# CHAPTER 2

## Knowledge and Comprehension Problems:

**Problem 2.1** Describe the laws of (a) multiple proportions and (b) mass conservation as related to atoms and their chemical properties.

**Answer 2.1:** (a) The law of multiple proportions states that atoms of one pure substance are different from the atoms of other pure substances and, when combined in specific simple fractions, form different compounds. (b) The law of mass conservation states that a chemical reaction is explained by separation, combination, or rearrangement of atoms, and that a chemical reaction does not lead to creation or destruction of matter.

**Problem 2.2** How did scientists find out that atoms themselves are made up of smaller particles?

**Answer 2.2:** Henri Becquerel and Marie and Pierre Curie showed that some atoms spontaneously emit rays and named this phenomenon *radioactivity*. The radiation was shown to consist of · (alpha), · (Beta), and · (gamma) rays. It was also shown that a and b particles have both charge and mass while g particles have no detectable mass or charge.

**Problem 2.3** How was the existence of electrons first verified? Discuss the characteristics of electrons.

**Answer 2.3:** Joseph J. Thompson, using cathode ray tube experiments, concluded that atoms in all matter are made of smaller particles that are negatively charged. The negatively charged plate (cathode) emits an invisible ray that is attracted by the positively charged plate (anode). The invisible ray is called a cathode ray and is made up of electrons. He also calculated the ratio of mass to charge of these electrons to be 5.60 X 10-19 g/C where *Coulomb*, C, is the unit of electrical charge. Robert Millikan, in his oil-drop experiments, determined the fundamental quantity of charge or the charge of an electron (regardless of the source) to be 1.60 X $10^{-19}$ C. For an electron, this quantity of charge is represented by -1. Using the ratio of mass to charge of the electron measured by Thompson and the charge of the electron measured by Millikan, the mass of an electron was determined to be 8.96 X $10^{-28}$ g.

**Problem 2.4** How was the existence of protons first verified? Discuss the characteristics of protons.

**Answer 2.4:** Ernest Rutherford bombarded a very thin foil of gold with positively charged a are slightly deflected, and a few are either largely deflected or completely bounce back. He concluded that 1) most of the atom must be made up of empty space (thus most particles pass through without deflection) and 2) a small neighborhood at the center of the atom, the *nucleus*, houses positively charged particles of its own. He suggested that those alpha particles that deflected intensely or bounced back must have interacted closely with the positively charged nucleus of the atom. The positively charged particles in the nucleus were called *protons*. It was later determined that the proton carries the same quantity of charge as an electron but opposite in sign and has a mass of 1.672 X $10^{-24}$ g (1840 times the mass of the electron). For a proton this quantity of charge is represented by +1. Also, since atoms are electrically neutral, they must have an equal number of electrons and protons.

**Problem 2.5** What are the similarities and differences among protons, neutrons, and electrons? Compare in detail.

**Answer 2.5:** Using the information in Table 2.1, one can summarize that protons and neutrons have significantly higher mass, and basically constitute the total mass of the atom (the mass of electron is minimal in comparison). On the other hand, the charge of the atom comes equally from its electrons (negative) and protons (positive). Neutrons are not charged.

| Particle | Mass (g) | Charge | |
|---|---|---|---|
| | | Coulomb (C) | Charge Unit |
| Electron | 9.10939 X $10^{-28}$ | -1.06022 X $10^{-19}$ | -1 |
| Proton | 1.67262 X $10^{-24}$ | +1.06022 X $10^{-19}$ | +1 |
| Neutron | 1.67493 X $10^{-24}$ | 0 | 0 |

**Problem 2.6** One mole of iron atoms has a mass of 55.85 grams, without any calculations determine the mass in amu of one iron atom.

**Answer 2.6:** One mole of iron corresponds to the number of atoms needed to create a mass in units of grams (55.85 grams) numerically equal to the atomic mass in amu of the substance under consideration. Thus one atom of iron has an atomic mass of 55.85 amu.

**Problem 2.7** One atom of oxygen has a mass of 16.00 amu, without any calculations determine the mass in grams of one mole of oxygen atoms.

**Answer 2.7:** One mole of oxygen corresponds to the number of atoms needed to create a mass in units of grams (16.00 grams) numerically equal to the atomic mass in amu of the substance under consideration. Thus one mole of oxygen has a weight of 16 grams.

**Problem 2.8** Define a) atomic number, b) atomic mass, c) atomic mass unit (amu), d) mass number, e) isotopes, f) mole, g) relative atomic mass, h) average relative atomic mass, and i) Avogadro's number.

**Answer 2.8:**

a) atomic number – the number of protons in the nucleus of an atom is the atomic number (Z).
b) atomic mass – the mass of one atom of a substance expressed in amu.
c) atomic mass unit (amu) - one amu is defined as exactly 1/12th the mass of a carbon atom with 6 protons and 6 neutrons.
d) mass number - the sum of protons and neutrons in a nucleus of an atom.
e) isotopes – atoms with the same atomic number but different mass numbers.
f) mole - one mole or gram-mole (mol) of any element is defined as the amount of substance that contains $6.02 \times 10^{23}$ atoms.
g) relative atomic mass - the mass in grams of one mole of an element is called the *relative atomic mass*, *molar mass*, or the *atomic weight*.
h) average relative atomic mass
i) Avogadro's number – number of atoms in one mole of an element.

**Problem 2.9** Explain the law of chemical periodicity.

**Answer 2.9:** The law of chemical periodicity states that the properties of elements are functions of their atomic number in a periodic manner.

**Problem 2.10** What is the nature of visible light? How is the energy released and transmitted in visible light?

**Answer 2.10:** Light is in the form of electromagnetic radiation. Energy is released and transmitted in the form electromagnetic waves.

**Problem 2.11** (a) Rank the following emissions in increasing magnitude of wavelength: microwave oven emissions, radio waves, sun lamp emissions, x-ray emissions, and gamma ray emissions from the sun. (b) Rank the same emissions in terms of frequency. Which emission has the highest energy?

**Answer 2.11:**

(a) Increasing wave length : Gamma Ray – x-ray – sun lamp – microwave – radio wave
(b) Increasing frequency of emission: Gamma Ray– x-ray– sun lamp – microwave – radio wave. The highest frequency emission, gamma ray, has the highest energy (E = hn).

**Problem 2.12** Describe the Bohr model of the hydrogen atom. What are the shortcomings of the Bohr model?

**Answer 2.12:** Bohr suggested that electrons travel in circular paths around the nucleus with discrete angular momenta (a product of velocity and radius). Furthermore, he suggested that the energy of the electron is restricted to a specific energy level that places the electron at that fixed circular distance from the nucleus. He called this the orbit of the electron.

Bohr's model worked very well for a simple atom such as hydrogen but it did not explain the behavior of more complex (multi-electron) atoms and left many unanswered questions. Also, Bohr's model required that we know the position and speed (momentum) of a particle at a given instant. However, Werner Heisenberg proposed the uncertainty principle stating that "it is impossible to simultaneously determine the exact position and the exact momentum (product of speed and mass) of a body, for instance an electron. Heisenberg also rejected Bohr's concept of an "orbit" of fixed radius for an electron; he asserted that the best we can do is to provide the probability of finding an electron with a given energy within a given space.

**Problem 2.13** Describe $Z_{eff}$ and when is it used?

**Answer 2.13.** In multielectron atoms, each electron is influenced by other electron charges. $Z_{eff}$ is the effective (net positive) nuclear charge of an electron in a multielectron atom and is determined by subtracting the average number of electrons between the nucleus and the electron under consideration, $S$, from the number protons in the nucleus, $Z$.

$$Z_{eff} = Z - S$$

The electric charge is equal to the number of protons in the nucleus times the elementary charge. In contrast, the effective nuclear charge is the attractive positive charge of nuclear protons acting on valance electrons, which is always less than the total number of protons present in a nucleus due to the shielding effect.

**Problem 2.14** Describe Slater's rule. How is this rule applied in determination of energy associated with a transition?

**Answer 2.14.** Slater's Rule is used to determine the effective nuclear charge of an electron in a multi electron atom. It states that if the electron under consideration lies within s or p orbitals with principal quantum number n, then the electrons in n - 1 energy levels contribute 0.85, each, to *S*. The electrons in the same shell as the electron under consideration, contribute 0.35, each, to S (excluding the electron under consideration). Electrons in all other lower energy levels (n -2, n – 3, etc.), contribute 1 to *S*.

**Problem 2.15** Describe the Uncertainty Principle. How does this principle contradict Bohr's model of the atom?

**Answer 2.15:** The uncertainty principle states that it is impossible to simultaneously determine the exact position and the exact momentum (product of speed and mass) of a body. Any attempt at measurement of, for instance, position would alter the velocity and vice versa. For Bohr's theory to work, he needed the knowledge of the position and momentum of an electron simultaneously.

**Problem 2.16** Describe the following terms (give a diagram for each): a) electron density diagram, b) orbital, c) boundary surface representation, and d) radial probability.

**Answer 2.16:**

a) Electron density diagram - An array of dots representing the probability of finding an electron (electron density) of a given energy level in a given region of space (see Figure 2.7 a).

b) Orbital – not to be confused with Bohr's "orbit" is a wave function that is the solution to the wave equation. An orbital has a characteristic energy level as well as a characteristic distribution of electron density (expressed geometrically in Figure 2.8).

c) Boundary surface representation - Another way to probabilistically represent the location of an electron with a given energy level is by drawing the boundary inside which we have a 90% chance of finding that electron (see Figure 2.7 b).

d) Radial probability - radial probability also called total probability, considers the probability of the electron being at a spherical layer with respect to the volume of that layer (see Figure 2.7 c).

**Problem 2.17** Name and describe all quantum numbers.

**Answer 2.17:**

1-The Principal Quantum Number, n - represents the principal energy level and only takes on integer values of one or greater than one, n = 1, 2, 3, ... Each principal energy level is also known as a shell representing a collection of subshells and orbitals with the same principal number, n. As n increases, so does the energy of the electron under consideration, indicating that the electron is less tightly bonded to the nucleus (easier to ionize).

2-The Subsidiary Quantum Number, l - within each principal shell, n, there exists a subshell. The shape of the electron cloud or the boundary space of the orbital is determined by this number. The quantum number l may be represented by an integer ranging from 0 to n-1, or by letters.

3-The Magnetic Quantum Number, $m_l$ - represents the orientation of the orbitals within each subshell. The quantum number, $m_l$ - will take on values ranging from + l to – l.

4-The Spin Quantum Number, *ms* – represents the direction of the spin of the electron. The spin quantum number can take on either +1/2 or -1/2.

**Problem 2.18** Explain the Pauli's exclusion principle.

**Answer 2.18:** No more than two electrons can occupy the same orbital of an atom, and the two electrons must have opposite spins. In other words, no two electrons can have the same set of four quantum numbers.

**Problem 2.19** Describe (a) the nucleus charge effect and (b) the shielding effect in multi-electron atoms.

**Answer 2.19:**

(a) The higher the charge of the nucleus (more protons), the higher the attraction force on an electron, and the lower the energy of the electron (a more stable system); this is called the nucleus charge effect.

(b) The shielding effect takes place when there is more than one outer electron. In this case, the outer electrons repel each other because of their charge similarity. This repulsion energy works against the attraction energy between the nucleus and the electrons. As a result, it is easier to remove these electrons from the nucleus compared to a situation where only one outer electron exists.

**Problem 2.20** Describe the terms a) metallic radius, b) covalent radius, c) first ionization energy, d) second ionization energy, e) oxidation number, f) electron affinity, g) metals, i) nonmetals, k) metalloids, and l) electronegativity.

**Answer 2.20:**

a) Metallic radius- is a measure of the size of an atom equal to half the distance between the nuclei of two adjacent atoms in a solid metallic element.

b) Covalent radius - is a measure of the size of an atom equal to half the distance between the nuclei of the identical atoms within the covalent molecule.

c) First ionization energy (IE1)- the energy required for the removal of the outermost electron in an atom.

d) Second ionization energy - the energy required for the removal of the second outermost electron in an atom after the first outermost electron has already been removed.

e) Oxidation number - the number of outer electrons that an atom can give up or receive through the ionization process.

f) Electron affinity – the tendency to easily accept an outermost electron.

g) Metals – those elements with atoms that have low ionization energies and little to no electron affinity (Groups 1A and 2A are exclusively metallic).

i) Nonmetals – elements with atoms that have a high ionization energy and very high electron affinity (Group 6A and 7A are exclusively nonmetallic).

k) Metalloids – elements that can behave either in a metallic or a nonmetallic manner (some elements in group 3A, 4A, and 5A; see the periodic table).

l) Electronegativity - indicates the degree by which an atom attracts electrons to itself.

**Problem 2.21** Compare and contrast the three primary bonds in detail (draw a schematics for each). Explain what the driving force in the formation of such bonds or, in other words, why do atoms want to bond at all?

**Answer 2.21:**

**Ionic Bonds -** metals and nonmetals bond through electron transfer and ionic bonding. Ionic bonding is typically observed between atoms with large differences in their electronegativities; for instance atoms of group 1A or 2A (reactive metals) with atoms of group 6A or 7A (reactive nonmetals). In short, one atom loses an electron and forms a cation, another atom gains the electron lost by the first atom and forms an anion. After the electron transfer process is completed, both atoms will have completed their outer electronic structure and take on the structure of a noble gas. The electrostatic attraction forces between the two ions will then hold the ions together to form an ionic bond.

**Covalent Bonds -** are typically observed between atoms with small differences in their electronegativities, and mostly between non-metals. At first, the nucleus of one atom attracts the electron cloud of the other; the atoms get closer to each other. As they get close, the two electron clouds interact, and both atoms start to take ownership of both electrons (share electrons). The atoms keep getting closer until they reach the equilibrium point in which the two atoms will form a bond by sharing their electrons, both completing their outer electronic structure, and reaching the lowest state of energy.

**Metallic Bonds –** during solidification, from a molten state, the atoms of a metal pack tightly together, in an organized and repeating manner. All the atoms contribute their valence electrons to a "sea of electrons" or the "electron charge cloud". These valence electrons (free electrons) are delocalized, move freely in the sea of electrons, and do not belong to any specific atom. The nuclei and the remaining core electrons of tightly packed atoms form a cationic or a positive core. What keeps the atoms together, in solid metals, is the attraction force between the positive ionic core (metal cations) and the negative electron cloud.

The driving for atoms to bond with other atoms through primary bonds is to lower their potential energy levels and become more stable.

**Problem 2.22** Describe the factors that control packing efficiency (number of neighbors) in ionic and covalent solids. Give an example of each type of solid.

**Answer 2.22:** In ionic solids, the number of cations that can pack around an anion (packing efficiency) is determined by two factors: 1) their relative sizes and 2) charge neutrality. Example six $Cl^-$ anions can pack around one $Na^+$ cation.

In covalent solids the number of neighbors (packing efficiency) around an atom will depend on the bond order (the number of shared pairs). The number of neighbors cannot be greater than four.

**Problem 2.23** Describe the five stages leading to formation of an ionic solid. Explain which stages require energy and which stages release energy.

**Answer 2.23:**

Stage 1: Solid metal to gaseous metal (atomization) – requires energy – $DH^1$

Stage 2: Nonmetal molecules to nonmetal atoms – requires energy – $DH^2$

Stage 3: Metal atoms, removing outer electrons – requires energy – $DH^3$

Stage 4: Nonmetal atoms, adding outer electrons – produces energy – $DH^4$

Stage 5: Formation of ionic solid from gaseous ions – produces energy – $DH^5$

**Problem 2.24** Describe a) Hess law, b) lattice energy and b) heat of formation.

**Answer 2.24:**

(a) Hess Law states that the total heat produced during formation of an ionic solid is the sum of the heats required in each five stage: $DH^0 = DH^1 + DH^2 + DH^3 + DH^4 + DH^5$.
(b) Lattice energy, $DH^5$, is the energy released when gaseous ions form solid ions due to electrostatic attraction forces.
(c) The total heat, $DH^0$, is called the heat of formation.

**Problem 2.25** Describe the terms a) shared pair, b) bond order, c) bond energy, d) bond length, e) polar and non-polar covalent bonds, and f) network covalent solid.

**Answer 2.25:**

a) shared pair – the pair of electrons active in the bond between two covalently bonded atoms.

b) bond order – the number of shared pairs between covalently bonded atoms (maximum of four)

c) bond energy – the energy required to break the bonds between covalently bonded atoms.

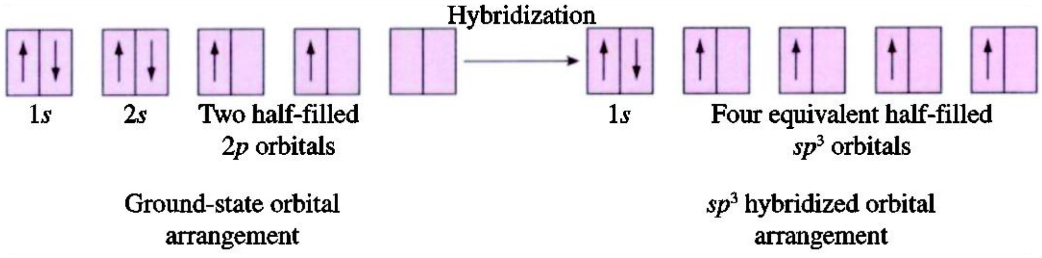
d) bond length – the distance between the nuclei of two covalently bonded atoms at the equilibrium point (point of minimum energy).

e) polar and non-polar covalent bonds – if the difference between the electronegativities of the covalently bonded atoms is zero, the bond is non-polar (no dipole). As the difference in electronegativity between the atoms increases, the bonds become polar. This means the shared electron will lean toward the more electronegative element, thus creating a dipole.

f) network covalent solid – when all atoms in a solid are bonded through covalent bonds in the form of an ordered network (example diamond).

**Problem 2.26** Explain the hybridization process in carbon. Use orbital diagrams.

**Answer 2.26:** The full 2S orbital is promoted to a 2p orbital to form 4 partially occupied hybrid SP$^3$ orbitals. Each hybrid orbital will be available for bonding with another atom for a maximum of 4.



**Problem 2.27** Describe the properties (electrical, mechanical, etc…) of materials that are exclusively made up of a) ionic bonds, b) covalent bonds, and c) metallic bonds. Name a material for each type.

**Answer 2.27:**

a) ionic solids – will be hard (difficult to indent), strong (difficult to deform permanently or fail), stiff (difficult to deform elastically), brittle (deform little before they fail), electrically insulating (in the solid state), and have a high melting temperature. Examples are MgO and CsCl.

b) network covalent solids – will be hard (difficult to indent), strong (difficult to deform permanently or fail), stiff (difficult to deform elastically), brittle (deform little before they fail), have a low thermal conductivity, and have a high melting temperature. Examples are quartz and diamond.

c) metallic solids – in a pure state, they are generally more malleable (soft and deformable), and are less stiff than ionic or covalent networked materials. Strength can be increased through alloying. Highly conductive (both heat and electricity). Examples are copper and aluminum.

**Problem 2.28** What are secondary bonds? What is the driving force for formation of such bonds? Give example of materials in which such bonds exist.

**Answer 2.28:** The bonding formed between molecules or atoms of noble gasses. Significantly weaker than primary bonds. The driving force is the electrostatic attraction between polar molecules and atoms.

**Problem 2.29** Discuss various types of mixed bonding.

**Answer 2.29:** Although different solids may be more inclined to form predominantly by a certain type of bond, other types of bonds will also be normally present. For instance, it is possible to have i) ionic-covalent, ii) metallic-covalent, and iii) metallic-ionic combinations.

**Problem 2.30** Define the following terms: a) dipole moment, b) fluctuating dipole, c) permanent dipole, d) van der Waals bonds, and e) hydrogen bond.

**Answer 2.30:**

a) dipole moment – a moment that produces a temporary or permanent separation of positive and negative charge centers in an atom or a molecule (moment = q*d).

b) fluctuating dipole - Fluctuating dipole bonding is a secondary type of bonding between atoms which contain electric dipoles. These electric dipoles, formed due to the asymmetrical electron charge distribution within the atoms, change in both direction and magnitude with time. This type of bond is electrostatic in nature, and is very weak and nondirectional.

c) permanent dipole - Permanent dipole bonding is also a secondary type of bonding between molecules possessing permanent electric dipoles. The bonds, formed by the electrostatic attraction of the dipoles, are directional in nature. They are slightly stronger than the fluctuating dipole.

d) van der Waals bonds – all bonds involving dipoles are collectively called van der Waals bonds (forces).

e) hydrogen bond – a special class of permanent dipole bonds forming between polar molecules containing hydrogen.

**Problem 2.31** The diameter of a soccer ball is approximately 0.279 m (11 inches). The diameter of the moon is $3.476 \times 10^6$ m. Give an "estimate" of how many soccer balls it will take to cover the surface of the moon (assume moon is a sphere with a flat terrain). Compare this number to Avogadro's number. What is your conclusion?

**Solution 2.31**

Surface area of the moon = $4\pi R^2$ where R is the radius of the moon

$$A_{s,\text{moon}} = 4\pi\left(\frac{3.476}{2} \times 10^6\right)^2 = 3.79 \times 10^{13}\, m^2$$

Cross-sectional area of soccer ball = $\pi R^2$

$$A_{\text{cs,ball}} = \pi\left(\frac{0.279}{2}\right)^2 = 0.0611m^2$$

Number of soccer balls required to cover the surface of the moon

$$= \frac{3.79\times 10^{13}\,m^2}{0.0611m^2} = 6.22\times 10^{14}\text{ balls}$$

Avogadro's number is $\sim$ 1 billion times larger.

**Problem 2.32** Each quarter produced by the US mint is made up of a copper and nickel alloy. In each coin, there is 0.00740 moles of Ni and 0.0886 moles of copper. (a) What is the total mass of a quarter? (b) What percentage of the mass of a quarter is nickel and what percentage is copper?

**Solution 2.32**

In each coin:

0.00740 moles of Ni

0.0886 moles of Cu

The masses of Ni & Cu are

$$0.0074\text{ moles}\times\frac{58.69\text{ gr}}{\text{mole}} = 0.434\text{ gr Ni}$$

$$0.0886\text{ moles}\times\frac{63.55\text{ gr}}{\text{mole}} = 5.63\text{ gr Cu}$$

Total mass = 0.434 + 5.63 = 6.06 gr

$$\%\text{ Ni} = \frac{0.434}{6.06}\times 100 = 7.16\%$$

$$\%\text{ Cu} = \frac{5.63}{6.06}\times 100 = 92.9\%$$

**Problem 2.33** Sterling silver contains 92.5 wt% silver and 7.5 wt% copper. Copper is added to silver to make the metal stronger and more durable. A small sterling silver spoon has a mass of 100 grams. Calculate the number of copper and silver atoms in the spoon.

**Solution 2.33**

Sterling silver: 92.5 wt % Ag + 7.5 wt % Cu

Mass of spoon = 100 gr

Mass of silver = $m_{Ag}$ = 92.5 gr ⇒ number of silver atoms $= \frac{6.02\times10^{23}\frac{\text{atoms}}{\text{mol}}}{107.9\frac{\text{gr}}{\text{mol}}} \times 92.5$ g

$= 5.2 \times 10^{23}$ atoms (Ag)

Mass of copper = $m_{Cu}$ = 7.5 gr · number of copper atoms $= \frac{6.02\times10^{23}\frac{\text{atoms}}{\text{mol}}}{63.55\frac{\text{gr}}{\text{mol}}} \times 7.5$ g

$= 7.1 \times 10^{22}$ atoms (Cu)

Total number of atoms in the spoon = $5.91 \times 10^{23}$ atoms

Note: it is smaller than $N_A$

**Problem 2.34** There are two naturally occurring isotopes for boron with mass numbers 10 (10.0129 amu) and 11 (11.0093 amu); The %abundances are 19.91 and 80.09 respectively. (a) Find the average atomic mass and (b) the relative atomic mass (or atomic weight) of boron. (c) Compare your value with that presented in the periodic table.

**Solution 2.34**

Boron isotopes : $^{10}$B – 10.0129 amu 19.91%

$^{11}$B – 11.0093 amu80.09%

Average atomic mass = $[(10.0129 \times 0.1991) + (11.0093 \times 0.8009)]$

$= [1.993 + 8.817]$ = 10.81 amu

Relative atomic mass = 10.81

Comparing with the value in the periodic table for B, we have a match.

**Problem 2.35** A monel alloy consists of 70 wt % Ni and 30 wt % Cu. What are the atom percentages of Ni and Cu in this alloy?

**Solution 2.35** Using a basis of 100 g of alloy, there are 70 g of Ni and 30 g of Cu. The number of gram-moles of each element is thus,

$$\text{No. of gram} - \text{moles of Cu} = \frac{30\ g}{63.54\ \text{g/mol}} = 0.472\ \text{mol}$$

$$\text{No. of gram} - \text{moles of Ni} = \frac{70\ g}{58.71\ \text{g/mol}} = \underline{1.192\ \text{mol}}$$

$$\text{Total gram-moles} = 1.664\, mol$$

The atomic percentages may then be calculated as,

$$\text{Atomic \% Cu} = \left[\frac{0.472\ \text{mol}}{1.664\ \text{mol}}\right](100\%) = 28.4\ at\%$$

$$\text{Atomic \% Ni} = \left[\frac{1.192\ \text{mol}}{1.664\ \text{mol}}\right](100\%) = 71.6\ at\%$$

**Problem 2.36** What is the chemical formula of an intermetallic compound that consists of 15.68 wt % Mg and 84.32 wt % Al?

**Solution 2.36** The chemical formula, $Mg_xAl_y$, may be determined based on the gram-mole fractions of magnesium and aluminum. Using a basis of 100 g of intermetallic compound,

$$\text{No. of gram} - \text{moles of Mg} = \frac{15.68\ g}{24.31\ \text{g/mol}} = 0.645\ \text{mol}$$

$$\text{No. of gram} - \text{moles of Al} = \frac{84.32\ g}{26.98\ \text{g/mol}} = \underline{3.125\ \text{mol}}$$

$$\text{Total gram-moles} = 3.770\, mol$$

$$x = \text{Gram-mole fraction of Mg} = \left[\frac{0.645\ \text{mol}}{3.770\ \text{mol}}\right] = 0.17$$

$$y = \text{Gram-mole fraction of Al} = \left[\frac{3.125\ \text{mol}}{3.770\ \text{mol}}\right] = 0.83$$

Thus we have $Mg_{0.17}Al_{0.83}$ or, multiplying by 6, **$MgAl_5$**.

**Problem 2.37** In order to raise the temperature of 100 grams of water from room temperature (20°C) to boiling temperature (100°C), an energy input of 33,440.0 J is required. If one uses a microwave oven (I, of radiation of 1.20 cm) to achieve this, how many photons of the microwave radiation are required?

**Solution 2.37**

33,440 joules of energy to raise water temp from 20°C to 100°C

- microwave = 1.20 cm = 0.012 m

$$E_{\text{mw}}\text{ photon} = \frac{h_c}{\lambda_{\text{mw photon}}} = \frac{(6.63\times10^{-34} J\cdot\text{sec})\left(3.00\times10^{8}\,\frac{m}{\text{sec}}\right)}{0.012\ \text{m}}$$

$E_{\text{mw}}$ photon = $1.66\times10^{-23}$ J

- Number of photons needed $= \frac{33{,}440\ \text{J}}{1.66\times10^{-23} J} = 2.0\times10^{27}$

Note: number is greater than $N_A$ by ~ 10,000 times.

**Problem 2.38** For problem 2.35, determine the number of photons to achieve the same increase in temperature if (a) ultraviolet (I = 1.0 x 10-8 m), visible (I = 5.0 x 10-7 m), and infrared (I = 1.0 x 10-4 m) lights were used. What important conclusions can you draw from this exercise?

**Solution 2.38**

Repeating problem 2.35 for UV and visible red lights

a) Ultraviolet
- $= 1\times10^{-8} m$

$$E_{\text{UV}}\text{ photon} = \frac{(6.63\times10^{-34} J\cdot\text{sec})\left(3.00\times10^{8}\,\frac{m}{\text{sec}}\right)}{1.0\times10^{-8}\ \text{m}}$$

$E_{\text{UV}}$ photon = $1.989\times10^{-17} J$

number of UV photons needed $= \frac{33{,}440\ \text{J}}{1.98\times10^{-17} J} = 1.68\times10^{21}$

b) Visible
- $= 5.0\times10^{-7} m$

$$E_{\text{vis}}\text{ photon} = \frac{(6.63\times10^{-34} J\cdot\text{sec})\left(3.00\times10^{8}\,\frac{m}{\text{sec}}\right)}{5.0\times10^{-7}\ \text{m}}$$

$E_{\text{vis}}$ photon = $4.0\times10^{-19} J$

number of vis photons needed $= \frac{33{,}440\ \text{J}}{4.0\times10^{-19} J} = 8.36\times10^{22}$

c) Infrared
- $= 1\times10^{-4} m$

$$E_{\text{IR}}\text{ photon} = \frac{(6.63\times10^{-34} J\cdot\text{sec})\left(3.00\times10^{8}\,\frac{m}{\text{sec}}\right)}{1.0\times10^{-4}\ \text{m}}$$

$E_{\text{IR}}$ photon = $2.0\times10^{-21} J$

number of IR photons needed $= \frac{33{,}440\ \text{J}}{2.0\times10^{-21} J} = 1.6\times10^{25}$

Note: $E_{\text{UV}} > E_{\text{vis}} > E_{\text{IR}} > E_{\text{mw}}$

**Problem 2.39** In order for the human eye to detect the visible light, its optical nerves must be exposed to a minimum energy of 2.0 x 10-17 J. (a) Calculate the number of photons of red light needed to achieve this (I = 700 nm). (b) Without any additional calculations, determine if you would need more or less photons of blue light to excite the optical nerves?

**Solution 2.39**

a) Energy for detection of light by optic nerve = $2.0 \times 10^{-17}$ J

number of photons of red light needed ( ▪ = 700 nm)

$$\text{Ered} = \frac{(6.63\times10^{-34} J\cdot\text{sec})(3.00\times10^{8}\,\frac{m}{\text{sec}})}{700\times10^{-9}\text{ m}} = 2.84\times10^{-19}J$$

number of red photons = $\frac{2.0\times10^{-17}\text{ J}}{2.84\times10^{-19}J} \approx 70.4$ photons
(very little energy is needed).

b) Blue light has a lower wavelength 450-495 nm. Thus, each photon of blue light will have more energy. You would need less photons to detect blue light.

**Problem 2.40** Represent the wave length of the following rays by comparing each to the length of a physical object (example: a ray with a wavelength of 1m (100 cm) would be approximately that of a baseball bat) : a) rays from a dental ray, b) rays in a microwave oven, c) rays in a sun lamp, d) rays in a heat lamp, and e) an FM radio wave.

**Solution 2.40**

Dental X-ray : ▪ $\sim 10^{-1}$ nm ( 100 times the diameter of an atom)

Micro wave : $\lambda\sim 1$ mm to $1\ m$ (thickness of 10 sheets of paper to the length of a baseball bat)

Heat lamp (Infrared) : ▪ $\sim 750$ nm to $1$ mm (a bacterial cell to thickness of 10 sheets of paper)

Sun lamp (UV) : ▪ $\sim 400$ nm to $10$ nm (size of a virus)

FM radio (radio wave) : ▪ $\sim 1$ m to $10\ m$ (baseball bat to a flagpole)

**Problem 2.41** For the rays in problem 2.38, without any calculations, rank them in increasing order of the energy of the radiation.

**Solution 2.41**

As wavelength increases, energy decreases

As wavelength decreases, energy increases

FM radio → microvave → heat lamp → sun lamp → X ray

(low energy) (high energy)

**Problem 2.42** In a commercial x-ray generator, a stable metal such as copper (Cu) or tungsten (W) is exposed to an intense beam of high-energy electrons. These electrons cause ionization events in the metal atoms. When the metal atoms regain their ground state they emit x-rays of characteristic energy and wavelength. For example, a "tungsten" atom struck by a high-energy electron may lose one of its K shell electrons. When this happens, another electron, probably from the tungsten L shell will "fall" into the vacant site in the K shell. If such a 2p → 1s transition occurs in tungsten, a tungsten $K_\alpha$ x-ray is emitted. A tungsten $K_\alpha$ x-ray has a wavelength $\lambda$ of 0.02138 nm. What is its energy? What is its frequency?

**Solution 2.42**

$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ \text{m/s})}{(0.02138\ \text{nm})(10^{-9}\ \text{m/nm})} = 9.30 \times 10^{-15}\ J$$

$$\nu = \frac{E}{h} = \frac{9.30 \times 10^{-15}\ J}{6.63 \times 10^{-34}\ J \cdot s} = 1.40 \times 10^{19}\ Hz$$

**Problem 2.43** What is the difference between equations 2.4 and 2.8?

**Solution 2.43.** Equation 2.4 is used strictly for a single electron atom such as hydrogen. Equation 2.8 is updated to include Zeff for multielectron atoms.

**Problem 2.44** Calculate the $Z_{eff}$ value for the following electrons: a) 2s electrons for Be, b) 3p electrons for Mn, and c) 4d electrons for Ag.

**Solution 2.44.**

a) 2s electron of Be

Zeff must be determined using the Slater's rule. The value of Z, number of protons in Be, is 4. To determine the value of S, the average number of electrons between the nucleus and the electron under consideration, we will first write the full electronic structure.

Be $\underline{1s^{2,}, 2s^{2}}$

The electron of interest is in the S orbital, n=2. All electrons within the n=2 energy level will contribute 0.35 to S. In n=2, we have one additional electron in the 2S orbital resulting in a contribution of 1 x 0.35 to the S value. The two electrons in the lower energy level, the 1s level, will contribute 2 x 0.85 to the S value. Thus, overall S is calculated as follows:

S = (1 x 0.35) + (2 x 0.85) = 2.05

2s 1s

Zeff is then calculated as follows:

$Z_{eff} = Z - S = 4 - 2.05 = 1.95$

b) 3p electron of Mn

Mn $\underline{1s^{2}, 2s^{2}, 2p^{6}, 3s^{2}, 3p^{6}}, 3d^{5}, 4s^{2}$

The electron of interest is in the P orbital, n=3. Five 5 electrons in the 3p orbital and two in the 3S orbital contribute 0.35 each (the 3d electrons have energy higher than 3p and won't contribute). There are 8 electrons in the n=2 level, each contributes 0.85. and finally, two electrons in the n=1 energy level, each contributes 1 to the the S value. Thus, overall S is calculated as follows:

S = (5 x 0.35) + (2 x 0.35) + (6 x 0.85) + (2 x 0.85) + (2 x 1) = 11.25

3p 3s 2p 2s 1s

Zeff is then calculated as follows:

$$Z_{eff} = Z - S = 25 - 11.25 = 13.75$$

c) 4d electron of Ag

Ag $\underline{1s^{2}, 2s^{2}, 2p^{6}, 3s^{2}, 3p^{6}, 3d^{10}, 4s^{2}, 4p^{6}, 4d^{9}}, 5s^{2}$

The electron of interest is in the d orbital, n=4. The remaining 8 electrons in 4d will contribute 0.35 each. All other electrons in lower levels, including 4p and 4s electrons will contribute 1. Thus, overall S is calculated as follows:

S = (8 x 0.35) + (6 x 1) + (2 x 1) + (10 x 1) + (6 x 1) + (2 x 1) + (8 x 1) + (2 x 1) = 33.6

4d 4p 4s 3d 3p 3s 2s+2p 1s

Zeff is then calculated as follows:

$$Z_{eff} = Z - S = 47 - 38.8 = 8.2$$

**Problem 2.45** Calculate the energy levels associated with all Li and He electrons. Compare the values and draw a conclusion.

**Solution 2.45**.

He:

1s electrons: S = (1 x 0.3); Zeff = 2- 0.3 = 1.7; E= -13.6 ($1.7^2/1^2$) = -39.30ev

Li:

1s: S = (1 x 0.3); Zeff = 3- 0.30 = 2.7; E= -13.6 ($2.7^2/1^2$) = -99.14 ev

2s: S= (2 x 0.85); Zeff = 3-1.7 = 1.3; E= -13.6 ($1.3^2/2^2$) = -5.75 ev

Conclusions:

1s electrons in He are much easier to separate from the nucleus than the 1s electrons in Li

2s electrons in Li are easier to remove than the 1s electrons in He

2s electrons in Li are significantly easier to remove thank the 1s electrons in Li

These effects are all due to nuclear charge effects due to existence of multielectrons.

**Problem 2.46** A hydrogen atom exists with its electron in the *n* = 4 state. The electron undergoes a transition to the *n* = 3 state. Calculate (*a*) the energy of the photon emitted, (*b*) its frequency, and (*c*) its wavelength in nanometers (nm).

**Solution 2.46**

(a) Photon energy emitted is:

$$\Delta E = \Delta\left[\frac{-13.6}{n^2}\right] = \left[\frac{-13.6}{4^2}\right] - \left[\frac{-13.6}{3^2}\right] = 0.66\ eV = 1.06 \times 10^{-19}\ J$$

(b) Photon frequency is found as:

$$\nu = \frac{\Delta E}{h} = \frac{1.06 \times 10^{-19}\ J}{6.63 \times 10^{-34}\ J \cdot s} = 1.6 \times 10^{14}\ Hz$$

(c) The wavelength is given as:

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63\times10^{-34}\ J\cdot s)(3.00\times10^{8}\ \text{m/s})}{(1.06\times10^{-19}\ J)(10^{-9}\ \text{m/nm})} = 1876\ nm$$

**Problem 2.47** A hydrogen atom exists with its electron in the *n* = 6 state. The electron undergoes a transition to the *n* = 2 state. Calculate (*a*) the energy of the photon emitted, (*b*) its frequency, and (*c*) its wavelength in nanometers.

**Solution 2.47**

(a) Photon energy emitted is:

$$\Delta E = \Delta\left[\frac{-13.6}{n^2}\right] = \left[\frac{-13.6}{6^2}\right] - \left[\frac{-13.6}{2^2}\right] = 3.02\ eV = 4.84 \times 10^{-19}\ J$$

(b) Photon frequency is found as:

$$\nu = \frac{\Delta E}{h} = \frac{4.84 \times 10^{-19}\ J}{6.63 \times 10^{-34}\ J \cdot s} = 7.3 \times 10^{14}\ Hz$$

(c) The wavelength is given as:

$$\lambda = \frac{hc}{\Delta E} = \frac{(6.63 \times 10^{-34}\ J \cdot s)(3.00 \times 10^{8}\ \text{m/s})}{(4.84 \times 10^{-19}\ J)(10^{-9}\ \text{m/nm})} = 410\ nm$$

**Problem 2.48** Using the information given in Examples 2.4 and 2.5 determine the uncertainty associated with the electron's position if the uncertainty in determining its velocity is 1%. Compare the calculated uncertainty in the position with the estimated diameter of the atom. What is your conclusion?

**Solution 2.48**

Speed of electron = 16.67 % of speed of light

= $0.166 \times 3 \times 10^{8} = 50.0 \times 10^{6}$ m/sec

1% uncertainty in measurement of speed = $0.01 \times 50.0 \times 10^6 = 500 \times 10^3$

Uncertainty in position $\Delta x \geq \frac{h}{4\pi \, m \, \Delta u} = \frac{6.62\times10^{-34} \frac{\text{kg}.m^2}{s}}{4\,\pi\,(9.11 \times 10^{-31}\text{ kgr})\times(500 \times 10^3\text{ kgr})}$

$\Delta x \geq 1.15 \times 10^{-10}$ m $\approx$ 0.115 nm

Note: diameter of atom is $\sim$ 0.1 nm. Uncertainty in the position of electron will be close to the size of the atom.

**Problem 2.49** Repeat problem 2.43 to determine the uncertainty associated with the electron's position if the uncertainty in determining its velocity is 2%. Compare the calculated uncertainty in the position with that of problem 2.37. What is your conclusion?

**Solution 2.49**

2% uncertainty associated with speed

$= 0.02 \times 50.0 \times 10^6$ m/s $= 1000 \times 10^3$ m/s

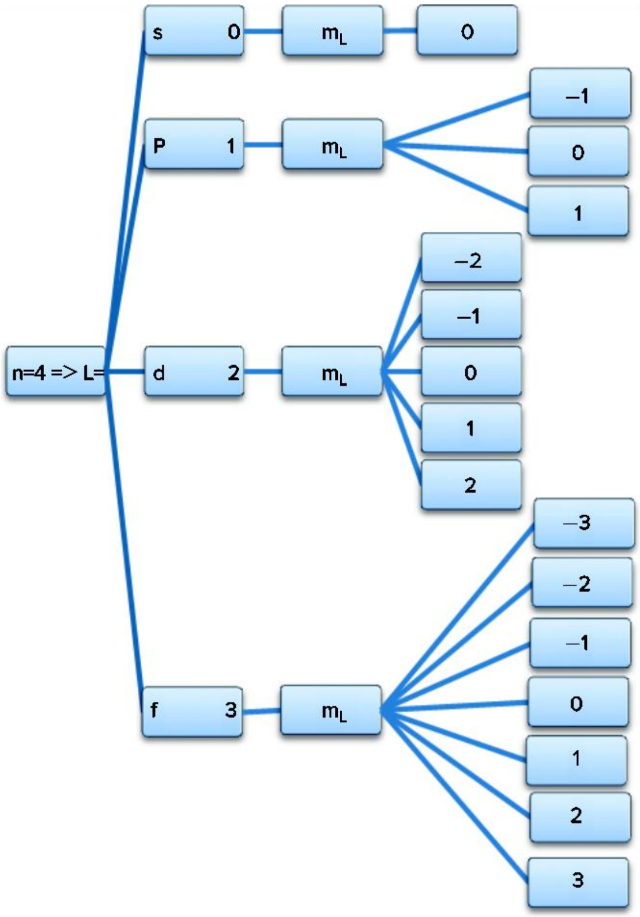
- The corresponding uncertainty in position will be $\Delta x \geq \frac{6.62\times10^{-34} \frac{\text{kg}.m^2}{s}}{4\,\pi\,(9.11 \times 10^{-31}\text{ kgr})(1000 \times10^3 m/s)}$

$\Delta x \geq 5.268 \times 10^{-11}$ m $\approx$ 0.0526 nm

Note: the uncertainty in position decreases to the radius of the atom as the uncertainty in the speed of electron increases to 2%.

**Problem 2.50** For the principal quantum number, n, of value 4, determine all other possible quantum numbers for l and $m_l$.

**Solution 2.50**



**Problem 2.51** For each pair of n and l given below, give the sublevel name, possible values of $m_l$, and the corresponding # of orbitals.

(a) n=1, l=0; (b) n=2, l=1; (c) n=3, l=2; (d) n=4, l=3

## Solution 2.51

| | | | | Sublevel name | $m_l$ | number of orbitals |
|---|---|---|---|---|---|---|
| a) | n=1, | l=0 | ▪ | 1s | 0 | 1 |
| b) | n=2, | l=1 | ▪ | 2p | -1,0,+1 | 3 |
| c) | n=3, | l=2 | ▪ | 3d | -2,-1,0,+1,+2 | 5 |
| d) | n=4, | l=3 | ▪ | 4f | -3,-2,-1,0,+1,+2,+3 | 7 |

**Problem 2.52** Determine if the following combinations of quantum numbers are acceptable.

(a) n=3, l=0, $m_l$=+1; (b) n=6, l=2, $m_l$=-3; (c) n=3, l=3, $m_l$=-1; (d) n=2, l=1, $m_l$=+1

## Solution 2.52

a) n=3, l=0 $m_l = +1$ : not possible, $m_l = 0$

b) n=6, l=2 $m_l = -3$ : not possible, $-2 < m_l < +2$

c) n=3, l=3 $m_l = -1$: not possible, l = 0,1,2 for n=3

d) n=2 l=1 $m_l = +1$ : possible

**Problem 2.53** In each row (a through d) there is only one piece of information that is wrong, highlight the information that is wrong (explain why).

| | **n** | **l** | **$m_l$** | **Name** |
|---|---|---|---|---|
| **(a)** | **3** | **0** | **1** | **3s** |
| **(b)** | **2** | **1** | **-1** | **2s** |
| **(c)** | **3** | **1** | **+2** | **3d** |
| **(d)** | **3** | **3** | **0** | **4f** |

## Solution 2.53

(Note: only one entry is wrong)

n l $m_l$ name

a) 3 0 $\boxed{1}$ 3s : $m_l \rightarrow$ 0 since l=0

| | | | | | | |
|---|---|---|---|---|---|---|
| b) | 2 | 1 | -1 | 2s | : | name → 2p since l=1 |
| c) | 3 | $\boxed{1}$ | +2 | 3d | : | l → 2 since $m_l$ > +2 and name is 3d |
| d) | $\boxed{3}$ | 3 | 0 | 4f | : | n → 4 since l=3 and name is 4f |

**Problem 2.54** Determine the four quantum numbers for the 3rd, 15th, and 17th electrons of the Cl atom.

**Solution 2.54**

Cl has 17 electrons

3rd electron of Cl atom: $n = 2$, $l = 0$, s, $m_l = 0$ $m_s = +1/2$

15th electron of Cl atom: $n = 3$, $l = 1$, p, $m_l = +1$ $m_s = +1/2$

17th electron of Cl atom: $n = 3$, $l = 1$,p, $m_l = 0$ $m_s = -1/2$

**Problem 2.55** Determine the electron configuration and group number of the atom in the ground state based on the given partial (valence level) orbital diagram. Identify the element.

**Solution 2.55**

The outer electron structure is $4s^2p^1$.

This means 28 subvalent electrons.

Total number of electrons will be 31.

The element is Ga (Gallium)

(Alternatively, the element is in the 3rd column since it has 3 outer electrons and fourth period since $4s^24p^7$ → Ga)

**Problem 2.56** Write the electron configurations of the following elements by using *spdf* notion:

(*a*) yttrium, (*b*) hafnium, (*c*) samarium, (*d*) rhenium.

**Solution 2.56**

(a) Y ($Z$ = 39): [Kr] $4d^15s$ (c) Sm (Z = 62): [Xe] $4f^66s^2$

(b) Hf ($Z = 72$): [Xe] $4f^{14}5d^{2}6s^{2}$ (d) Re (Z = 75): [Xe] $4f^{14}5d^{5}6s^{2}$

**Problem 2.57** Write the electron configuration of the following ions by using *spdf* notation:

(*a*) $Cr^{2+}$, $Cr^{3+}$, $Cr^{6+}$; (*b*) $Mo^{3+}$, $Mo^{4+}$, $Mo^{6+}$; (*c*) $Se^{4+}$, $Se^{6+}$, $Se^{2-}$.

**Solution 2.57**

| | | | | | | | | |
|---|---|---|---|---|---|---|---|---|
| (a) | Cr | [Ar] $3d^{5}\,4s^{1}$ | (b) | Mo | [Kr] $4d^{5}\,5s^{1}$ | (c) | Se | [Ar] $3d^{10}\,4s^{2}4p^{4}$ |
| | $Cr^{2+}$ | [Ar] $3d^{4}$ | | $Mo^{3+}$ | [Kr] $4d^{3}$ | | $Se^{4+}$ | [Ar] $3d^{10}\,4s^{2}$ |
| | $Cr^{3+}$ | [Ar] $3d^{3}$ | | $Mo^{4+}$ | [Kr] $4d^{2}$ | | $Se^{6+}$ | [Ar] $3d^{10}$ |
| | $Cr^{6+}$ | [Ar] | | $Mo^{6+}$ | [Kr] $4d^{3}$ | | $Se^{2-}$ | [Ar] $3d^{10}\,4s^{2}4p^{6}$ |

**Problem 2.58** Rank the following atoms in (a) increasing atomic size and (b) decreasing first ionization energy, IE1. Use only the periodic table to answer the questions. Check your answer using Figures 2.10 and 2.11.

i) K, Ca, Ga
ii) Ca, Sr, Ba
iii) I, Xe, Cs

**Solution 2.58**

i)

a) Atomic size (increasing)

As we move to the left in a period, the atomic size increases

$$\left.\begin{array}{l}\text{Ga : period 4 group 3}\\ \text{Ca: period 4 group 2}\\ \text{K: period 4 group 1}\end{array}\right\} \cdot \quad \text{Ga, Ca, K (increasing order)}$$

b) Ionization energy (decreasing)

As we move to the left in a period, the ionization energy increases.

Ga, Ca, K (decreasing order)

ii)

a) Atomic size (increasing)

Ca : period 4 group 2

Sr: period 5 group 2

Ba: period 6 group 2

As we move down in a group, atomic size increases

∴ Ca, Sr, Ba (increasing size)

b) Ionization energy (decreasing)

Ionization energy will have the opposite trend as that of atomic size.

(as we move down in a group, ionization energy decreases)

Thus, Ca, Sr, Ba (decreasing)

iii)

a) Atomic size (increasing)

I : period 5 group 7

Xe: period 5 group 8

Cs: period 6 group 1

In period 5, Xe (group 8) will have a smaller radius than I (group 7).

Cs in period 6 will be larger than both.

Cs, I, Xe (increasing size)

b) Ionization energy (decreasing)

In period 5, I will have a lower ionization energy than Xe.

Cs, in period 6, will have the lowest ionization energy.

Xe, I, Cs (decreasing)

**Problem 2.59** Rank the following atoms in (a) increasing atomic size and (b) decreasing first ionization energy, IE1. Use only the periodic table to answer the questions. Check your answer using Figures 2.10 and 2.11.

i) Ar, Li, F, O, Cs C

ii) Sr, H, Ba, He, Mg, Cs

**Solution 2.59**

a) Increasing atomic size

i)

Ar : period 3, group 8
Li : period 2, group 1
F : period 2, group 7
O : period 2, group 6
Cs : period 6, group 1
C : period 2, group 4

Period 2 elements will be smaller than period 3, and both will be smaller than period 6.

period 2 Li, F, O, C
1st organization period 3 Ar ↓ Increasing size
Period 6 Cs

Within each period, the size decreases as you move to the right

2nd organization ↓ Li F
C, O

Increasing size O, C

F, Li

Argon (Ar), the last element in period 3, will be smaller than Li but larger than F

F

$$\left.\begin{array}{l}\mathrm{O}\\ \mathrm{C}\\ \mathrm{Li}\end{array}\right\}\ \mathrm{Ar}$$

Cs

Cs, will be the largest since it is in period 6.

**Problem 2.60** The first ionization energies of two atoms with electronic configurations (a) $1S^22s^22p^6$ and (b) $1s^22s^22p^63s^1$ are given to be 2080 kJ/mol and 496 kJ/mol. Determine which IE1 belongs to which electronic structure and justify your answer.

**Solution 2.60**

| | |
|---|---|
| $1s^22s^22p^6$ | This atom belongs to group VIII. Its outer electron structure is complete. It will be difficult to ionize this atom. Thus, it must have a high $IE_1$. |
| $1s^22s^22p^63s^1$ | This atom belongs to group I, period III. There is a single electron in n=3.<br>It will be easier to ionize, so it must have a lower $IE_1$. |

**Problem 2.61** The first ionization energies of three atoms with electronic configurations (a) $[He]2s^2$, (b) $[Ne]3s^1$, and (c) $[Ar]4s^1$ and (d) $[He]2S^1$ are given to be 496 kJ/mol, 419 kJ/mol, 520 kJ/mol, and 899 kJ/mol. Determine which IE1 belongs to which electronic structure and explain your answer.

**Solution 2.61**

$[He]\ 2s^2$ $[Ne]3s^1$ $[Ar]4s^1$ $[He]2s^1$ $[He]2s^1$ $[Ne]3s^1$ and $[Ar]4s^1$ all belong to group I. The atom that is located in the lowest period will have the lowest $IE_1$. Thus,

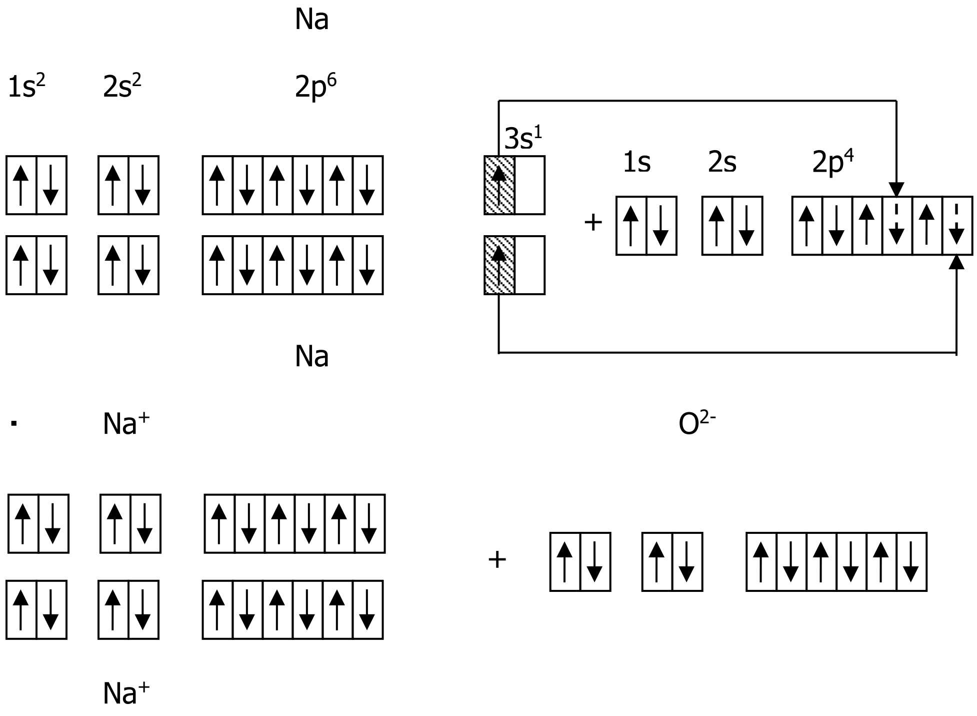
$IE_1$ $[Ar]4s^1$ = 419 KJ/mol; $IE_1$ $[Ne]3s^1$ = 496 KJ/mol; $IE_1$ $[He]2s^1$ = 520 KJ/mol

$[He]2s^2$ belongs to the second group in period II. Thus, its $IE_1$ must be higher than $[HE]2s^1$. And so, $IE_1$ $[He]2s^1$ = 899 KJ/mol

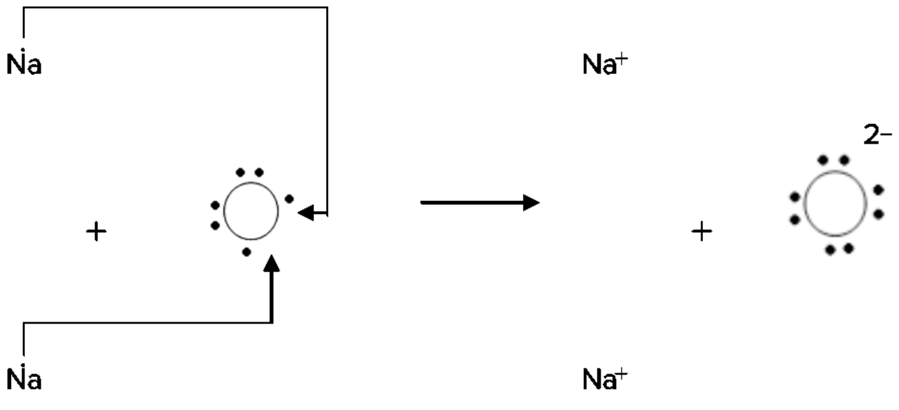
**Problem 2.62** Similar to Figure 2.15 use (a) orbital diagrams and (b) Lewis symbols to explain the formation of $Na^+$ and $O^{2-}$ ions and the corresponding bonding. What is the formula of the compound?

**Solution 2.62**

a) Orbital diagram



b) Lewis symbols



Chemical formula = $Na_2O$