Chapter 3. Elements, Atoms, Ions, and the Periodic Table

The Periodic Law and the Periodic Table

In the early 1800's many elements had been discovered and found to have different properties. In 1817 **Döbreiner**'s triads –with regularly varying properties: (Mg, Ca, Ba) (F, Cl, Br) and (S Se Te).1865: **Newlands** – "law of octaves", about 55 elements: pattern of reactivity follows after 8 elements. However, no one had found a clear "order" in their properties until **Mendeleev**, Dmitri (1834-1907) arranged 63 then known elements in the order of **increasing atomic mass** in a periodic table and showed some **chemical properties** would reappear periodically. In certain cases, he placed a lighter slightly heavier element before a lighter element so that the chemical properties of the vertical columns would be preserved. Even though in a different and much less clear form **Meyer,** Lothar (1830–1895) also came up with a graph showing periodic properties similar to Medeleev.

In Mendeleev's table, there was a gap. He purposely left blank position in his table so that the consistent vertical columns with the same chemical properties would be preserved. These **missing elements** were later discovered.

The **periodic law** is an organized "map" of the elements that relates their structure to their chemical and physical properties. The periodic table is the result of the periodic law, and provides the basis for prediction of such properties as relative atomic and ionic size, ionization energy, and electron affinity, as well as metallic or non-metallic character and reactivity.

The modern periodic table exists in several forms. The most important variation is in group numbering. The tables in the text use the two most commonly accepted numbering systems.

Numbering Groups in the Periodic Table

Periods and Groups

Periods are the **horizontal rows** of elements in the periodic table; the **columns** represent **groups** or **families**.

Elements in a **vertical group** have similar chemical properties. The vertical groups are currently named by numbers ranging from 1 to 18. An older way to identify the vertical groups is to use a Roman number and the capital letters A or B. Vertical groups of **main group elements** (or representative elements) were given a Roman numeral plus the letter **A**. Vertical groups of **transition elements** were given a Roman numeral plus the letter **B**.

Representative elements are elements that always **lose or gain the same number of electrons** in chemical reactions.
Transition elements are elements that can lose or **gain variable numbers of electrons** in chemical reactions.

The **lanthanide series** and the **actinide series** are parts of periods 6 and 7, respectively, and groups that have been named include the alkali metals, the alkaline earth metals, the halogens, and the noble gases. Group A elements are called representative elements; Group B elements are transition elements. Metals, metalloids, and nonmetals can be identified by their location on the periodic table.

These groups are number from 1 - 18, left to right and groups have their **Roman numbers** and **A** or **B** classification..

	Name	Elements	Common Valence Electron Configuration
Group 1 (IA) -	Alkali metal:	Li, Na, K Rb, Cs, Fr	ns^1
Group 2 (IIA) -	Alkaline earth metals:	Be, Mg, Ca, Sr, Ba, Ra	ns^2
Group 13 (IIIA) -	No specific name	B, Al, Ga, In, Tl	$ns^2\ 3p^1$
Group 14 (IVA) -	No specific name	C, Si, Ge, Sn, Pb	$ns^2\ 3p^2$
Group 15 (VA) -	No specific name	N, P, As, Sb, Bi	$ns^2\ np^3$
Group 16 (VIA) -	No specific name	O, S, Se, Te, Po	$ns^2\ np^4$
Group 17 (VIIA) -	Halogens:	Cl, Br, I, At	$ns^2\ np^5$
Group 18 (VIIIA) -	Noble gases:	He, Ne, Ar, Kr, Xe, Rn	$ns^2\ np^6$

In addition to groups in the periodic table there are three **blocks** of elements called **transition elements (**which are labeled with **B)**, **Lanthanides** and **Actinides** (placed bottom of the table.

Metals, Nonmetals and Metalloids

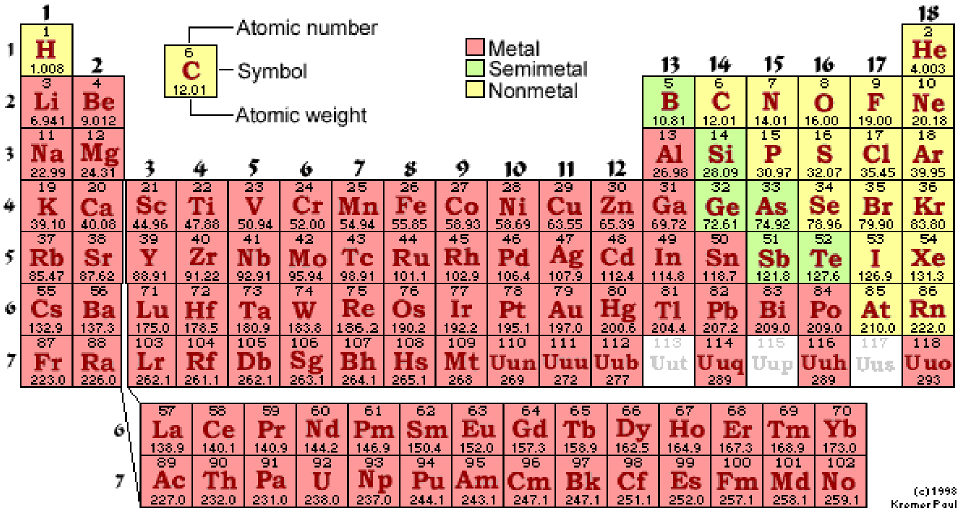
Most of the elements in the periodic table are **metals**. Note the stair step line in the periodic table. Elements to the left of the line are metals. Elements to the right of the line are **nonmetals**. In between metal and non-metals there are **semi-metals** or **metalloids**. Metals lose electrons and nonmetals gain electrons.

Ionic Compounds are formed when electrons are exchanged in this way between metals and nonmetals.

Covalent or Molecular Compounds are formed between **non metals** and **non metals** react by sharing electrons.

Atomic Number and Atomic Mass

The **atomic number (Z)** of an element represents the number of protons in the nucleus of atoms of that specific element. No two element has that same number of protons. Atomic number after it was discovered proved to be the best order without any discrepancies to arrange the elements in the periodic table and is shown on top of the space for each element. The atomic number will always be a whole number value without decimals. At the bottom **average atomic mass** calculated based on isotopes of each elements is written.



Problem: Pick the a) representative elements, b) transition elements, c) inert gas elements, d) elements that from anions, e) semi- metals, and f) elements that from cations from the following list: Ca, Si, K, Ar, Cu, Fe Zn, Ge, Kr, Cl, O, F.

Answer:

a) representative elements: Ca, Cl, O, F
b) transition elements: Cu, Fe
c) inert gas elements: Ar, Kr
d) elements that from anions: O, F
e) semi- metals: Si, Ge
f) elements that from cations: Ca, K, Cu, Fe Zn

Look on a periodic chart at the elements listed below. Do you know how to find an elements atomic number?

Problem: Use your periodic table to find the symbol, atomic number and atomic mass rounded to two decimal place of each of the following elements:

a) Magnesium b) Neon c) Selenium d) Gold

Answer

Mg, atomic number = 12, mass = 24.31 amu

Ne, atomic number = 10, mass = 20.18 amu
Se, atomic number = 34, mass 78.96 amu
Au, atomic number 79, mass 197.0 amu

Electron Arrangement and the Periodic Table

Bohr concluded that the energy levels of an atom can handle only a certain number of electrons at a time.

The Quantum Mechanical Atom

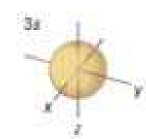
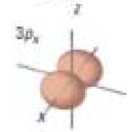
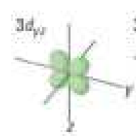
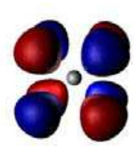
J. J. Thomson had demonstrated the particle properties of the electron earlier.
Because electrons can exhibit diffraction patterns, they have a dual nature of both wave and particle.
In 1924, **Louis de Broglie** suggested that the electron should have wave properties.
Light waves exhibit "diffraction."
Erwin Schrodinger developed equations to describe the regions around the nucleus where electrons had the probability of being 95% of the time.

These regions of high probability for finding an electron around the nucleus were called **orbitals**. Three dimensional models of the probability regions or orbitals can be constructed. Electron cloud representations are used to show the space that can be occupied by electrons in different energy levels.

Building Atoms by Orbital Filling

Schrodinger's work showed that each orbital could have a maximum of two electrons. Energy levels could contain different numbers of orbitals. Energy levels further from the nucleus can accommodate more orbitals than energy levels nearer the nucleus.
Energy levels can have sublevels when multiple orbitals are present.

Orbital Shapes

Sub-level	**Shape**	**# of orbitals/energy level**	**Picture**
s	spherical	1	
p	dumbbell	3	
d	complex	5	
f	very complex	7	

Energy Levels	n=1	n=2	n=3	n=4
number of sublevels	one	two	three	four
Sublevel Names	s	s and p	s, p and d	s. p, d and f
Sublevels and orbitals	1s (1)	2s(1) 2p(3)	3s(1) 3p(3) 3d(5)	4s(1) 4p(3) 4d(5) 4f(7)
Number orbitals	1	4	9	16
maximum number of electrons per sublevel	$2(2n^2)$	$2 + 6= 9(2n^2)$	$2 + 6 + 10= 18\ (2n^2)$	$2 + 6 + 10 + 14 = 32\ (2n^2)$

The maximum number of electrons that can be in an energy level is $2n^2$, where n is equal to the energy level being considered.

Energy Level	maximum number of electrons in an Energy Level	# of Sublevels	sublevels names	maximum number of electrons per sublevel
n = 1	2	1	s	2
n = 2	8	2	s, p	2, 6 =8
n = 3	18	3	s, p, d	2, 6, 10 =18
n = 4	32	4	s, p, d, f	2, 6, 10, 14 =32

Problem:
How many electrons are found:
Within principle shells? a) n = 1 b) n = 2 c) n = 3 d) n =4 e) n = 5
Answer:

a. $n = 1; 2n^2 = 2(1)^2 = 2$
b. $n = 2; 2n^2 = 2(2)^2 = 8$
c. $n = 3; 2n^2 = 2(3)^2 = 18$
d. $n = 4; 2n^2 = 2(4)^2 = 32$
e. $n = 5; 2n^2 = 2(5)^2 = 50$

Problem: With in a sub-shells: a) s, b) p c) d, d) f
Answer: a) s = **2**, b) p=**6** c) d=**10**, d) f= **14**

Problem: With in a Orbital?
Answer: Two electrons.

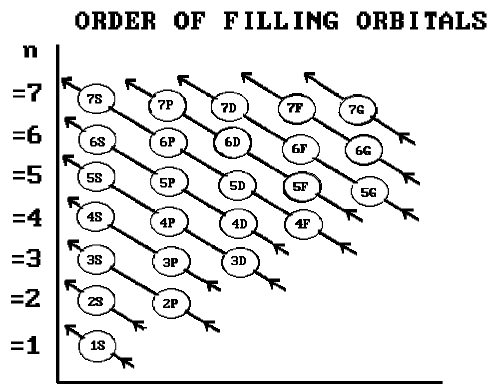
Energy Levels and Sublevels

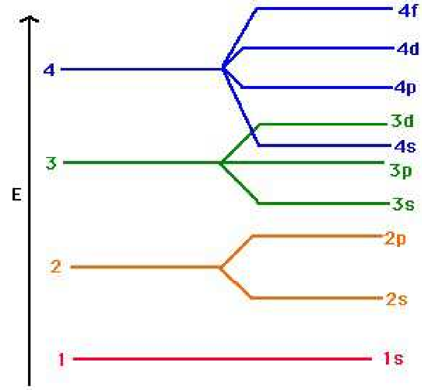
A **sublevel** is a part of a **principal energy level** and is designated *s, p, d,* and *f.* Each sublevel may contain one or more orbitals, regions of space containing a maximum of two electrons with their spins paired.

Schrodinger's work showed that

- Eeach orbital could have a maximum of two electrons.
- Energy levels could contain different numbers of orbitals.
- Energy levels further from the nucleus can accommodate more orbitals than energy levels nearer the nucleus.

- Energy levels can have sublevels when multiple orbitals are present.





Building Atoms by Orbital Filling

Amazingly, the **"electron configurations" of the elements are "embedded" in the Periodic Table.**
Honk, if you can see this "embedded" information in the Periodic Table?
Analogy: The periodic table is actually a packing slip that tells how the electrons are packed around the nucleus.

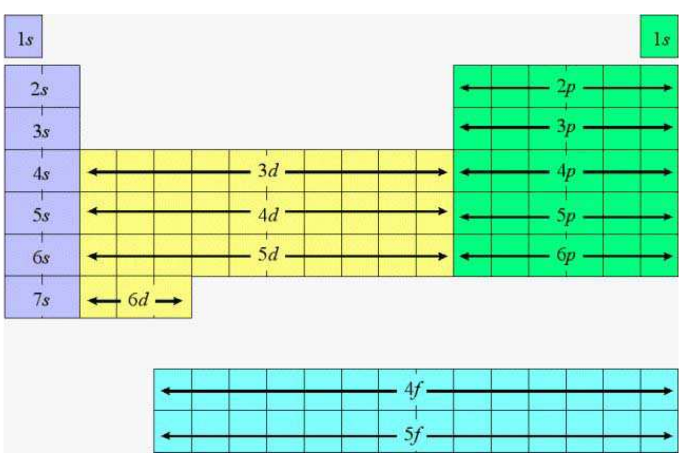
Electronic Configuration - the arrangement of electrons, in orbits or orbitals, around a nucleus of an atom.

Electron Configuration and the Aufbau (Building Up) Principle

A scheme used by chemist to obtain electronic configuration of a multi-electron atom in the ground state by filling atomic orbital starting with lowest energy.

1s 2s 2p3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d… (**building up principle**)

If two or more orbitals exist at the same energy level, they are degenerate. Do not pair the electrons until you have to.



Problem: What is the electron configuration of a) K and b) P?
Answer:
Using Aufbau principle or periodic table

a. Potassium: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$
b. Phosphorus: $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$

Problem: Examine the electron configurations below, and name the element.
$1s^2\ 2s^2$ $1s^2\ 2s^2\ 2p^3$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

Answer: Going through the periodic table.
$1s^2\ 2s^2$ (**He**) $1s^2\ 2s^2\ 2p^3$ (**N**) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ (**Al**)

Problem: If a neutral atom in its ground state contains only 5 electrons in its outermost p sublevel, it is an atom in what "vertical group" of elements?
Answer: group 17 or VIIB or Halogen family.

Problem: If a neutral atom in its ground state contains 2 electrons in its outermost s sublevel, it is an atom in what "vertical group" of elements?
Answer: group 2 or IIA or Alkaline Earth family.

Problem: State what is similar and what is different about the electron configuration of fluorine and chlorine.
Answer: F: $1s^2\ 2s^2\ 2p^5$ Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ valance shell electron configuration is similar but electron configurations are different.

Problem: Fluorine and chlorine have similar chemical properties. Oxygen and sulfur have similar chemical properties. However, oxygen and sulfur have chemical properties different from fluorine and chlorine. What does electron configuration have to do with this observation?
Answer: F: $1s^2\ 2s^2\ 2p^5$ Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ both have same vala.nce electron configurations and similar chemical properties.
O: $1s^2\ 2s^2\ 2p^4$ S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ both have same valance electron configurations and similar chemical properties..
However, two groups F, Cl: $ns^2\ np^5$ and O, S: $ns^2\ np^4$ have different valance electron configurations creating different chemical properties.

Valence Electrons

The **outermost electrons** in an atom are valence electrons. For representative elements, the number of valence electrons in an atom corresponds to the group or family number. If atoms of different elements have the same electron arrangement in their valence shell electrons, then they can have similar chemical properties even if their atomic numbers or atomic masses are quite different. Metals tend to have fewer valence electrons than nonmetals. **Valence** electrons are involved in chemical interactions and bonding (valence comes from the Latin valere, "to be strong"). Valence shell electrons are available to be lost, gained, or shared in chemical reactions.

Problem: How many total electrons and valance electrons are in the following atoms:
a) K, b) F, c) P, d) O and e) Ca
Answer
For counting valance electrons go to the period the element is found and count (**excluding transition element blocks**) from left to right until element is found.

a. Total electrons = 19 (same as atomic number), valence electrons = 1
b. Total electrons = 9 (same as atomic number), valence electrons = 7
c. Total electrons = 15 (same as atomic number), valence electrons = 5
d. Total electrons = 8 (same as atomic number), valence electrons = 6

e. Total electrons = 20 (same as atomic number), valence electrons = 2

Abbreviated Electron Configurations
Abbreviated electronic configuration is separating valance electrons from core electrons and designating core electrons as a noble gas.

E.g. What is the abbreviated electron configurations of a) K, b) P and Sn?
Answer
First, obtain the electron configuration then find the valence electrons.

a. Potassium (K): $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^1$
b. Phosphorus (P): $1s^2, 2s^2, 2p^6, 3s^2, 3p^3$
c. Tin (Sn): $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^2$

Second, lump all non valance electrons as "core" abbreviated as a noble gas con figuration.

a. Ar: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$ = [Ar]
b. Ne: $1s^2, 2s^2, 2p^6, 3s^2$ = [Ne]
c. Kr: $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$, $3d^{10}$, $4p^6$ = [Kr]

Final answer

a. Potassium (K): [Ar] $4s^1$
b. Phosphorous (P): [Ne] $3s^2, 3p^3$
c. Tin (Sn): [Kr] $5s^2$, $4d^{10}$, $5p^2$

Electron configuration of the elements is predictable, using the **Aufbau Principle**. Knowing the electron configuration, we can identify valence electrons and begin to predict the kinds of reactions that the elements will undergo.

Elements in the last family, the noble gases, have either two or eight valence electrons. Their most important properties are their extreme stability and lack of reactivity. A full energy level is responsible for this unique stability.

The Octet Rule
Noble gases are non-reactive because they all have a complete outer shell. An atom chemically reacts to fill its valance shell. A full valance shell contains eight electrons there fore the name **octet**. The octet rule tells us that in chemical reactions atoms of elements will gain, lose or share the minimum number of electrons necessary to achieve the electron configuration of the nearest **noble gas**.
octet rule - the rule which predicts that atoms form the most stable molecules or ions when they are surrounded by eight electrons in their highest occupied energy (valance) level.

Electronic configuration of ions
Series of negative ions, noble gas atom, and positive ions with the same number electrons and electronic configuration. Electron configuration of ions is obtained by adding more electrons (**anions**) or removing electrons (cations) from a neutral atom. In the process atoms achieves a noble gas electron configuration.
Group 1 (or IA), Alkali Metals have one valence electron.
They all form +1 cations when the single valence electron is lost.
Metals lose electrons and achieve electron configuration of preceding noble gas.
E.g. Potassium (K):

$K \rightarrow K^+$ (cation) + e^-

Oxygen (O):

$O + 2e^- \rightarrow O^{2-}$ (anion)

Metallic elements tend to form cations and nonmetals form anions that are isoelectronic with their nearest noble gas neighbor.

Isoelectronic electronic configurations

If atom and a cation or anion have same number of electrons they are called isoelectronic.

E.g. K^+ and Ar

O^{2-} and Ne

Problem: Which of the following are isoelectronic: F^-, Cl^-, K^+, Ar

Answer:

a. F^-, 10e⁻; Cl^-, 18 e⁻; Not isoelectronic
b. K^+, 18 e⁻; Ar, 18e⁻; Isoelectronic

Ion Formation and the Octet Rule

Metals lose electrons and achieve a an octet of valance electrons similar to electron configuration of preceding noble gas.

E.g. Potassium (K): [Ar] $4s^1$

K ([Ar] $4s^1$) → K^+ ([Ar]) + e^-

Oxygen (O): [He] $2s^2 2p^4$

O ([He] $2s^2 2p^4$) + $2e^-$ → O^{2-} ([Ne])

a. I^- (54 e⁻) = Xe= $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6$
b. Ba^{2+} (54 e⁻)= Xe= $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6$
c. Se^{2-} (36 e⁻) =Kr = $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6$
d. Al^{3+} (10 e⁻)=Ne = $1s^2, 2s^2, 2p^6$

Trends in the Periodic Table

Atomic Size

Atomic size increases from top to bottom but decreases from left to right in the periodic table. Cations are smaller than the parent atom. Anions are larger than the parent atom. Ions with multiple positive charge are even smaller than their corresponding monopositive ion; ions with multiple negative charge are larger than their corresponding less negative ion.

Problem: Arrange the following list of elements in order of increasing atomic size.

a) Al, Si, P, Cl, S
b) In, Ga, Al, B, Tl
c) Sr, Ca, Ba, Mg, Be
d) O,N, Sb, Bi, As

Answer:

a. (Smallest) Cl, S, P, Si, Al (Largest)
b. (Smallest) B, Al, Ga, In, Tl (Largest)
c. (Smallest) Be, Mg, Ca, Sr, Ba (Largest)
d. (Smallest) N, P, As, Sb, Bi (Largest)

Ionization Energy

The energy required to remove an electron from an atom in the gas phase.

The energy required to remove an electron from the atom is the ionization energy. Descending a group, the ionization energy decreases. Proceeding across a period, the ionization energy increases.

Problem: Arrange the following list of elements in order of increasing ionization energy.

a) N,F, O
b) Li, K, Cs
c) Br, I, Cl

Answer:

a) (Smallest) N, O, F(Largest)
b) (Smallest) Cs, K, Li (largest)
d) (Smallest) Cl, Br, I (Largest)

Electron Affinity

The energy released when a single electron is added to neutral atom in the gaseous state is known as the electron affinity. Electron affinities generally decrease proceeding down a group and increase proceeding across a period.

Exceptions exist for periodic trends. They are generally small anomalies, and do not detract from the predictive power of the periodic table.

Problem: Arrange the following list of elements in order of increasing ionization energy.

a. Na, Li, K
b. Br, F, Cl
c. S, O, Se

Answer:

a. (Smallest) Li, Na, K (Largest)
b. (Smallest) F, Br, Cl (Largest)
c. (Smallest) Se, S, O (Largest)

Ion Size

Ions follows same trends as for atomic radius in a group, fro example taking oxide and sulfide ion: radius of $O^{2-} < S^{2-}$.

Cation or positive ions have fewer electrons than neutral atom and nuclear charge being same attract remaining electrons strongly making cation smaller than the neutral atom

Anions or negative ions larger than neutral atom. Anions are larger than the atoms from which there are formed. Adding electrons to an atom increases the repulsion between electrons. Anion has a harder time holding on to the electrons.

CHEM 120 Homework 3. Chapter 3

1. In the modern periodic table, the elements are arranged according to increasing ________.
a. atomic masses b. number of neutrons c. atomic number d. mass number

2. How many periods are found on the periodic table?
a. 2 b. 7 c. 18 d. 32

3. Which period contains the element Cesium?
a. 2 b. 4 c. 6 d. 7

4. Where are the alkaline earth metals located on the periodic table?
a. Group 1 (IA) b. Group 2 (IIA) c. Group 13 (IIIA) d. Group 14 (IVA) e. Group 17 (VIIA)

5. Which one of the following is not a representative element
a. Na b. As c. Ca d. Fe e. Cl

6. How many orbitals are in an *s* sublevel? How many in a *p* sublevel?
a. 2;6 b. 1;1 c. 1; 3 d. 3; 5

7. Which of the following correctly gives the electron capacity of a principal energy level in terms of the number *n*?
a. n b. $2n$ c. $2n + 2$ d. n^2 e. $2n^2$

7. What requirement must be met in order for two electrons to coexist in the same orbital?
a. they go to a s orbital
b. they go to a p orbital
c. they must have opposite spins
d. they must have parallel spins

8. How many valence electrons are present in an atom of silicon?
a. 2 b. 3 c. 4 d. 5 e. 7

9. The electronic configuration in an atom of argon,
a. $1s^22s^2$
b. $1s^22s^22p^6$
c. $1s^22s^22p^63s^23p^6$
d. $1s^22s^22p^63s^23p^64s^23d^{10}4p^6$

10. Common valence electron configuration of halogens
a. ns^1 b. ns^2 c. $ns^2\ 3p^2$ d. $ns^2\ np^3$ e. $ns^2\ np^5$

11. What is not isoelectronic with K^+?
a. S^{2-} b. Ar c. Cl^- d. Na^+ e. Ca^{2+}

12. Which of the following atoms has the biggest size (radius)?
a. Na b. Al c. Cl d. Rb e. I

13. Which of the following elements has the highest ionization energy?
a. Li b. B c. O d. F e. Ne

14. Which one of the following elements has the highest electron affinity?
a. Li b. K c. Kr d. O e. Cl

15. What charge is found on a ion from Al?
a. +1 b. -2 c. +3 d. -3

Sample Test Chapter 3

1. Which two scientists in 1869 arranged the elements in order of increasing atomic masses to form a precursor of the modern periodic table of elements?
Ans. Mendeleev and Meyer

2. Who stated that the elements, when arranged according to their atomic masses, showed a distinct periodicity of their properties?
Ans. Dimitri Mendeleev

3. In the modern periodic table, the elements are arranged according to what system?
Ans. increasing atomic number

4. The modern periodic law states that the physical and chemical properties of the elements are periodic functions of what property?
Ans. atomic number

5. What do we call the horizontal row of elements on the periodic table?
Ans. periods

6. How many periods are found on the periodic table?
Ans. seven

7. Which period contains the element sodium?
Ans. three

8. What do we call the columns of elements on the periodic table?
Ans. groups

9. What number for an atom gives the number of electrons and protons found in that atom?
Ans. atomic number

10. Where are the alkaline earth metals located on the periodic table?
Ans. Group IIA (2)

11. What is the general name given to the elements of Group VIIA (17)?
Ans. Halogens

12. What term is used for the elements straddling the "staircase" boundary between the metals and nonmetals?
Ans. Metalloids or semi-metals

13. For a representative element, how can we deduce the number of valence electrons in a neutral atom from the position of the element in the
Periodic Table?
Ans. the group number (Roman numbers with Bs) is also the number of valence electrons

14. How many orbitals are in an *s* sublevel? How many in a *p* sublevel?

Ans, 1; 3

15. In what way(s) are the three orbitals in the 2p sublevel similar; in
what way(s) are they different?
Ans. they have the same shape and the same energy; they are oriented
differently in space

16. What requirement must be met in order for two electrons to coexist in
the same orbital?
Ans. they must have opposite spins

17. State the Aufbau Principle.
Ans. Electrons occupy the available orbital of lowest energy first.

18. How many electrons are present in an atom of silicon?
Ans. Fourteen

19. Give the electronic configuration in an atom of argon, element number
18.
Ans. 1*s*22*s*22*p*63*s*23*p*6

20. Give the electronic arrangement in an atom of strontium, element number 38.
Ans. 1*s*22*s*22*p*63*s*23*p*64*s*23*d*104*p*65*s*2

21. How many electrons are present in a chloride ion?
Ans. Eighteen

22. State the Octet Rule.
Ans. Elements tend to react in such a way as to attain the electron configuration of the atoms of the noble gas nearest to them in the Periodic Table

23. Give the name of a Group IA (1) ion that has the following electronic arrangement:
$1s^22s^22p^6$
Ans. sodium ion

24. Give the name of a VIIA (17) ion that has the following electronic
arrangement: 1*s*22*s*22*p*63*s*23*p*6
Ans. chloride

25. What ion carries a 2- charge and is isoelectronic with K^+?

Ans. S^{2-}

26. Give the complete electronic arrangement of a sulfide ion, S^{2-}.
Ans. 1*s*22*s*22*p*63*s*23*p*6

27. Atoms with the biggest radii occur in the _______ ____ region of the Periodic Table.
Ans. bottom left

28. How would you expect an Al^{3+} ion to compare in size with an Al atom? Explain why?
Ans. The ion will be much smaller. In forming the ion, the atom loses all its outermost electrons. The net positive charge on the ion ensures that all the electrons in the ion are strongly attracted to the nucleus, keeping the ion small.

29. Which group of elements has the highest ionization energies? Which group has the lowest?
Ans. Group VIIIA (18) are highest; Group IA (1) are the lowest.

30. Explain what is meant by electron affinity.
Ans. It is the energy released when a neutral atom gains an electron
to form an anion.

31. In Mendeleev's table of the elements, they were arranged according to

A. atomic number
B. mass number
C. atomic mass
D. neutron number
E. density

Ans. C. atomic mass

32. The modern periodic table is arranged according to what property?

A. atomic number
B. mass number
C. atomic mass
D. neutron number
E. density

Ans. A. atomic number

33. What do we call a complete horizontal row of elements on the periodic
table?

A. group
B. period
C. family
D. representative elements
E. transition elements

Ans, B

34. What are all the elements in the A-groups often called?

A. transition elements
B. lanthanides
C. metals
D. non-metals
E. representative elements

Ans. E

35. Which of the following elements is a metalloid?

A. C B. Ge C. Pb D. N E. P

Ans. B

36. Where are the alkali metals located on the periodic table?

A. representative elements
B. transition metals
C. Group IA (1)
D. Group IIA (2)
E. Group IIIA (3)

Ans. C

37. How many valence electrons are in an atom of carbon?

A.8 B. 6 C. 4 D. 1 E. 0

Ans. C

38. What is the lowest energy sublevel of a principal level?

A. *d* B. *e* C. *f* D. *s* E. *p*

Ans. D

39. How many sublevels are there in the third principal energy level?

A. 3 B. 2 C. 1 D. 0 E. 4

Ans. A

40. How many orbitals are there in a *p* sublevel?

A. 2 B. 3 C. 1 D. 0 E. 4

Ans. B

41. Which of the following correctly gives the electron capacity of a
principal energy level in terms of the number *n*?

A. n B. $2n$ C. $2n + 2$ D. n^2 E. $2n^2$

Ans. E

42. What is the electron configuration of sulfur, atomic number 16?

A. $1s^2 1p^6 2s^2 2p^6$
B. 1*s*22*s*22*p*62*d*6
C. 1*s*22*s*22*p*63*s*23*p*4
D. 1*s*22*s*22*p*63*s*23*d*4
E. 1*s*22*s*22*p*63*s*22*d*4

Ans. C

43. Which one of the following electron configurations is appropriate for a
normal atom?

A. 1*s*12*s*1
B. 1*s*22*s*1
C. 1*s*22*s*22*p*8
D. 1*s*22*s*22*p*43*s*1
E. 1*s*22*s*22*p*63*d*1

Ans. B

44. Which of the following elements is most likely to form a 3+ ion?

A. Li B. K C. Al D. N E. Cu

Ans. C

45. Give the complete electronic configuration of a sodium ion.

A. 1*s*22*s*22*p*5
B. 1*s*22*s*22*p*6
C. 1*s*22*s*22*p*63*s*1
D. 1*s*22*s*22*p*63*s*2
E. 1*s*22*s*22*p*63*s*23*p*64*s*1

Ans. B

46. Which of the following ions does not follow the octet rule?

A. Na+ B. Ca2+ C. Al3+ D. N3- E. Cl2-

Ans. E

47. Which of the following atoms has the biggest size (radius)?

A. Na B. Al C. Cl D. Rb E. I

Ans. D

48. Which of the following elements has the highest ionization energy?

A. Li B. B C. O D. F E. Ne

Ans. E

49. Which of the following elements has the lowest ionization energy?

A. Li B. B C. O D. F E. Ne

Ans. A

50. The electron affinity is

A. the energy required to remove an electron from an isolated atom
B. the force between two electrons in the same orbital
C. the force between two ions of opposite charge
D. the energy released when an isolated atom gains an electron
E. the attraction of an atom for an electron in a chemical bond

Ans. D

51. Which one of the following elements has the highest electron affinity?

A. Li B. K C. Kr D. O E. Cl

Ans. E

52. T F In Mendeleev's table, the elements were arranged according to
their atomic numbers.

Ans. F

53. T F There are nine periods on the periodic table.

Ans. F

54. T F Sulfur (S) is one of the representative elements.

Ans. T

55. T F Platinum (Pt) is a lanthanide element.

Ans. F

56. T F Tin (Sn) is a metalloid.

Ans. F

57. T F Valence electrons are involved when atoms form bonds.

Ans. T

58. T F There are a maximum of 50 electrons in principal energy level
number five.

Ans. T

59. T F Atoms of the noble gas elements, Group VIII A (18), do not form
bonds with any other elements.

Ans. F

60. T F There are eight valence electrons in a chloride ion.

Ans. F

61. T F The ions formed from Group IIA (2) atoms have charges of 2+.

Ans. T

62. T F Cations tend to be formed from metal atoms, while anions are
formed from non-metal atoms.

Ans. T

63. T F The atoms of smallest radius are those of elements in top left
hand part of the periodic table.

Ans. F

64. T F The halogens (Group VII A (17)) have the lowest ionization
energies of any group in the periodic table.