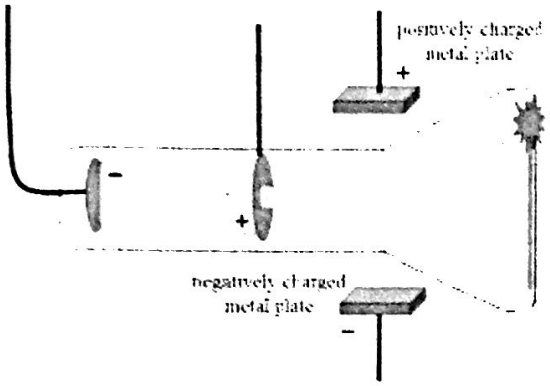
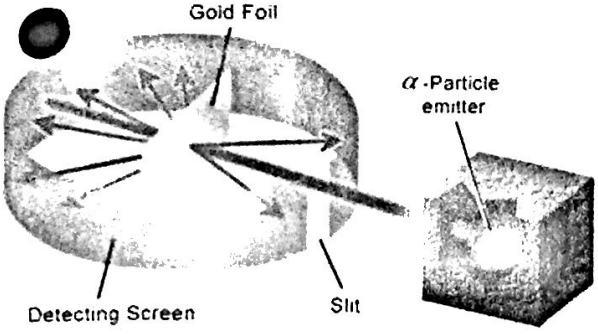
Station 1: Early Atomic Theory Scientists

The development of atomic theory took place over a long period of time through the work of many scientists.

Examine the diagrams below, depicting two major experiments that contributed to the development of the atomic theory. Complete the table using your knowledge of Chemistry.



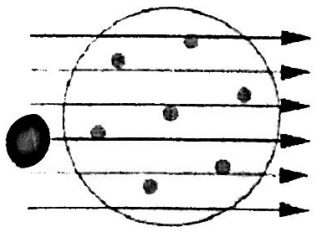
Name of Experiment	Cathode Ray Tube Exp.
Analyzed By (Scientist's Name)	JJ Thomson
Observation(s) Made	Cathode ray bends towards + plate, and away from - plate
Conclusion(s)	Cathode rays are composed of a stream of - particles, called electrons



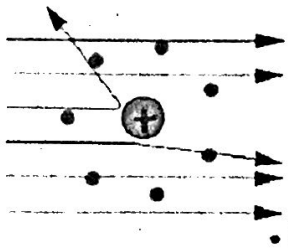
Name of Experiment	Gold Foil Experiment
Analyzed By (Scientist's Name)	Ernest Rutherford
Observation(s) Made	1) most alpha particles passed through foil undeflected 2) some alpha particles were deflected from foil
Conclusion(s)	1) Most of atom is empty space, which is why most alpha particles passed through undeflected 2) Each atom must contain a small, dense, + charged nucleus, the + α particles that hit this were repelled/deflected due to the same charge.

The diagrams below represent the models proposed by both Thomson and Rutherford. Label the model with the appropriate scientist. Then explain how the two models differ and the experimental evidence that supported the revision to atomic theory.

Thomson



Rutherford



- Thomson - Plum Pudding Model, (-) electrons embedded evenly throughout a (+) substance
  - Why? cathode ray experiment discovered (-) electrons
- Rutherford - nuclear model w/ (+) charged nucleus surrounded by (-) electrons (see experiment above)

Station 2: Quantum Mechanical Model of the Atom

The modern model of the atom (also referred to as the wave mechanical model) uses theoretical quantum physics and mathematical probability to predict the regions within the atom with the most probable location of an electron.

Identify the scientist that is responsible for each of the following contributions to the modern atomic theory.

1. It is impossible to predict the exact position and momentum of an electron at any time. Heisenberg
2. Created a wave-function to predict electron probability. Schrödinger
3. First to propose that electrons can behave as both particles and waves. deBroglie

Using your knowledge of the modern atomic theory answer the following questions.

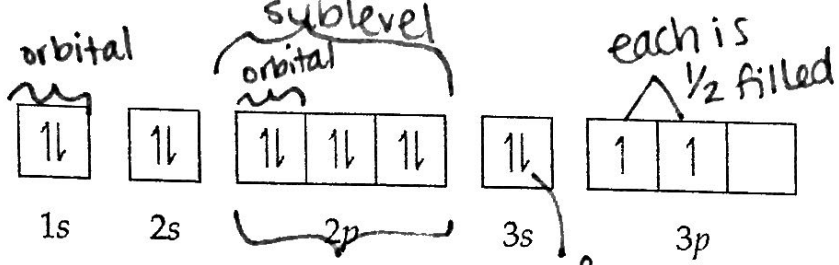
4. Circle the sublevels that do not exist: 4p (1p) (2f) 5s 3d 7p (2d) 3s
5. The number of orbitals in a d sublevel. 5
6. The number of orbitals in n = 2 shell. 4
   S sublevel = 1 orbital, P sublevel = 3 orb.
7. The number of sublevels in n = 5. 5
   # sublevels = n
8. The max. number of electrons in n=3. 18 e-
   $2n^2 = 2(3^2)$
9. The max. number of electrons in a p sublevel. 6 e-
   3 orbitals x 2 e- per orbital
10. The number of orbitals in l = 1. 3
    l=1 is P sublevel, which has 3 orbitals
11. The sublevels present in n = 2. 2
    # sublevels = n
12. The max. number of electron in l = 0. 2 e-
    l=0 is S sublevel, which only has 1 orbital. Each orbital can hold 2e-.

Questions 13-15 refer to answer choices a through d. Identify all of the answer choices that apply to the given question.

a. 5s and 6s orbitals b. 3p and 4p orbitals c. 3s and 4s sublevels d. 2s and 2p orbitals

13. They have the same shape. a, b, c
14. The maximum number of electrons is the same. a, b, c
15. They are in the same energy level. d

Station 3: Electron Configurations



1. Given the Orbital Fill Diagram:
   a. Identify the element. Si
      14e = 14p = Si
   b. How many orbitals are completed filled? 6
      each box = orbital
   c. How many orbitals are half filled? 2
   d. How many sublevels are completely filled? 4
      (1s, 2s, 2p, 3s)

✱complete/full sublevel = 2e- in each box of every orbital in that sublevel.

full orbital = 2 e- in box

2. Complete the chart below using your knowledge of chemistry and the NYS Reference Tables.

Element	Aluminum, Al (Z=13)
Complete electron configuration	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
shell configuration	2 - 8 - 3
Number of valence electrons	3 $(3s^{②} + 3p^{①})$

↳ all the e- in the outermost/highest principal energy level

Element	Nickel, Ni (Z=28) ✱Atomic # over 18, so use diagonal rule!
Complete electron configuration	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^8$
shell configuration	2 - 8 - 16 - 2
Number of valence electrons	2 $(4s^{②})$

Station 4: Ground State vs. Excited State

1. Using your knowledge of chemistry, define the terms below.

Ground State	when electrons occupy the lowest possible principal energy levels + sublevels
Excited State	when electrons occupy higher energy levels after absorbing energy.

2. Explain, in terms of energy and electrons, the production of the bright line spectrum of an element.

The BLS of an element is produced when electrons fall from the higher-energy excited state to the lower-energy ground state, emitting energy in the form of light.

3. Assume each of the following represents a neutral atom. Determine the name of the element and if the configuration is written in the ground state or excited state.

Electron Configuration	Element Name - find total # e-	Ground vs. Excited
$1s^2 2s^2 2p^2 3s^1$ (2p not completely full, yet 3s is written, so excited)	nitrogen	excited
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	calcium	ground
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9 5s^1$ (3d not full)	zinc	excited

*Remember, s sublevel can hold 2e-, p sublevel can hold 6e-, d sublevel can hold 10e-, f sublevel can hold 14e-

4. Write the ***ground state electron configuration*** for a neutral atom of calcium.

$_{20}Ca: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

5. Write a possible ***excited state electron configuration*** for a neutral atom of calcium.

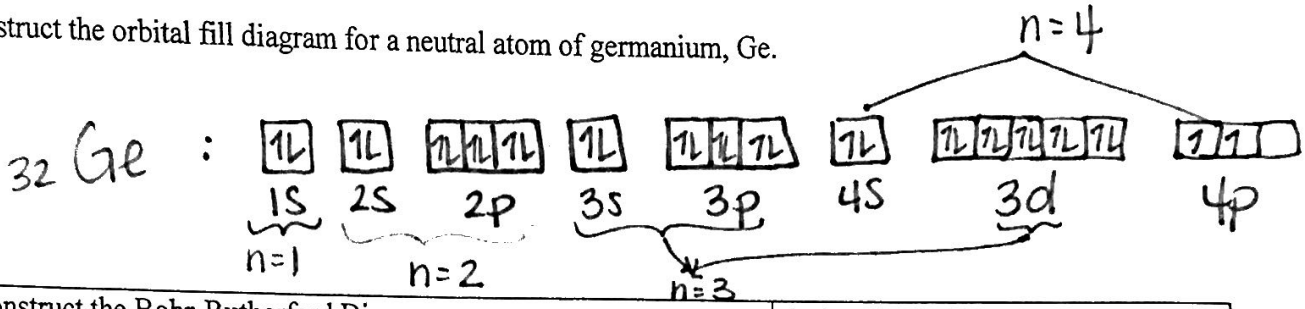
*Many possible. To check your work, write the shell config. It should have e- in higher energy levels than 2-8-8-2

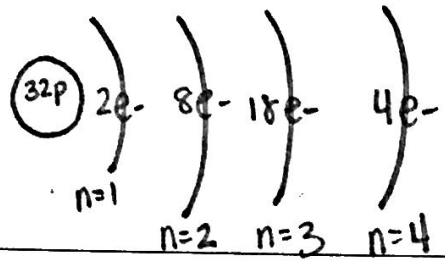
$1s^2 2s^2 2p^6 3s^2 3p^5 4s^2 4p^1$

check w/ shell config → 2 8 7 3

Station 5: Representing Atoms

1. Construct the orbital fill diagram for a neutral atom of germanium, Ge.



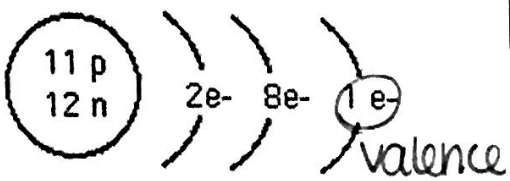
2. Construct the Bohr-Rutherford Diagram	3. Construct the Lewis Dot Diagram (only valence)
	

4. How is a Lewis Dot Diagram different from a Bohr-Rutherford Diagram?

The Lewis Dot diagram shows only the valence shell electrons, while the Bohr-Rutherford Diagram show the # of e- in each principal energy level.

Questions 5 through 8 refer to the Bohr-Rutherford Diagram provided.

5. What is the symbol for the element? Na
6. What is the nuclear charge of the atom? +11 (↳ look @ # of p)
7. What is the mass number for the atom? 23 (↳ = + # protons; #p + #n)
8. How many valence electrons? 1



9. Complete the chart for the elements represented by the dot diagrams below.

Dot Diagram	Number of Valence Electrons	Full Orbitals in valence shell	Half-full orbitals in valence shell	Empty orbitals in valence shell
:Cl· $1s^2 2s^2 2p^6 3s^2 3p^5$	7	$3s^2$ $3p^5$ 3s 3p 3	3p 1	0
:Ne: $1s^2 2s^2 2p^6$	8	2s 2p all full 4	0 (all full)	0
B· $1s^2 2s^2 2p^1$	3	2s 2p 1	2s 2p 1	2s 2p 2

Station 6: Quantum Numbers

1. State the four quantum numbers and explain what each represents.

n	principal energy level, gives info about size of e- cloud
l	sublevel, gives info about shape of e- cloud
m	orbital, gives info about orientation in space
s	spin, gives info about direction of electron spin

2. Identify the n and l values for the following orbitals.

n → 1s ← l=0 is s, l=1 is p, l=2 is d, l=3 is f

1s: n = 1, l = 0

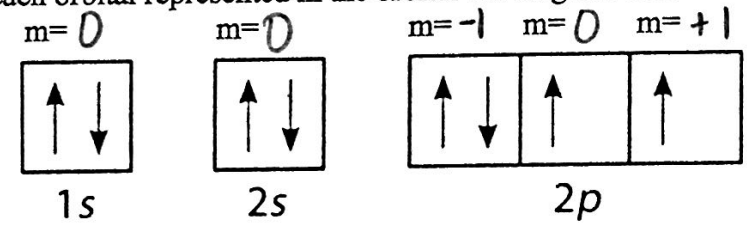
3s: n = 3, l = 0

2p: n = 2, l = 1

4d: n = 4, l = 2

5f: n = 5, l = 3

3. Label the m value for each orbital represented in the orbital fill diagram below.



4. State the number of possible electrons described by the quantum numbers below.

a. n = 3, l = 0 → 3s; 2e-

b. n = 3, l = 1 → 3p; 6e-

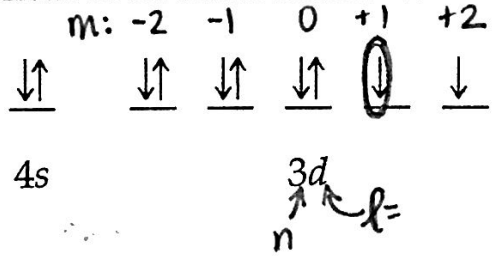
↳ 3 p orbitals x 2e- each

c. n = 3, l = 2, m = -1 → 2e-

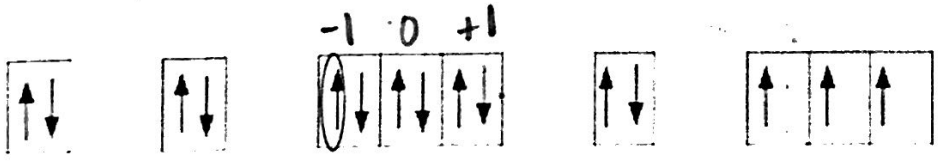
d. n = 5, l = 0, m = -2 → 2e-

5. Identify the four quantum numbers for the circled electron in each of the following diagrams.

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



n = 3
l = 2
m = +1
s = -1/2



n = 2
l = 1
m = -1
s = +1/2

2p: n ↑ l=1

6. Explain why each of the following set of quantum numbers is *not* possible.

a. n=2, l = 2, m = -2, s= +½

↳ l=2 would be the d sublevel, which does not exist for the second PEL

b. n = 3, l = 0, m = 1, s = -½

↳ The l=0 is the s sublevel. There is only 1 possible orbital of m=0.

m ranges from -l to +l.

Name: _____

HONORS CHEMISTRY – Modern Atomic Theory Review Questions

DIRECTIONS: Answer each of the following questions using your knowledge of chemistry and the periodic table. Be sure to include any work associated with each problem and underline/circle key words that will help in remembering the PROCESS of answering questions.

1. An orbital is a region of space where there is a high probability of finding
   (A) A proton
   (B) A positron
   (C) A neutron
   (D) An electron

2. The maximum number of electrons in a single orbital of the 3d sublevel may contain is
   (A) 5 (B) 2 (C) 3 (D) 4

3. What is the total number of orbitals in the f sublevel?
   (A) 1 (B) 7 (C) 3 (D) 5

4. Which electron transition would result in the emission of energy?
   (A) 3s to 4s
   (B) 3p to 4p
   (C) 3s to 3p
   (D) 4p to 3s

   ↳ from higher energy to lower energy

5. What is the total number of orbitals containing only one electron in an atom of nitrogen in the ground state?
   (A) 1
   (B) 2
   (C) 3
   (D) 4

   $_7N$: 1L (1s) 1L (2s) 1 1 1 (2p)

6. An atom in the excited state can have an electron configuration of
   (A) $1s^2 2s^2$
   (B) $1s^2 2s^2 2p^5$
   (C) $1s^2 2p^1$
   (D) $1s^2 2s^2 2p^6$

   ↳ e- in higher energy levels (no correct answer)

7. Which sublevels are occupied in the outermost principal energy level of an argon atom in the ground state?
   (A) 3s and 3d
   (B) 3s and 3p
   (C) 2s and 3p
   (D) 2p and 3d

   $_{18}Ar: 1s^2 2s^2 2p^6 \underbrace{3s^2 3p^6}_{\text{valence e-}}$

skip for now

8. Which particle has an electron configuration $1s^2 2s^2 2p^6$?
   (A) Cl
   (B) Cl^-
   (C) Na
   (D) Na^+

   10 e- total

9. What is the total number of unpaired electron in an atom of nickel in the ground state?
   (A) 0
   (B) 2
   (C) 3
   (D) 4

   $_{28}Ni: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

   1L 1L 1L 1 1 (3d)

[1][1][1]
3p

10. Which atom in the ground state has three half filled orbitals?
(A) P $1s^2 2s^2 2p^6 3s^2 3p^3$
(B) Si $1s^2 2s^2 2p^6 3s^2 3p^2$
(C) Al $1s^2 2s^2 2p^6 3s^2 3p^1$
(D) Li $1s^2 2s^1$

11. How many electrons can exist in the quantum state: 2, 1, -1, $\frac{1}{2}$? } Pauli: each e- has a unique set of 4 quantum #s
(A) 0 (B) 1 (C) 2 (D) 6 (E) 10

12. What is the nuclear charge of an atom with an atomic number of 6 and a mass number of 14?
(A) +6 (B) 0 (C) +14 (D) +8

13. An electron cannot exist in the energy state described by which of the following set of quantum numbers? n, l, m, S
(A) 1, 0, 0, $\frac{1}{2}$ ✓
(B) 3, 2, 2, $-\frac{1}{2}$ ✓
(C) 4, 3, 3, $\frac{1}{2}$ ✓
(D) 2, 1, -3, $\frac{1}{2}$ (l, m) m can only range from -l to +l
(E) 2, 0, 0, $-\frac{1}{2}$

14. Which is the shape of a p-orbital?
(A) Spherical
(B) Pointed
(C) Propeller } either choice.
(D) Dumbbell

15. Which of the following scientists discovered that atoms contain electrons?
(A) Niels Bohr
(B) Ernest Rutherford
(C) Wolfgang Pauli
(D) Erwin Schrodinger
(E) J.J. Thomson – cathode ray tube experiment

16. Which of the following is a conclusion of the gold foil experiment?
(A) An atom has a uniform positive density.
(B) Most of the atom is empty space.
(C) The nucleus of an atom is neutral.
(D) Atoms contain electrons.

17. According to Heisenberg's uncertainty principle at any given time, if one knows the exact position of a particle, it is impossible to determine the particle's exact
(A) Charge
(B) Mass
(C) Density
(D) Size
(E) Momentum