ATOMIC STRUCTURE – TIME LINE

B.C. – Ancient Greeks: Democritus and Leucippus

Vs.
Continuous theory of Matter

Evidence in support of the atomic theory did not come until almost 2000 years later.

1770—Lavoisier

1799—Proust

1803—Dalton

1803—Dalton's Atomic Theory

1.

2.

3.

4.

MODERN ATOMIC THEORY

Dalton's theory had existed for almost 100 years. As new discoveries were made Dalton's theory had to be revised to include the structure of the atom. The major differences from Dalton include:

1.

2.

1897—J.J. Thomson

1909—Robert A. Millikan

1911—Rutherford

1913—Bohr Model

1920—Schroedinger Wave Model

1932—James Chadwick

Name _____ Class _____ Date _____

REVIEW and REINFORCEMENT

An Atomic Model of Matter

Section **5-1**

KEY CONCEPTS

- ▲ After much observation and questioning, Democritus concluded that matter could not be divided into smaller and smaller pieces forever. Eventually the smallest possible piece would be obtained.
- ▲ All elements are composed of atoms. Atoms are indivisible and indestructible particles.
- ▲ Atoms of the same element are exactly alike.
- ▲ Atoms of different elements are different.
- ▲ Compounds are formed by the joining of atoms of two or more elements.
- ▲ According to Thomson's atomic model, the atom was made of a puddinglike positively charged material throughout which negatively charged electrons were scattered, like plums in a pudding.
- ▲ Rutherford reasoned that all of an atom's positively charged particles were contained in the nucleus. The negatively charged electrons were scattered outside the nucleus around the atom's edge.
- ▲ According to Bohr's atomic model, electrons move in definite orbits around the nucleus, much like planets circle the sun. These orbits, or energy levels, are located at certain distances from the nucleus.
- ▲ According to the modern atomic model, an atom has a small positively charged nucleus surrounded by a large region in which there are enough electrons to make the atom neutral.

■ Vocabulary Skills: Expanding Definitions

Write five facts about the term **nucleus**, using the following words as cues.

1. Size _____
2. Location _____
3. Rutherford _____
4. Charge _____
5. Density _____

Models of the Atom: Understanding the Main Ideas

Decide which model of the atom each of the following sentences describes. Then fill in the blank before each sentence according to the following key:

DM = Democritus

DL = Dalton

R = Rutherford

T = Thomson

B = Bohr

W = Wave model

If a sentence seems to describe more than one atomic model, choose the model that *first* pictured the atom this way.

_____ 1. Atoms are small, hard particles.

_____ 2. An atom contains negatively charged particles called "corpuscles."

_____ 3. Atoms of the same element are exactly alike.

_____ 4. In an atom, electrons move in definite orbits around the nucleus, much like planets circle the sun.

_____ 5. An atom is the smallest piece of matter.

_____ 6. An atom is mostly empty space with a dense, positively charged nucleus in the center.

_____ 7. Atoms are indivisible.

_____ 8. An atom has a small, positively charged nucleus surrounded by a large region in which scientists can predict where an electron is likely to be found.

_____ 9. An atom is made of positively charged, puddinglike material through which negatively charged particles are scattered.

_____ 10. In an atom, electrons are located in energy levels that are a certain distance from the nucleus.

SUBATOMIC PARTICLES

1. Subatomic Particles:

Particle	Symbol(s)	Charge	Mass (amu)	Location

2. Atomic Number (Z):

3. Mass Number (A):

Try the following:

Sample atom	Notation	Protons	Neutrons	Electrons
Carbon-12 C-12				
Chlorine-35 Cl-35				
Chlorine-37 Cl-37				
Neon-20 Ne-20				
Oxygen-16 O-16				

4. Isotopes:

5. Atomic mass:

Try the following:

1. Determine the number of protons, neutrons and electrons for the following isotopes of hydrogen:
   hydrogen-1 (protium)

   hydrogen-2(deuterium)

   hydrogen-3 (tritium)

2. Naturally occurring chlorine consists of 75 % chlorine –35 and 25% chlorine-37. Find the average atomic mass.

3. Calculate the atomic mass of an element with isotope A occurring 70.0% of the time with a mass of 13.0 amu and isotope B occurring 30.0% with a mass of 15.0 amu.

4. An element X has three isotopes x-30 has a 50.0% abundance, x-28 has a 30.0% abundance and x-31 has a 20.0% abundance. Calculate the average atomic mass.

5. There are two isotopes of element z, 60.0% of the atoms have a mass of 58.0 amu and 40.0% have a mass of 57.0 amu. Calculate the atomic mass of element Z.

Name _____ Class _____ Date _____

REVIEW and REINFORCEMENT

Se 5

Structure of the Atom

KEY CONCEPTS

▲ The three main subatomic particles are the proton, the neutron, and the electron.

Vocabulary Skills: Understanding Relationships

Explain how the following terms are related.

1. isotope: neutron

_____

2. quark: subatomic particle

_____

_____

3. atomic number: proton

_____

4. subatomic particles: atomic mass unit

_____

5. electron: energy level

_____

_____

6. mass number: atomic mass

_____

_____

_____

7. electron: electron cloud

_____

_____

8. isotope: mass number

_____

Subatomic Particles: Reviewing the Main Ideas

Complete the following chart:

Particle	Location	Mass (amu)	Charge
Proton			
Electron			
Neutron			

Find the Missing Numbers

Use your knowledge of atomic number and mass number to fill in the missing numbers:

Element	Atomic #	Mass #	HOW MANY?		
			Protons	Neutrons	Electrons
Iron	26	56			
Sulfur	16	32			
Carbon	6			6	
Fluorine		19	9		
Calcium	20	40			
Nitrogen		14			7
Copper	29			35	
Sodium		23	11		
Mercury		201			80
Silver				61	47

ATOMIC STRUCTURE

Name _____

An atom is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud). The atomic number is equal to the number of protons. The mass number is equal to the number of protons plus neutrons. In a neutral atom, the number of protons equals the number of electrons. The charge on an ion indicates an imbalance between protons and electrons. Too many electrons produces a negative charge, too few, a positive charge.

This structure can be written as part of a chemical symbol.

Example:

mass number → ^{15}; *charge* → $^{+3}$; *atomic number* → $_{7}$

$^{15}_{7}N^{+3}$

7 protons

8 neutrons (15 – 7)

4 electrons

Complete the following chart.

Element/ Ion	Atomic Number	Atomic Mass	Mass Number	Protons	Neutrons	Electrons
H						
H^+						
$^{12}_{6}C$						
$^{7}_{3}Li^+$						
$^{35}_{17}Cl^-$						
$^{39}_{19}K$						
$^{24}_{12}Mg^{2+}$						
As^{3-}						
Ag						
Ag^{+1}						
S^{-2}						
U						

Isotopes or Different Elements?

In each of the following statements, you are given a pair of elements and important information about each. Use this information to determine if the pair of elements are isotopes or different elements. Indicate your answer in the space provided.

1. Element D has 6 protons and 7 neutrons.
   Element F has 7 protons and 7 neutrons. _____

2. Element J has 27 protons and 32 neutrons.
   Element L has 27 protons and 33 neutrons. _____

3. Element X has 17 protons and 18 neutrons.
   Element Y has 18 protons and 17 neutrons. _____

4. Element Q has 56 protons and 81 neutrons.
   Element R has 56 protons and 82 neutrons. _____

5. Element T has an atomic number of 20 and an atomic mass of 40.
   Element Z has an atomic number of 20 and an atomic mass of 41. _____

6. Element W has 8 protons and 8 neutrons.
   Element V has 7 protons and 8 neutrons. _____

7. Element P has an atomic number of 92 and an atomic mass of 238.
   Element S has 92 protons and 143 neutrons. _____

ISOTOPES AND AVERAGE ATOMIC MASS

Name _____

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

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

Example: A sample of cesium is 75% ^{133}Cs, 20% ^{132}Cs and 5% ^{134}Cs. What is its average atomic mass?

Answer: .75 x 133 = 99.75
.20 x 132 = 26.4
.05 x 134 = 6.7
Total = 132.85 amu = average atomic mass

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

1. 80% ^{127}I, 17% ^{126}I, 3% ^{128}I _____

2. 50% ^{197}Au, 50% ^{198}Au _____

3. 15% ^{55}Fe, 85% ^{56}Fe _____

4. 99% ^{1}H, 0.8% ^{2}H, 0.2% ^{3}H _____

5. 95% ^{14}N, 3% ^{15}N, 2% ^{16}N _____

6. 98% ^{12}C, 2% ^{14}C _____

THE BOHR MODEL OF THE ATOM

Niels Bohr, the Danish physicist, created an atomic model that explained the answer to the following questions:

1. Why do not atoms not collapse?
2. How does an excited atom radiate energy?

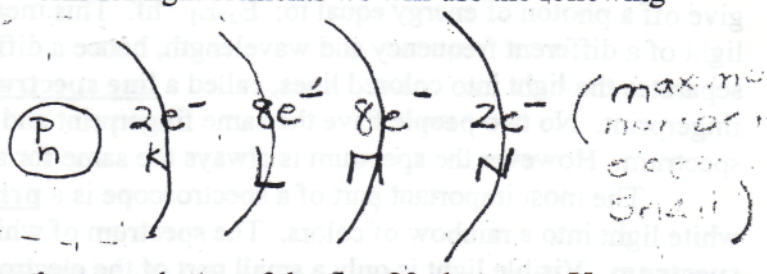
According to Bohr's Atomic Model, electrons move in definite orbits around the nucleus, much like planets circle the sun. These orbits or energy levels are located at certain distances from the nucleus. Bohr did work with hydrogen and other light atoms.

These concentric orbits or shells are called principal energy levels and can be designated by the letters: K, L, M, N, O, P, Q or by the numbers 1, 2, 3, 4, 5, 6, 7,

Electrons in energy levels closer to the nucleus are in a lower energy state than those electrons found in energy levels farther from the nucleus. When electrons are in the lowest available energy level they are said to be in the **ground state.**

If these electrons absorb energy and shift to higher energy levels the atom is said to be in the **excited state.** The excited state of the atom is unstable and the electron or electrons fall back to lower energy levels while simultaneously radiating energy.

Since the Bohr Model only works for light elements we will use the following guidelines:



Activity: Draw the Bohr model for the ground state of the first 20 elements. Use hydrogen as a guideline and complete the table below:

Name of the Element	Symbol	Bohr Model
Hydrogen	$^{1}H_{1}$	

If these electrons absorb energy and shift to higher energy levels the atom is said to be in the **excited state.** The excited state of the atom is unstable and the electrons fall back to lower energy levels while simultaneously radiating energy, which is equal to the difference in energy before and after the shift in energy levels.

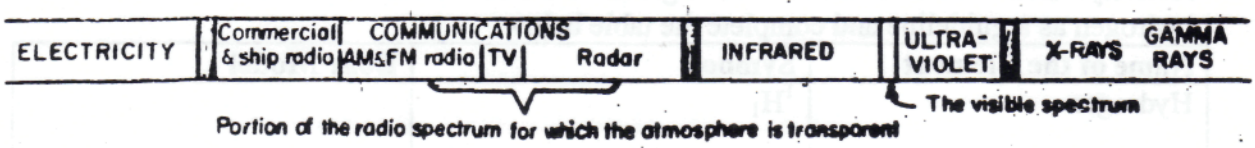
Bohr assumed that the gain or lose of energy, or the absorption or emission of energy was not continuous but rather emitted in discrete or separate bursts of energy called "quanta". This means that the shift of an electron is directly proportional to the frequency of light. (Number of waves per second). Since frequency is inversely proportional to wavelength, the energy of the photon is indirectly proportional to the wavelength.

Use the following symbols to represent the qualities above:

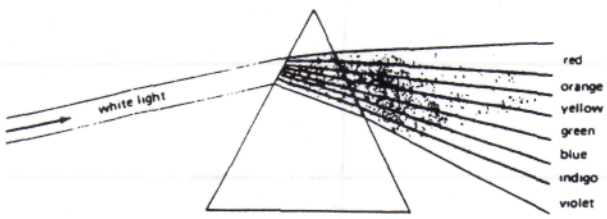
V= velocity
F= frequency
λ= wavelength
E_2= energy of the atom when the electrons are in a higher energy state
E_1= energy of the atom when the electrons are in a lower energy state
h= Planck's constant

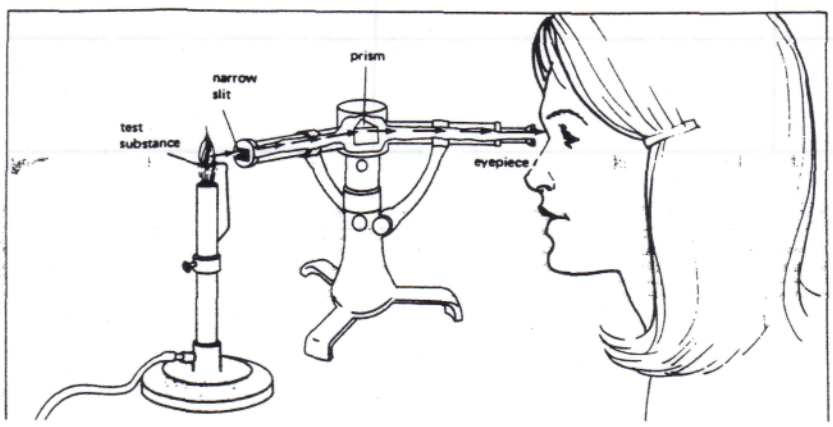
When an electron goes from a higher energy state to a lower energy state it will give off a photon of energy equal to: $E_2-E_1= hf$. This means that each transition gives off light of a different frequency and wavelength, hence a different color. A spectroscope separates the light into colored lines, called a **line spectrum.** A spectrum is like a fingerprint. No two people have the same fingerprint and no two elements have the same spectrum. However the spectrum is always the same for a particular element.

The most important part of a spectroscope is a **prism.** A prism can break up white light into a rainbow of colors. The spectrum of white light is called a **continuous spectrum.** Visible light is only a small part of the electromagnetic spectrum.

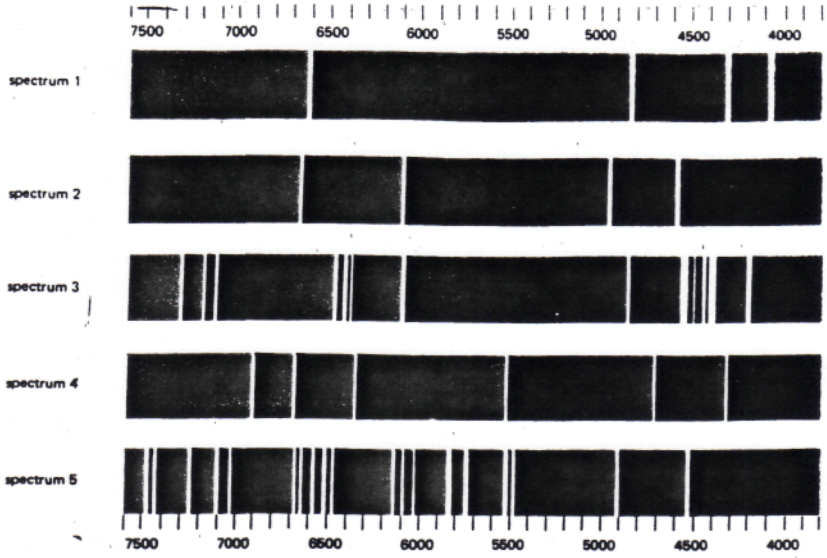


In a spectroscope a prism forms spectrums for individual elements.





Below are five unidentified elements. Use the spectrums of known elements from the previous page to identify the five spectrums below. Write your answers in the table:



SPECTRUM	ELEMENT
1	
2	
3	
4	
5	

MODERN ORBITAL MODEL

The Bohr model of the atom was capable of explaining the spectra of simple atoms like hydrogen. However it could not explain fully the spectra of heavier or more complicated atoms that contain additional spectral lines. These additional spectral lines are more fully explained by the modern orbital model of the atom.

In the Modern Atomic Model (Orbital Model) the electrons do not move in planetary orbits around the nucleus (Bohr Model) but rather they move through average regions called "orbitals". These orbitals differ in size, in shape, and in space orientation. The Orbital Model of the atom shows the following:

1. The total number of electrons outside the nucleus equals the number of protons in the nucleus. (Atomic number)

2. There are principal energy levels of electrons within the atom that are represented by principal quantum number (n). The principal energy levels are the same as in the Bohr model. The lowest energy level is number 1, next n=2, then n=3 etc. The number is the same as the period number on the periodic table. (n=1-7)

3. The maximum number of electrons for each principal energy level (n) is $2n^2$. This means that for the lowest energy level n=1, the maximum number of electrons is $2 \times 1^2 = 2 \times 1 = 2$. However the maximum number of electrons in the outermost level is limited to 8 and the next to outermost is 18. Complete the column for "Maximum Number of Electrons" on the separate handout.

4. The unexplained additional spectral lines mentioned above, can be explained by assuming that the principal energy levels or shells are divided into sublevels or subshells, each representing a different energy level. The total number of possible sublevels for each principal quantum number (n) equals n. However the number of occupied sublevels does not exceed 4 even when n is greater than 4. These sublevels or subshells are named s, p, d, and f, derived from the observations of line spectra. These letters are abbreviations for a series of spectral lines originally named "sharp", "principal", "diffuse" and "fundamental". The first principal quantum number (n=1) consists of 1 sublevel 1s; the second principal quantum number (n=2) has 2 sublevels 2s and 2p. Complete the column for sublevel on the handout.
   Note: Within a given principal energy level the lowest sublevel is the "s" and the highest is the "f" sublevel. The principal quantum number and one of the letters s, p, d, and f are employed to describe the energy of an electron in a particular energy level. For example the electron in the 3s sublevel is slightly lower in energy than the electrons in the 3p sublevel.

5. Each sublevel may consist of one or more orbitals with each orbital having a slightly different orientation in space. The s sublevel consists of 1 orbital, the p sublevel consists of 3 orbitals, the d sublevel consists of 5 orbitals and the f sublevel consists of 7 orbitals. The total number of orbitals within the same principal energy (n) is n^2. The first energy level (n=1) has 1^2 or 1 orbital, the second energy level (n=2) has 2^2 or 4 orbitals. Complete the column for orbitals on the handout.

6. Each electron of an atom occupies an orbital. Each orbital can hold no more than two electrons, each of different spin. (Pauli Exclusion Principal) Since the s sublevel has 1 orbital, and an orbital holds 2 electrons, the maximum number of electrons in an s sublevel is 2. The p sublevel has 3 orbitals, each orbital holds 2 electrons, giving a maximum number of p electrons as 6. The d sublevel has 5 orbitals giving a maximum number of d electrons as 10. The f sublevel has 7 orbitals giving a maximum number of electrons as 14. Complete the column marked total number of electrons. Compare this column with the maximum number of electrons.

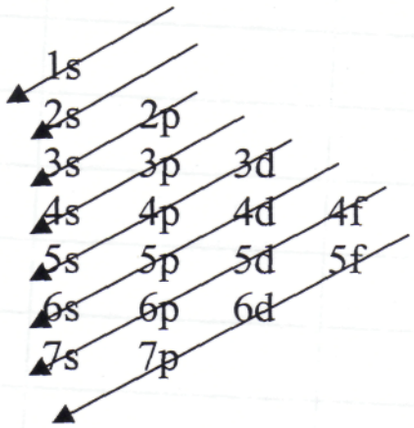
MODERN ORBITAL MODEL

Principal Energy Level (n)	Maximum # of Electrons (____)	# of Sublevels (____)	# of Orbitals (____) s p d f	# of Electrons s p d f =
1				
2				
3				
4				
5				
6				
7				

ELECTRON CONFIGURATION, ORBITAL NOTATION and ELECTRON DOT

Electron Configuration: Starting with hydrogen, atomic number 1, the electron configuration of the atoms in order of increasing atomic number can be built by adding one electron at a time. However irregularities exist in the electron configuration as we proceed to elements with higher atomic numbers. For example the 4s sublevel is filled before the 3d because it is lower in energy and the average probable location of the electrons in the 4s orbital is closer to the nucleus than the average probable location of the 3d orbitals.

An easy way to learn which sublevels are filled first is to use the following diagram:



For example hydrogen is $1s^1$. The "1" indicates the energy level, "s" indicates the s sublevel and the superscript "1" indicates one electron.

Orbital Notation: Use a box to indicate an orbital. A second electron is not added to an orbital until each orbital in the sublevel contains one electron. (Hund's Rule). For example hydrogen is. You use an up and down arrow in a box to indicate two electrons with opposite spin.

Electron dot: The electrons in the outermost principal energy level of an atom, which usually determines the chemical properties of the element, are referred to as **valence electrons.** The valence electrons can be represented using an electron dot notation. The **kernel** or core of the atom, which includes the nucleus and all the electrons except the valence electrons, is represented by the symbol for the element. The valence electrons are shown by dots placed around the symbol.

Each side of the symbol represents an orbital, where two dots representing two electrons go on a side, for a maximum of eight valence electrons. For example the electron dot for hydrogen is **H**·. Note: It does not matter where the dots are placed around the kernel. It does matter if they are paired or not like in an orbital.

Activity: Complete the table for elements with atomic number 1-20.

Element name/symbol	Electron Configuration	Orbital Notation	Electron dot (Lewis electron dot or E-dot)
Hydrogen (H)	$1s^1$	↑ 1s	H•

ELECTRON CONFIGURATION (LEVEL ONE)

Name _____

Electrons are distributed in the electron cloud into principal energy levels (1, 2, 3, ...), sublevels (s, p, d, f), orbitals (s has 1, p has 3, d has 5, f has 7) and spin (two electrons allowed per orbital).

Example: Draw the electron configuration of sodium (atomic #11).

Answer: $1s^2$ $2s^2$ $2p^6$ $3s^1$

↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑

Draw the electron configurations of the following atoms.

1. Cl

2. N

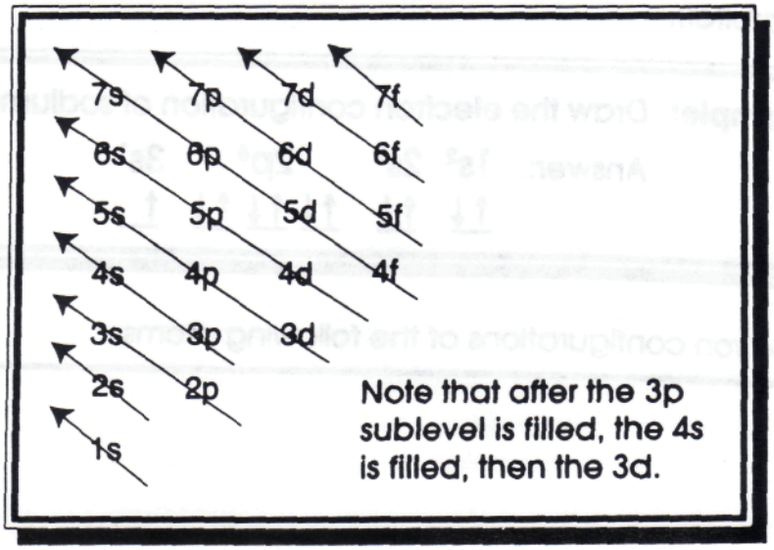
3. Al

4. O

ELECTRON CONFIGURATION (LEVEL TWO)

Name _____

At atomic number greater than 18, the sublevels begin to fill out of order. A good approximation of the order of filling can be determined using the diagonal rule.



Draw the electron configurations of the following atoms.

1. K

2. V

3. Co

4. Zr

VALENCE ELECTRONS

Name _____

The valence electrons are the electrons in the outermost principal energy level. They are always "s" or "s and p" electrons. Since the total number of electrons possible in s and p sublevels is eight, there can be no more than eight valence electrons.

Determine the number of valence electrons in the atoms below.

Example: carbon

Electron configuration is $1s^2$ $\boxed{2s^2\ 2p^2}$.

Carbon has 4 valence electrons.

1. fluorine _____
2. phosphorus _____
3. calcium _____
4. nitrogen _____
5. iron _____
6. argon _____
7. potassium _____
8. helium _____
9. magnesium _____
10. sulfur _____
11. lithium _____
12. zinc _____
13. carbon _____
14. iodine _____
15. oxygen _____
16. barium _____
17. aluminum _____
18. hydrogen _____
19. xenon _____
20. copper _____

LEWIS DOT DIAGRAMS

Name _____

Lewis diagrams are a way to indicate the number of valence electrons around an at

Na·, :Cl:, ·N:
are all examples of this type of diagram.

Draw Lewis dot diagrams of the following atoms.

1. calcium
2. potassium
3. argon
4. aluminum
5. bromine
6. carbon
7. helium
8. oxygen
9. phosphorus
10. hydrogen