleus and destroy the atom

Robert Millican 1911

- *Measured mass and charge on electron* as 1.6 x 10-19 coulombs by using "Oil Drop Experiment"

Neils Bohr 1913

- *Electrons are outside the nucleus and orbit in distinct energy paths called orbits*
- Each shell has an energy level which keeps the electrons in their orbit.
- Outer levels have more energy than inner levels
- Electron levels shell labeled: KLMNO
- number of electrons can occupy an energy level depending on that levels energy
    - K - 2 electrons
    - L - 8 electrons
    - M - 18 electrons
    - N - 32 electrons
    - O - 50 electrons

- Normally electrons exist at a Ground State
    - Occupy lowest available energy level
    - Electrons can move to higher energy levels by absorbing energy
    - Electrons can absorb only a specific amount of energy to move to an excited state
    - Electrons always drop back from excited state to ground state
    - When electron drops state, it releases a quantum of energy, called photon
    - Some of the energy is our visible light spectrum
- Problems in Bohr's Model

    Only explained properties of Hydrogen, not other elements

Henry Moseley 1914

- *Atomic Number* - Each element has set number of protons that cannot be changed or element will be different
- Nucleus is made up of protons and neutrons
    - They give the atom its mass
    - Neutron has a mass equal to a proton

Werner Heisenberg 1927

- *Uncertainty principle* – path taken by the electron can not be predicted but only estimated
- Modern atomic theory - Electron cloud

James Chadwick 1932

- *Discovered the neutron* - Bombarded a Beryllium nucleus with an alpha particle — a particle without a charge was emitted a neutron

Name ______________________ Date ____________ Class ____________

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. Why is Democritus's view of matter considered only an idea, while Dalton's view is considered a theory?

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2. Give an example of a chemical or physical process that illustrates the law of conservation of mass.

____

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3. State two principles from Dalton's atomic theory that have been revised as new information has become available.

____

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4. The formation of water according to the equation

$$2H_2 + O_2 \rightarrow 2H_2O$$

shows that 2 molecules (made of 4 atoms) of hydrogen and 1 molecule (made of 2 atoms) of oxygen produce 2 molecules of water. The total mass of the product, water, is equal to the sum of the masses of each of the reactants, hydrogen and oxygen. What parts of Dalton's atomic theory are illustrated by this reaction? What law does this reaction illustrate?

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PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

5. ______________ If 3 g of element C combine with 8 g of element D to form compound CD, how many grams of D are needed to form compound CD_2?

6. A sample of baking soda, $NaHCO_3$, *always* contains 27.37% by mass of sodium, 1.20% of hydrogen, 14.30% of carbon, and 57.14% of oxygen.

a. Which law do these data illustrate?

____

b. State the law.

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7. Nitrogen and oxygen combine to form several compounds, as shown by the following table.

Compound	Mass of nitrogen that combines with 1 g oxygen (g)
NO	1.70
NO_2	0.85
NO_4	0.44

Calculate the ratio of the masses of nitrogen in each of the following:

______ **a.** $\frac{NO}{NO_2}$ ______ **b.** $\frac{NO_2}{NO_4}$ ______ **c.** $\frac{NO}{NO_4}$

d. Which law do these data illustrate?

____

Name ______________________ Date ______________ Class ______________

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

SECTION 2

SHORT ANSWER Answer the following questions in the space provided.

1. In cathode-ray tubes, the cathode ray is emitted from the negative electrode, which is called the

______________________.

2. The smallest unit of an element that can exist either alone or in molecules containing the same or different elements is the ______________________.

3. A positively charged particle found in the nucleus is called a(n) ______________________.

4. A nuclear particle that has no electrical charge is called a(n) ______________________.

5. The subatomic particles that are least massive and most massive, respectively, are the

______________________ and ______________________.

6. A cathode ray produced in a gas-filled tube is deflected by a magnetic field. A wire carrying an electric current can be pulled by a magnetic field. A cathode ray is deflected away from a negatively charged object. What property of the cathode ray is shown by these phenomena?

____

7. How would the electrons produced in a cathode-ray tube filled with neon gas compare with the electrons produced in a cathode-ray tube filled with chlorine gas?

____

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____

8. a. Is an atom positively charged, negatively charged, or neutral?

____

b. Explain how an atom can exist in this state.

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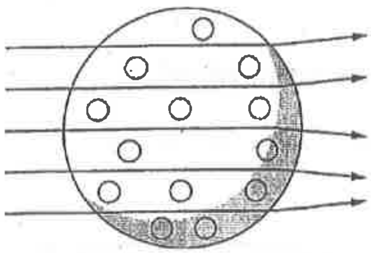
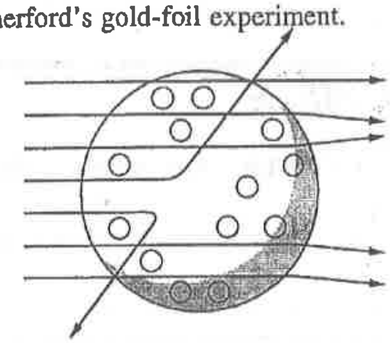
____

Name ______________________ Date ____________ Class ____________

SECTION 2 continued

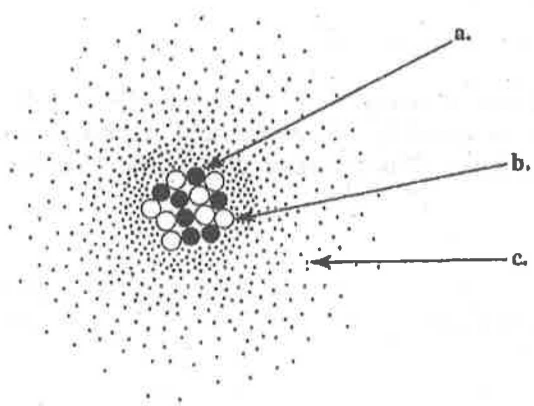
9. Below are illustrations of two scientists' conceptions of the atom. Label the electrons in both illustrations with a − sign and the nucleus in the illustration to the right with a + sign. On the lines below the figures, identify which illustration was believed to be correct before Rutherford's gold-foil experiment and which was believed to be correct after Rutherford's gold-foil experiment.





a. ______________________ b. ______________________

10. In the space provided, describe the locations of the subatomic particles in the labeled model of an atom of nitrogen below, and give the charge and relative mass of each particle.



a. proton

____

____

b. neutron

____

____

c. electron (a possible location of this particle)

____

____

Name ______________________ Date ______________ Class ______________

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

SECTION 3

SHORT ANSWER Answer the following questions in the space provided.

1. Explain the difference between the *mass number* and the *atomic number* of a nuclide.

2. Why is it necessary to use the average atomic mass of all isotopes, rather than the mass of the most commonly occurring isotope, when referring to the atomic mass of an element?

3. How many particles are in 1 mol of carbon? 1 mol of lithium? 1 mol of eggs? Will 1 mol of each of these substances have the same mass?

4. Explain what happens to each of the following as the atomic masses of the elements in the periodic table increase:

a. the number of protons

b. the number of electrons

c. the number of atoms in 1 mol of each element

Name ______________________ Date ____________ Class ____________

SECTION 3 continued

5. Use a periodic table to complete the following chart:

Element	Symbol	Atomic number	Mass number
Europium-151			
	$^{109}_{47}Ag$		
Tellurium-128			

6. List the number of protons, neutrons, and electrons found in zinc-66.

_____ protons

_____ neutrons

_____ electrons

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

7. ______________ What is the mass in grams of 2.000 mol of oxygen atoms?

8. ______________ How many moles of aluminum exist in 100.0 g of aluminum?

9. ______________ How many atoms are in 80.45 g of magnesium?

10. ______________ What is the mass in grams of 100 atoms of the carbon-12 isotope?

Name ______________________ Date ______________ Class ______________

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

MIXED REVIEW

SHORT ANSWER Answer the following questions in the space provided.

1. The element boron, B, has an atomic mass of 10.81 amu according to the periodic table. However, no single atom of boron has a mass of exactly 10.81 amu. How can you explain this difference?

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2. How did the outcome of Rutherford's gold-foil experiment indicate the existence of a nucleus?

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3. Ibuprofen, $C_{13}H_{18}O_2$, that is manufactured in Michigan contains 75.69% by mass carbon, 8.80% hydrogen, and 15.51% oxygen. If you buy some ibuprofen for a headache while you are on vacation in Germany, how do you know that it has the same percentage composition as the ibuprofen you buy at home?

____

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____

____

4. Complete the following chart, using the atomic mass values from the periodic table:

Compound	Mass of Fe (g)	Mass of O (g)	Ratio of O:Fe
FeO			
Fe_2O_3			
Fe_3O_4			

Name ______________________ Date ____________ Class ____________

MIXED REVIEW continued

5. Complete the following table:

Element	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
Sodium			22			
	F	9	19			
			80		45	
			40	20		
		1			0	
			222			86

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

6. ______________ **a.** How many atoms are there in 2.50 mol of hydrogen?

______________ **b.** How many atoms are there in 2.50 mol of uranium?

______________ **c.** How many moles are present in 107 g of sodium?

7. A certain element exists as three natural isotopes, as shown in the table below.

Isotope	Mass (amu)	Percent natural abundance	Mass number
1	19.99244	90.51	20
2	20.99395	0.27	21
3	21.99138	9.22	22

______________ Calculate the average atomic mass of this element.

Chapter 4

Name ______________________ Date ______________ Class ______________

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 1

SHORT ANSWER Answer the following questions in the space provided.

1. In what way does the photoelectric effect support the particle theory of light?

____

____

____

2. What is the difference between the ground state and the excited state of an atom?

____

____

3. Under what circumstances can an atom emit a photon?

____

____

4. How can the energy levels of the atom be determined by measuring the light emitted from an atom?

____

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____

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5. Why does electromagnetic radiation in the ultraviolet region represent a larger energy transition than does radiation in the infrared region?

____

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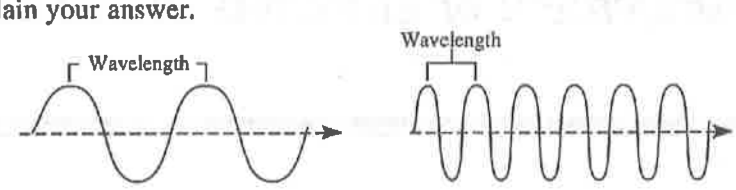
SECTION 1 continued

6. Which of the waves shown below has the higher frequency? (The scale is the same for each drawing.) Explain your answer.



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7. How many different photons of radiation were emitted from excited helium atoms to form the spectrum shown below? Explain your answer.

Spectrum for helium

____

____

____

____

____

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

8. ______________ What is the frequency of light that has a wavelength of 310 nm?

9. ______________ What is the wavelength of electromagnetic radiation if its frequency is 3.2×10^{-2} Hz?

Name ______________________ Date ______________ Class ______________

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 2

SHORT ANSWER **Answer the following questions in the space provided.**

1. ______ How many quantum numbers are used to describe the properties of electrons in atomic orbitals?

(a) 1 (c) 3
(b) 2 (d) 4

2. ______ A spherical electron cloud surrounding an atomic nucleus would best represent

(a) an *s* orbital. (c) a combination of two different *p* orbitals.
(b) a *p* orbital. (d) a combination of an *s* and a *p* orbital.

3. ______ How many electrons can an energy level of $n = 4$ hold?

(a) 32 (c) 8
(b) 24 (d) 6

4. ______ How many electrons can an energy level of $n = 2$ hold?

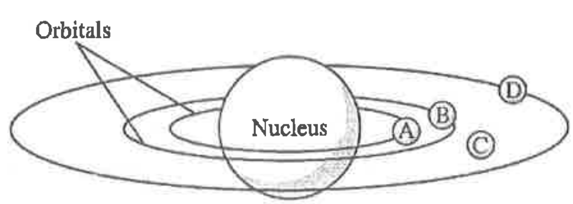
(a) 32 (c) 8
(b) 24 (d) 6

5. ______ Compared with an electron for which $n = 2$, an electron for which $n = 4$ has more

(a) spin. (c) energy.
(b) particle nature. (d) wave nature.

6. ______ According to Bohr, which is the point in the figure below where electrons cannot reside?

(a) point A (c) point C
(b) point B (d) point D



7. ______ According to the quantum theory, point D in the above figure represents

(a) the fixed position of an electron.
(b) the farthest position from the nucleus that an electron can achieve.
(c) a position where an electron probably exists.
(d) a position where an electron cannot exist.

8. How did de Broglie conclude that electrons have a wave nature?

9. Identify each of the four quantum numbers and the properties to which they refer.

10. How did the Heisenberg uncertainty principle contribute to the idea that electrons occupy "clouds," or "orbitals"?

11. Complete the following table:

Principal quantum number, *n*	Number of sublevels	Types of orbitals
1		
2		
3		
4		

Name ____________________ Date ____________ Class ____________

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 3

SHORT ANSWER **Answer the following questions in the space provided.**

1. State the Pauli exclusion principle, and use it to explain why electrons in the same orbital must have opposite spin states.

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2. Explain the conditions under which the following orbital notation for helium is possible:

↑	↑
1*s*	2*s*

____

____

Write the ground-state electron configuration and orbital notation for each of the following atoms:

3. Phosphorus

4. Nitrogen

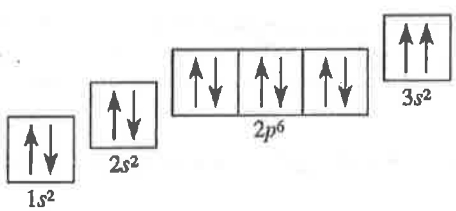
5. Potassium

6. Aluminum

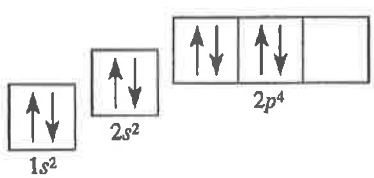
7. Argon

8. Boron

9. Which guideline, Hund's rule or the Pauli exclusion principle, is violated in the following orbital diagrams?



a. ______________________



b. ______________________

Name ______________________ Date ______________ Class ______________

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

MIXED REVIEW

SHORT ANSWER Answer the following questions in the space provided.

1. Under what conditions is a photon emitted from an atom?

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2. What do quantum numbers describe?

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3. What is the relationship between the principal quantum number and the electron configuration?

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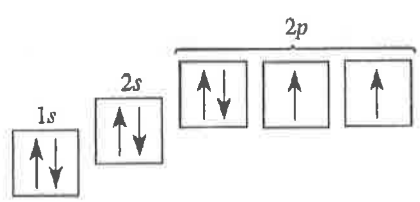


4. In what way does the figure above illustrate Hund's rule?

____

____

____

5. In what way does the figure above illustrate the Pauli exclusion principle?

____

____

Name ______________________ Date ______________ Class ______________

MIXED REVIEW continued

6. Elements of the fourth and higher main-energy levels do not seem to follow the normal sequence for filling orbitals. Why is this so?

____

____

____

7. How do electrons create the colors in a line-emission spectrum?

____

____

____

8. Write the ground-state electron configuration of the following atoms:

a. Carbon

____

b. Potassium

____

c. Gallium

____

d. Copper

____

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

9. ______________ What is the wavelength of light that has a frequency of 3×10^{-4} Hz in a vacuum?

10. ______________ What is the energy of a photon that has a frequency of 5.0×10^{14} Hz?

Waves & Particles
Electrons In Atoms

Where does our understanding of electrons come from?

How does light travel from the sun to us?

An electromagnetic wave consists of electric and magnetic fields that OSCILLATE at right angles to each other.

X-rays, GAMMA RAYS, radiowaves and MICROWAVES are some examples.

DRAW A WAVE HERE AND LABEL ALL OF THE PARTS:

Wavelength (λ) lambda - length of one complete wave

Frequency (ν) nu - # of waves that pass a point during a certain time period

hertz (Hz) = 1/s Hz = (cycles per second)

Visible light has a frequency of 4×10^{14}Hz to 7×10^{14} Hz

Amplitude (A) - distance from the origin to the trough or crest

Wavelength:http://www.colorado.edu/physics/2000/waves_particles/lightspeed-1.html

Frequency:http://id.mind.net/~zona/mstm/physics/waves/partsOfAWave/waveParts.htm

c= speed of light 3.00×10^{8} m/s or 3.00×10^{10}cm/s

However, the speed of light does vary, depending on the medium. Albert Einstein proved that light is the fastest moving thing in a vacuum, with its speed being roughly 186,000 miles per second.

In water, the speed of light slows to 140,000 miles per second. Its speed is 124,000 per second in glass. Through a diamond, the speed of light is 77,500 miles per second.

The speed of light = wavelength x frequency $c = \lambda\nu$

λ means *lambda* and represents wavelength ν means *nu* and represents frequency

Frequency is measured in cycles per second and is represented by Hz = hertz.

s^{-1} also represents frequency and is called a reciprocal second.

Frequency & wavelength are inversely proportional

EX: Find the frequency of a photon with a wavelength of 434 nm.

Calculate the wavelength of a yellow light emitted by a sodium lamp if the frequency of the radiation is 5.10×10^{14} Hz. (Use $c = 3.00 \times 10^{10}$ cm/s)

What is the wavelength (in cm) of radiation with a frequency of $1.50 \times 10^{18}\ s^{-1}$?

What is the frequency if wavelength is 5.00×10^{-6} cm?

Try this one on your own: If a light has a wavelength of 633 nanometers, what is the frequency?

Write information regarding the electromagnetic spectrum in the space below:

- What does ROY G BIV represent?
- Which has more energy? Red or violet light
- Which color of visible light has the longest wavelength?
- Why can red light be used in booths that develop film?

What if you could see in infrared? http://spaceplace.jpl.nasa.gov/en/kids/sirtf1/sirtf_action.shtml

Everything emits infrared light. Even things found in your freezer emit it. Infrared can be found in anything above absolute zero. The amount of infrared light that an object emits depends on its temperature. A human gives off much more infrared light than an ice cube would.

The Nobel Prize in Physics 1901 Wilhelm Conrad Roentgen (1845-1923) On November 8, 1895, his attention was drawn to a glowing fluorescent screen on a nearby table. Roentgen immediately determined that the fluorescence was caused by invisible rays. Surprisingly, these mysterious rays penetrated the opaque black paper wrapped around the tube. Roentgen had discovered X rays, a momentous event that instantly revolutionized the field of physics and medicine. For his discovery, Roentgen received the first Nobel Prize in physics in 1901. When later asked what his thoughts were at the moment of his discovery, he replied "I didn't think, I investigated. "It was the crowning achievement in a career beset by more than its share of difficulties.

Bohr Model e^- exist only in orbits with specific amounts of energy called energy levels

Therefore...

e^- can only gain or lose certain amounts of energy

only certain photons are produced

The amount of energy is specific. This amount of energy is called a QUANTUM. Energies of electrons are said to be QUANTIZED.

Quantum Theory

Planck (1900) A turtle sitting on a staircase can take on only certain discrete energies

Energy is required to move the turtle up the steps (absorption)

Energy is released when the turtle moves down the steps (emission)

Only discrete amounts of energy are absorbed or released (energy is said to be **quantized**)

A turtle on a ramp would represent continuous. It can move anywhere on the ramp.

http://www.upscale.utoronto.ca/PVB/Harrison/BohrModel/Flash/BohrModel.html

Bohr Model

Energy staircase diagram for atomic hydrogen

bottom step is called the **ground state**

higher steps are called **excited states**

Line Spectra = the fingerprint of an element

http://www.colorado.edu/physics/2000/quantumzone/index.html

http://jersey.uoregon.edu/vlab/elements/Elements.html

Quantum Theory

Planck (1900)

Observed - emission of light from hot objects

Concluded - energy is emitted in small, specific amounts (quanta)

Quantum - minimum amount of energy change

Quantum Theory

Einstein (1905)

Observed - photoelectric effect

Photoelectric effect
http://www.explorelearning.com/index.cfm?method=cResource.dspView&ResourceID=491

Einstein (1905)

Concluded - light has properties of both waves and particles "wave-particle duality"

Photon - particle of light that carries a quantum of energy

Quantum Theory

energy (J, joules)

h:Planck's constant (6.6262×10^{-34} J·s)

v: frequency (Hz)

Quantum Theory

EX: Find the energy of a red photon with a frequency of 4.57×10^{14} Hz.

Louis de Broglie: Stated that all matter has a wave like motion.

Black light information:

If you turn on a black light bulb in a dark room, what you can see from the bulb is a purplish glow. What you cannot see is the **ultraviolet light** that the bulb is also producing.

Our eyes can see visible light in a spectrum ranging from red through orange, yellow, green, blue and violet. Above violet is ultraviolet light, which we cannot see. A black light bulb produces **UVA** light (as opposed to UVB light, which is much more harmful).

Ultraviolet light frequencies are above violet on the spectrum and are invisible to the human eye.

What you see glowing under a black light, whether on a fluorescent poster or an invisible hand stamp or a newly washed white T-shirt, are **phosphors**.

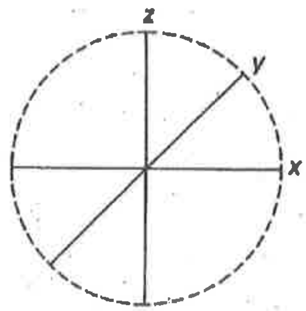
A phosphor is any substance that emits visible light in response to some sort of **radiation**. A phosphor converts the energy in the UV radiation from a black light into visible light.

Why Is My Shirt Glowing?

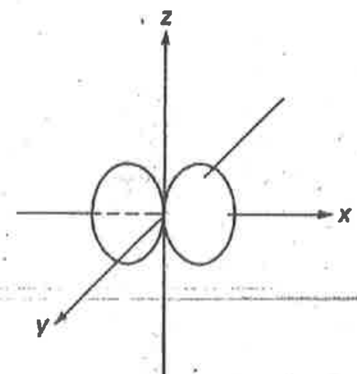
White T-shirts and socks normally glow under a black light because modern detergents contain phosphors that convert UV light into white light. This makes whites look "whiter than white" in normal sunlight. What you are seeing in sunlight is the normal reflection of visible white light from the cloth, as well as the emission of white light that the phosphors create from UV light in sunlight. The T-shirt really is whiter than white!

Name____

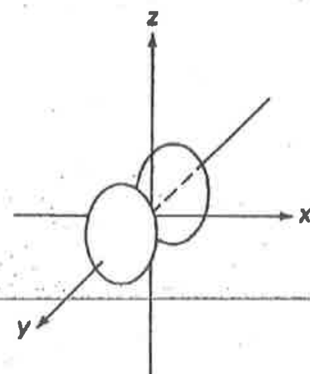
ATOMIC ORBITALS



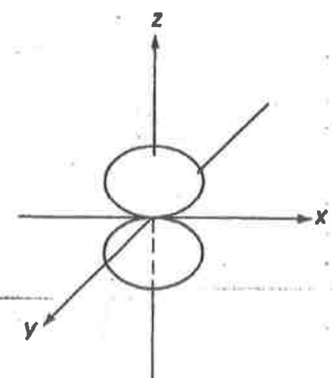
1s orbital



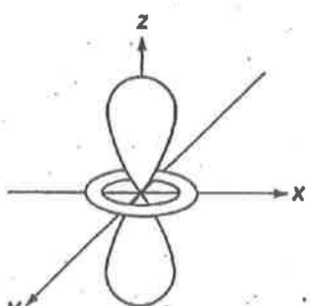
$2p_x$ orbital



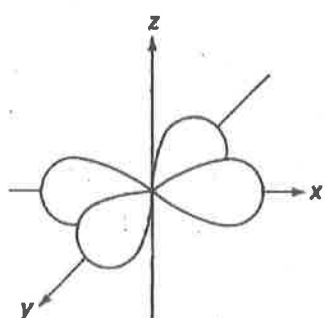
$2p_y$ orbital



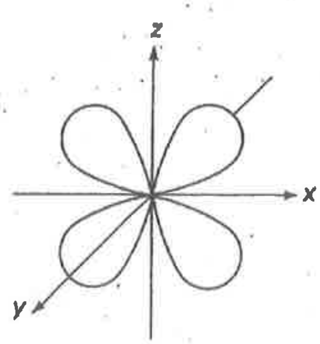
$2p_z$ orbital



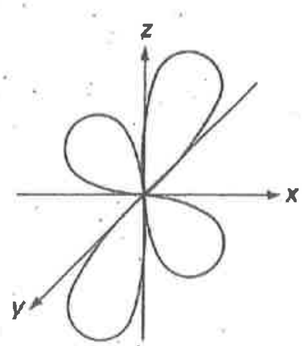
d_{z^2}



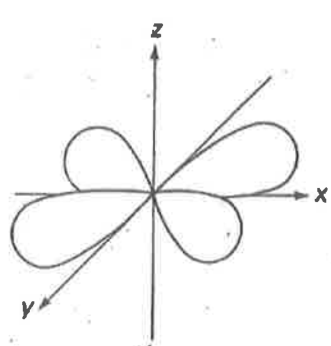
$d_{x^2-y^2}$



d_{xz}



d_{yz}



d_{xy}

Name____

ELECTRON CONFIGURATIONS OF THE ELEMENTS

Element	1s	2s	2p	3s	3p	3d	4s	4p	4d	4f	5s	5p	5d	5f	6s	6p	6d	6f	7s
1 H	1																		
2 He	2																		
3 Li	2	1																	
4 Be	2	2																	
5 B	2	2	1																
6 C	2	2	2																
7 N	2	2	3																
8 O	2	2	4																
9 F	2	2	5																
10 Ne	2	2	6																
11 Na	2	2	6	1															
12 Mg	2	2	6	2															
13 Al	2	2	6	2	1														
14 Si	2	2	6	2	2														
15 P	2	2	6	2	3														
16 S	2	2	6	2	4														
17 Cl	2	2	6	2	5														
18 Ar	2	2	6	2	6														
19 K	2	2	6	2	6		1												
20 Ca	2	2	6	2	6		2												
21 Sc	2	2	6	2	6	1	2												
22 Ti	2	2	6	2	6	2	2												
23 V	2	2	6	2	6	3	2												
24 Cr	2	2	6	2	6	5	1												
25 Mn	2	2	6	2	6	5	2												
26 Fe	2	2	6	2	6	6	2												
27 Co	2	2	6	2	6	7	2												
28 Ni	2	2	6	2	6	8	2												
29 Cu	2	2	6	2	6	10	1												
30 Zn	2	2	6	2	6	10	2												
31 Ga	2	2	6	2	6	10	2	1											
32 Ge	2	2	6	2	6	10	2	2											
33 As	2	2	6	2	6	10	2	3											
34 Se	2	2	6	2	6	10	2	4											
35 Br	2	2	6	2	6	10	2	5											
36 Kr	2	2	6	2	6	10	2	6											
37 Rb	2	2	6	2	6	10	2	6			1								
38 Sr	2	2	6	2	6	10	2	6			2								
39 Y	2	2	6	2	6	10	2	6	1		2								
40 Zr	2	2	6	2	6	10	2	6	2		2								
41 Nb	2	2	6	2	6	10	2	6	4		1								
42 Mo	2	2	6	2	6	10	2	6	5		1								
43 Tc	2	2	6	2	6	10	2	6	6		1								
44 Ru	2	2	6	2	6	10	2	6	7		1								
45 Rh	2	2	6	2	6	10	2	6	8		1								
46 Pd	2	2	6	2	6	10	2	6	10										
47 Ag	2	2	6	2	6	10	2	6	10		1								
48 Cd	2	2	6	2	6	10	2	6	10		2								
49 In	2	2	6	2	6	10	2	6	10		2	1							
50 Sn	2	2	6	2	6	10	2	6	10		2	2							
51 Sb	2	2	6	2	6	10	2	6	10		2	3							
52 Te	2	2	6	2	6	10	2	6	10		2	4							
53 I	2	2	6	2	6	10	2	6	10		2	5							
54 Xe	2	2	6	2	6	10	2	6	10		2	6							
55 Cs	2	2	6	2	6	10	2	6	10		2	6			1				
56 Ba	2	2	6	2	6	10	2	6	10		2	6			2				
57 La	2	2	6	2	6	10	2	6	10		2	6	1		2				
58 Ce	2	2	6	2	6	10	2	6	10	2	2	6			2				
59 Pr	2	2	6	2	6	10	2	6	10	3	2	6			2				
60 Nd	2	2	6	2	6	10	2	6	10	4	2	6			2				
61 Pm	2	2	6	2	6	10	2	6	10	5	2	6			2				
62 Sm	2	2	6	2	6	10	2	6	10	6	2	6			2				
63 Eu	2	2	6	2	6	10	2	6	10	7	2	6			2				
64 Gd	2	2	6	2	6	10	2	6	10	7	2	6	1		2				
65 Tb	2	2	6	2	6	10	2	6	10	9	2	6			2				
66 Dy	2	2	6	2	6	10	2	6	10	10	2	6			2				
67 Ho	2	2	6	2	6	10	2	6	10	11	2	6			2				
68 Er	2	2	6	2	6	10	2	6	10	12	2	6			2				
69 Tm	2	2	6	2	6	10	2	6	10	13	2	6			2				
70 Yb	2	2	6	2	6	10	2	6	10	14	2	6			2				
71 Lu	2	2	6	2	6	10	2	6	10	14	2	6	1		2				
72 Hf	2	2	6	2	6	10	2	6	10	14	2	6	2		2				
73 Ta	2	2	6	2	6	10	2	6	10	14	2	6	3		2				
74 W	2	2	6	2	6	10	2	6	10	14	2	6	4		2				
75 Re	2	2	6	2	6	10	2	6	10	14	2	6	5		2				
76 Os	2	2	6	2	6	10	2	6	10	14	2	6	6		2				
77 Ir	2	2	6	2	6	10	2	6	10	14	2	6	7		2				
78 Pt	2	2	6	2	6	10	2	6	10	14	2	6	9		1				
79 Au	2	2	6	2	6	10	2	6	10	14	2	6	10		1				
80 Hg	2	2	6	2	6	10	2	6	10	14	2	6	10		2				
81 Tl	2	2	6	2	6	10	2	6	10	14	2	6	10		2	1			
82 Pb	2	2	6	2	6	10	2	6	10	14	2	6	10		2	2			
83 Bi	2	2	6	2	6	10	2	6	10	14	2	6	10		2	3			
84 Po	2	2	6	2	6	10	2	6	10	14	2	6	10		2	4			
85 At	2	2	6	2	6	10	2	6	10	14	2	6	10		2	5			
86 Rn	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6			
87 Fr	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6			1
88 Ra	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6			2
89 Ac	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6	1		2
90 Th	2	2	6	2	6	10	2	6	10	14	2	6	10		2	6	2		2
91 Pa	2	2	6	2	6	10	2	6	10	14	2	6	10	2	2	6	1		2
92 U	2	2	6	2	6	10	2	6	10	14	2	6	10	3	2	6	1		2
93 Np	2	2	6	2	6	10	2	6	10	14	2	6	10	4	2	6	1		2
94 Pu	2	2	6	2	6	10	2	6	10	14	2	6	10	6	2	6			2
95 Am	2	2	6	2	6	10	2	6	10	14	2	6	10	7	2	6			2
96 Cm	2	2	6	2	6	10	2	6	10	14	2	6	10	7	2	6	1		2
97 Bk	2	2	6	2	6	10	2	6	10	14	2	6	10	8	2	6	1		2
98 Cf	2	2	6	2	6	10	2	6	10	14	2	6	10	10	2	6			2
99 Es	2	2	6	2	6	10	2	6	10	14	2	6	10	11	2	6			2
100 Fm	2	2	6	2	6	10	2	6	10	14	2	6	10	12	2	6			2
101 Md	2	2	6	2	6	10	2	6	10	14	2	6	10	13	2	6			2
102 No	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6			2
103 Lr	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6	1		2
104	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6	2		2
105	2	2	6	2	6	10	2	6	10	14	2	6	10	14	2	6	3		2

Name ____

<u>**Electron Configurations: Vocabulary**</u>

- **<u>Quantum Mechanics:</u>** A model of the atom which is also called wave mechanics. This model was developed by Heisenberg and Schroedinger. They used theories developed in higher mathematics and probability statistics to explain the structure of the atom. Schroedinger treated the electron mathematically as a wave.
- **<u>Aufbau Principle:</u>** Electrons will occupy orbitals of the lowest energy requirements before going into orbitals that require more energy. This Principle explains why we use the diagonal rule. Electrons first occupy 1s then 2s then 2p etc.
- **<u>Hund's Rule:</u>** An electron will always enter an empty orbital before it pairs up with another electron in the same orbital. (Think about entering a bus. People go to an empty row before sitting next to a person on a seat. You can also think about electrons preferring private rooms provided they don't cost them any more.)
- **<u>Heisenberg Uncertainty Principle:</u>** It is impossible to find the exact location and momentum of an electron within an atom at the same time.
- **<u>Paulis' Exclusion Principle:</u>** No two electrons in the same atom may have the same four quantum numbers. This limits the number of electrons per orbital to two providing they are spinning in opposite directions.

<u>Valence electrons:</u> Electrons in the outermost energy level of an atom. They are responsible for the atom entering into a chemical reaction.

<u>Energy level:</u> This identifies how far the electron is away from the nucleus. *In the apartment analogy, this was the floor of the apartment complex.*

<u>Sublevel:</u> This is an area within an energy level that is composed of orbitals of a specific shape with specific energy requirements. *In the apartment analogy, this was represented as the apartment. Some apartments had only one room, some had three rooms, some had five rooms, and some had seven rooms.*

<u>Orbital:</u> This is a specific area in an electron cloud where the probability of finding an electron is approximately 90%. *In the apartment analogy, this was the room within the apartment.*

<u>Shared pair:</u> When two electrons occupy one orbital.

<u>ıpaired electron:</u> When only one electron is occupying an orbital.

Quantum numbers signify the placement of an electron within the electron cloud. Pauli's exclusion principle states that no two electrons within an atom may have the same set of quantum numbers.

There are four quantum numbers. Think of these numbers like an address of an electron.

The first quantum number is called the PRINCIPAL quantum number.
It is represented as n and is equal to the numbers 1 through 7.
This number tells us the energy level of the electron. (The distance from the nucleus.)

n = 1 through 7

The second quantum number is called the SECONDARY quantum number or the AZIMUTHAL quantum number. We will use the term SECONDARY.
It is represented as l and is equal to a number that represents the shape of the area within the energy level. It is therefore associated with the sublevel.
If the sublevel is an s shape, the quantum number = 0.
If the sublevel is a p shape, the quantum number = 1.
If the sublevel is a d shape, the quantum number = 2.
If the sublevel is an f shape, the quantum number = 3.

Review:
Suppose an electron resides in the second energy level in a p sublevel. So far you have figured out two of the four quantum numbers. $n = 2$ and $l = 1$

The third quantum number is called the MAGNETIC quantum number and represents the electrons placement within the sublevel. It is represented as m. As you recall, each sublevel type consists of varying amounts of orbitals. An s sublevel consists of one orbital, a p sublevel consists of three orbitals, a d sublevel consists of five orbitals, and an f sublevel consists of seven orbitals. Each of these orbitals resides in a particular three dimensional place within the cloud. The magnetic quantum number is equal to $-l$ to $+l$.
In other words, each of the orbitals within the sublevel will be assigned a numerical value.
Example:
An s sublevel consists of only one s orbital. Since l for an s sublevel is 0, m is also 0.
A p sublevel consists of three p orbitals. Since l for a p orbital is 1, the orbitals receive a value of -1, 0 and +1 respectively.
Since l for a d orbital is 2, the five orbitals receive values of -2, -1, 0, +1 and +2.
Since l for an f orbital is 3, the seven orbitals receive values of -3,-2,-1,0, +1, +2, +3.

Review:
Remember our electron in the second energy level? Suppose it resides in the first p orbital. Now you know three of the quantum numbers for this electron. $n = 2$, $l = 1$, $m = -1$

The fourth quantum number is called the SPIN quantum number and represents the direction of the spin. It can either spin, clockwise or counterclockwise. This quantum number is represented with the letter s and can be either:
+1/2(means the electron is spinning clockwise)
-1/2 (means the electron is spinning counter clockwise)

Review:
Let's return to that electron in the second energy level. We now know its address. We know the four quantum numbers. $n = 2$, $l = 1$, $m = -1$, $s = +1/2$

Electromagnetic S[illegible]trum of EMR

Radio Waves	TV	Micro waves	Infrared (IR)	Visible	Ultraviolet (UV)	X rays	Gamma rays	Cosmic Rays
Communications	Communications	Communications Cooking food	Radiant heat		Sunburn Skin cancer	Medical imaging	Radiation	Outerspace
Long wavelength Waves			Sun light			High energy Rays		

Expanded spectrum to show visible colors

Radio Waes	TV	Micro Waves	IR	Visible λ–10^{-7} m : E –10^{-19} J : ν–10^{14} Hz						UV	X rays	Gamma rays
10	1	10^{-2}	10^{-6}	Red R	Orange O	Yellow Y	Green G	Blue B	Violet V	10^{-8}	10^{-9}	10^{-11}

λ (m)	0.1	10^{-4}	7.5×10^{-7}	6.47×10^{-7}	5.85×10^{-7}	5.75×10^{-7}	4.91×10^{-7}	4.24×10^{-7}	4×10^{-7}	10^{-8}	10^{-11}
λ (nm)	10^{8}	10^{5}	750	647	585	575	491	424	400	10	0.01
E (J)	2×10^{-24}	2×10^{-21}	2.7×10^{-19}	3.1×10^{-19}	3.4×10^{-19}	3.5×10^{-19}	4×10^{-19}	4.7×10^{-19}	5×10^{-19}	2×10^{-17}	2×10^{-14}
ν (Hz)	3×10^{9}	3×10^{12}	4×10^{14}	4.6×10^{14}	5.1×10^{14}	5.2×10^{14}	6.1×10^{14}	7.1×10^{14}	7.5×10^{14}	3×10^{16}	3×10^{19}

Variables: E = Energy (J) ν = frequency (Hz or 1/s) λ = wavelength (m)

Constants: c = speed of light in vacuum = 3.00×10^{8} m/s h = Planck's constant = 6.6262×10^{-34} Js (J/Hz)

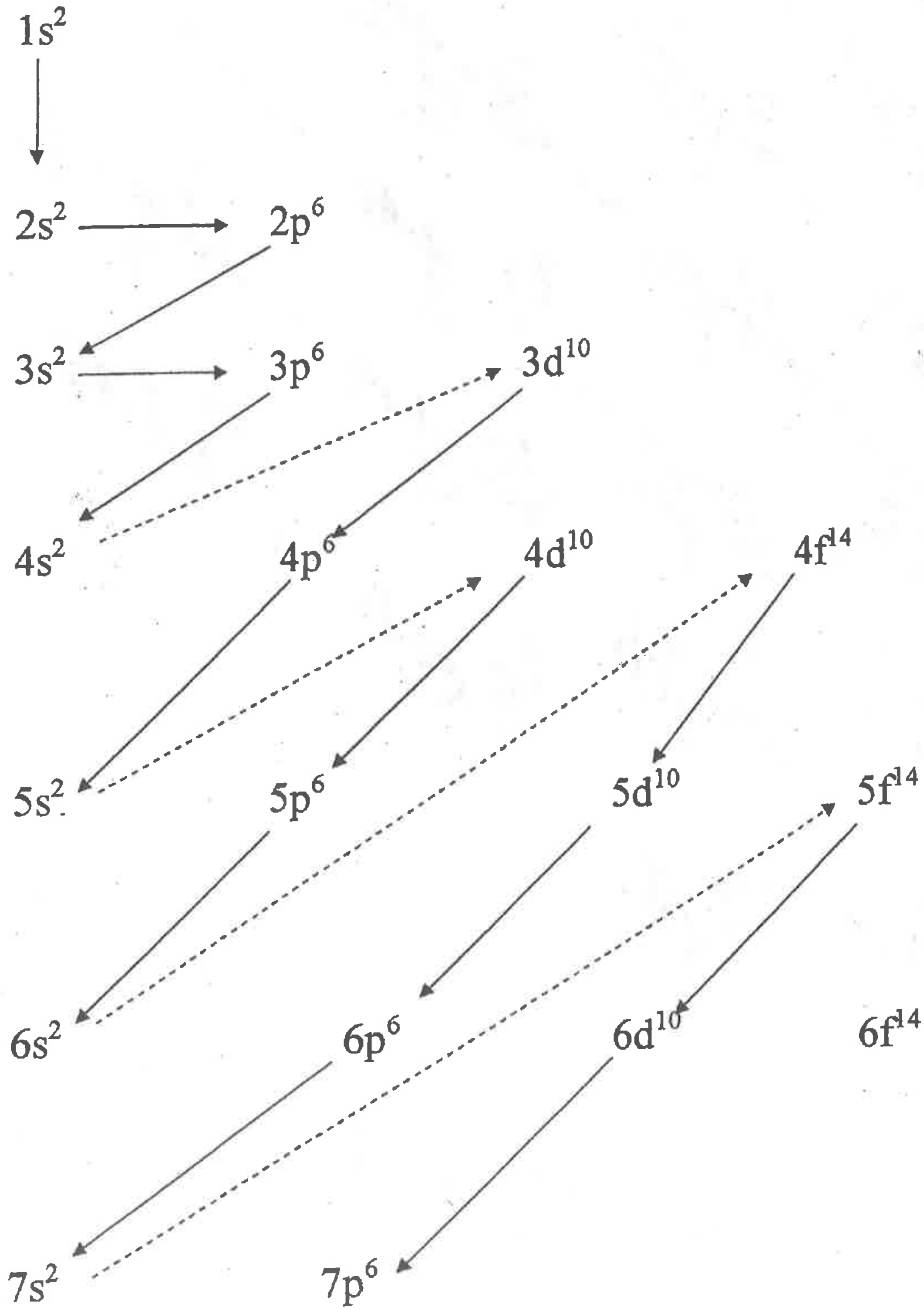
Equations: Energy equals Planck's constant times frequency: $E = \nu h$ Therefore: $\nu = E/h$

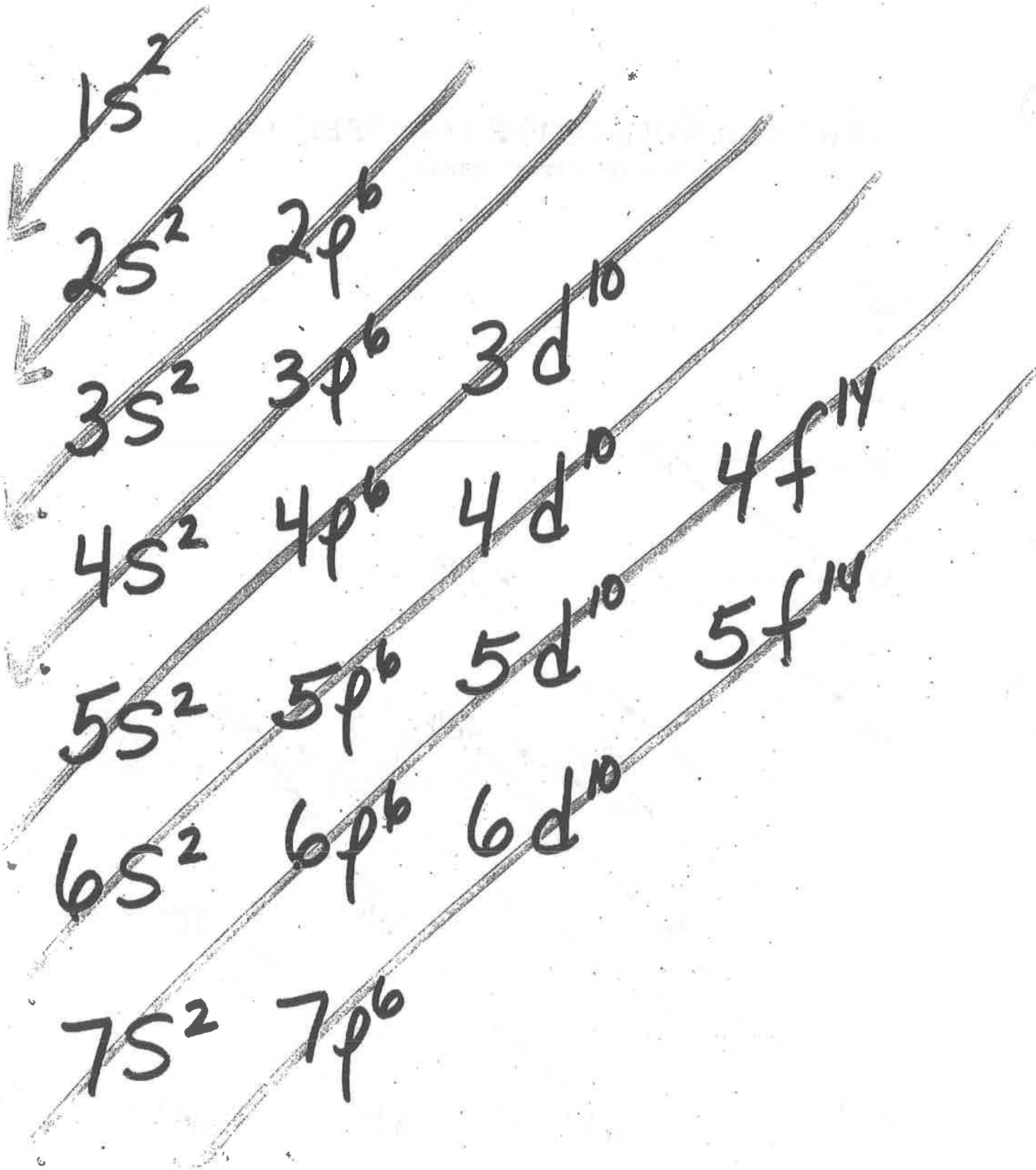
Speed equals frequency times wavelength: $c = \nu\lambda$ Therefore: $\lambda = c/\nu$ $\nu = c/\lambda$

Note: FM radio: 93.1 FM = frequency of 93.1 MHz = 93.1×10^{6} Hz

DIAGONAL RULE FOR FILLING ORBITALS

(there are some exceptions)





Diagonal Rule

Name ______________________ Date ____________ Class __________

5 Practice Problems

1. Write out the electron configurations for (a) potassium and (b) cobalt. How many unpaired electrons does each possess?

2. Which element has the following electron configuration: $1s^22s^22p^3$?

3. Write out the electron configurations for (a) silicon and (b) lithium. How many unpaired electrons does each possess?

4. Which element has the following electron configuration: $1s^22s^22p^63s^23p^3$?

5. Write out the electron configurations for (a) iridium and (b) selenium. How many unpaired electrons does each possess?

6. Which element has the following electron configuration: $1s^22s^22p^63s^23p^64s^23d^1$?

7. Write out the electron configurations for (a) bismuth and (b) vanadium. How many unpaired electrons does each possess?

8. Which element has the following electron configuration: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}$?

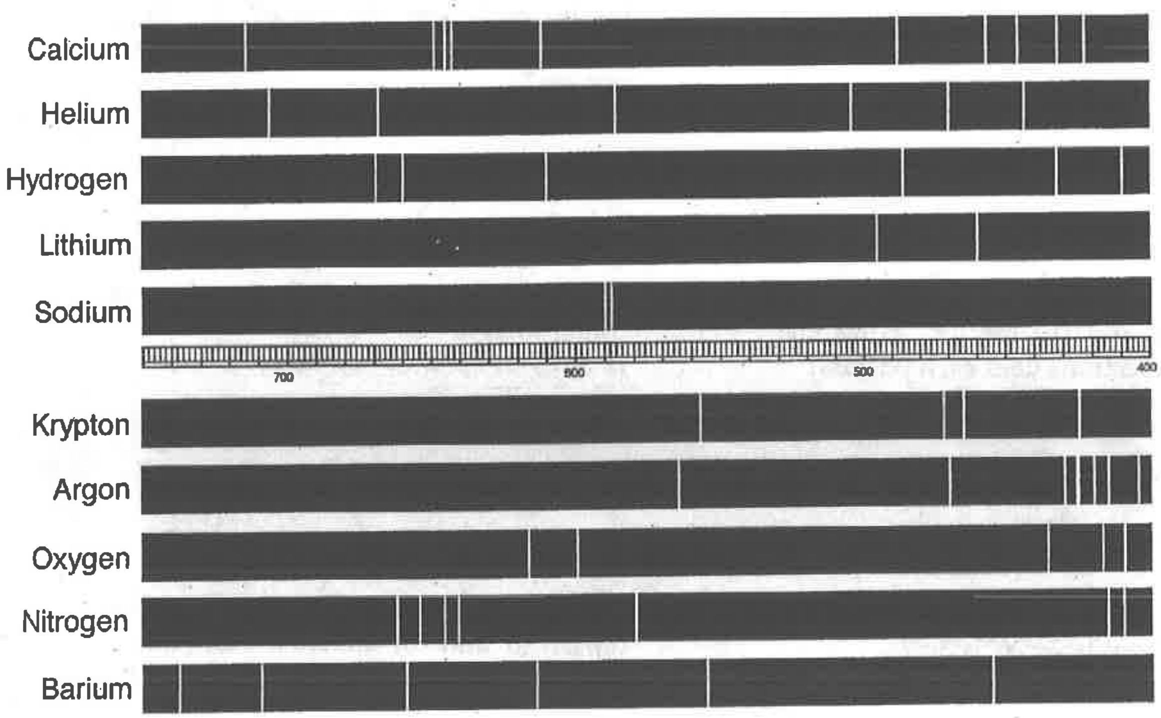
9. Write out the electron configurations for (a) sulfur and (b) mercury. How many unpaired electrons does each possess?

10. Which element has the following electron configuration: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^6$?

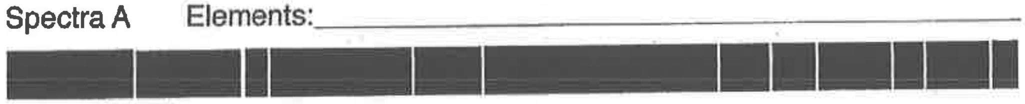
Mystery Spectra

Name:_______________

Identify the elements in spectra A, B, and C by comparing the bright lines present with the bright lines in the spectra for known elements.

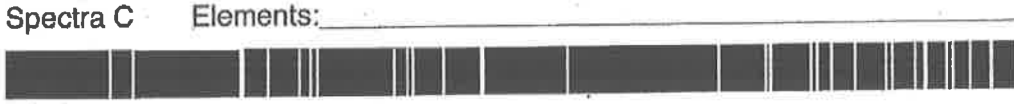


Spectra A Elements:_______________________________________



Spectra B Elements:_______________________________________

Spectra C Elements:_______________________________________



ELECTRON CONFIGURATION WORKSHEET

PRINCIPAL ENERGY LEVEL (N)

N=1,2,3,4 etc.

SUBLEVELS

NAME ______ ______ ______ ______

ORBITALS ______ ______ ______ ______

EACH ORBITAL CAN HOLD ____ ELECTRONS

OF ELECTRONS
ALTOGETHER ______ ______ ______ ______

PRINCIPAL ENERGY LEVEL	SUBLEVELS	TOTAL # OF ELECTRONS = $2N^2$
1		
2		
3		
4		

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

ARRANGE THE ELECTRONS FROM 1s TO 5s IN ORDER OF THEIR INCREASING ENERGY. SHOW THE MAXIMUM NUMBER OF ELECTRONS WHICH CAN BE HELD IN EACH SUBLEVEL.

$1s^2$ ____ ____ ____ ____ ____ ____ ____ ____

Name ______________________________ Date ____________________

Write the orbital diagram first AND THEN write the electron configuration of the following elements.

1. Calcium #20 ______________________________ orbital diagram

______________________________ electron configuration

2. Lithium #3 ______________________________ orbital diagram

______________________________ electron configuration

3. Argon #18 ______________________________ orbital diagram

______________________________ electron configuration

4. Iron #26 ______________________________ orbital diagram

______________________________ electron configuration

5. Sodium #11 ______________________________ orbital diagram

______________________________ electron configuration

6. Iodine # 53 ______________________________ orbital diagram

______________________________ electron configuration

How many unpaired electrons does Iodine have? ______________

How many pairs of electrons does sodium have? ______________

ORBITAL DIAGRAM, ELECTRON CONFIGURATION

ELEMENT	ATOMIC NUMBER	ORBITAL DIAGRAM ELECTRON CONFIGURATION	APPLICATION OF FILLING RULES
hydrogen	1	1s: [↑] $1s^1$	first electron found in lowest energy level
helium	2	1s: [↑↓] $1s^2$	second electron occupies remaining vacancy in lowest available energy level with opposite spin
lithium	3	1s: [↑↓] 2s: [↑] $1s^2$ $2s^1$	first energy level is now full. The third electron occupies the s sublevel in the second energy level the lowest available
beryllium	4	1s: [ ] 2s: [ ] $1s^2$ $2s^2$	fourth electron occupies remaining vacancy in lowest available energy level with opposite spin
boron	5	1s: [↑↓] 2s: [↑↓] 2p: [↑][ ][ ] $1s^2$ $2s^2$ $2p^1$	fifth electron occupies first orbital in the p sublevel
carbon	6	1s: [↑↓] 2s: [↑↓] 2p: [↑][↑][ ] $1s^2$ $2s^2$ $2p^2$	sixth electron occupies second orbital in the p sublevel
nitrogen	7	1s: [ ] 2s: [ ] 2p: [ ][ ][ ] $1s^2$ $2s^2$ $2p^3$	seventh electron occupies third orbital in the p sublevel
oxygen	8	1s: [↑↓] 2s: [↑↓] 2p: [↑↓][↑][↑] $1s^2$ $2s^2$ $2p^4$	eighth electron occupies the empty space in the first orbital in the p sublevel with opposite spin
fluorine	9	1s: [ ] 2s: [ ] 2p: [ ][ ][ ] $1s^2$ $2s^2$ $2p^5$	ninth electron occupies the empty space in the second orbital in the p sublevel with opposite spin
neon	10	1s: [↑↓] 2s: [↑↓] 2p: [↑↓][↑↓][↑↓] $1s^2$ $2s^2$ $2p^6$	tenth electron occupies the empty space in the third orbital in the p sublevel with opposite spin
sodium	11	1s: [ ] 2s: [ ] 2p: [ ][ ][ ] 3s: [ ] $1s^2$ $2s^2$ $2p^6$ $3s^1$	second energy level is now full. The eleventh electron occupies the s sublevel in the third energy

QUESTIONS :

1) WHICH PRINCIPAL ENERGY LEVEL CAN HOLD A MAXIMUM OF 18 ELECTRONS ?

2) IF N REPRESENTS THE PRINCIPAL ENERGY LEVEL, THE MAXIMUM NUMBER OF ELECTRONS POSSIBLE IN THAT PRINCIPAL ENERGY LEVEL IS EQUAL TO :

A) N B) 2N C) N^2 D) $2N^2$

3) WHAT IS THE MAXIMUM NUMBER OF SUBLEVELS IN THE THIRD PRINCIPAL ENERGY LEVEL ?

4) THE TOTAL NUMBER OF **d** ORBITALS IN THE THIRD PRINCIPAL ENERGY LEVEL IS :

5) WHICH ATOM IN THE GROUND STATE CONTAINS ONLY ONE ORBITAL THAT IS PARTIALLY OCCUPIED ?

A) Si B) Ne C) Ca D) Na

6) THE TOTAL NUMBER OF COMPLETELY FILLED ORBITALS IN AN ATOM OF NITROGEN IN THE GROUND STATE IS :

A) 1 B) 2 C) 3 D) 4

7) THE TOTAL NUMBER OF ORBITALS IN AN **f** SUBLEVEL IS :

A) 1 B) 3 C) 5 D) 7

8) WHAT IS THE ELECTRON CONFIGURATION FOR $_4Be^{2+}$?

A) $1S^1$ B) $1S^2$ C) $1S^2\ 2S^1$ D) $1S^2\ 2S^2$