

## CHAPTER 2

### Knowledge and Comprehension Problems:

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

**Answer 2.1:** Joseph J. Thompson using cathode ray tube experiments concluded that atoms in all matters 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  $5.60 \times 10^{-19} \text{ g/C}$  where *Coulomb*, C, is the unit of electrical charge. Robert Millikan<sup>9</sup>, in his oil-drop experiments, determined the fundamental quantity of charge or the charge of an electron (regardless of the source) to be  $1.60 \times 10^{-19} \text{ C}$ . 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  $8.96 \times 10^{-28} \text{ g}$ .

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

**Answer 2.2:** (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.

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

**Answer 2.3:** 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  $\alpha$  (alpha),  $\beta$  (Beta), and  $\gamma$  (gamma) rays. It was also shown that  $\alpha$  and  $\beta$  particles have both charge and mass while  $\gamma$  particles have no detectable mass or charge.

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

**Answer 2.4:** Using the information in Table 2.1, one can summarize that protons and neutrons significantly higher mass and basically constitute the total mass of the atom (the mass of electron is minimal in comparison). On the other hand, charge of the atom comes equally from its electrons (negative) and protons (positive). Neutron is not charged.

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

2.5 How was the existence of protons first verified? Discuss the characteristics of protons.

**Answer 2.5:** Ernest Rutherford, bombarded a very thin foil of gold with positively charged  $\alpha$  particles. He noticed that many of the  $\alpha$  particles pass through the foil without deflection, some 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  $1.672 \times 10^{-24}$  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.

2.6 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 2.6:** 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 atom of iron has an atomic mass of 16.00 amu.

2.7 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 2.7:**

- 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  $1/12^{\text{th}}$  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.

2.8 One mole of titanium atoms has a mass of 47.88 grams, without any calculations determine the mass in amu of one iron atom.

**Answer 2.8:** One mole of titanium corresponds to the number of atoms needed to create a mass in units of grams (47.88 grams) numerically equal to the atomic mass in amu of the substance under consideration. Thus one atom of titanium has an atomic mass of 47.88 amu.

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

**Answer 2.9:** Light is in the form of electromagnetic radiation. Energy is released and transmitted in the form electromagnetic waves.

**2.10** (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 2.10:**

- (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 ( $E = h\nu$ ).

**2.11** Explain the law of chemical periodicity.

**Answer 2.11:** State that the properties of elements are functions of their atomic number in a periodic manner.

**2.12** Describe the Uncertainty Principle. How does this principle contradict Bohr’s model of the atom?

**Answer 2.12:** 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 position and momentum of an electron simultaneously.

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

**Answer 2.13:**

- 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).

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

**Answer 2.14:** 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.

**2.15** 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, h) nonmetals, i) metalloids, and j) electronegativity.

**Answer 2.15:**

- 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 (IE<sub>1</sub>)- 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.

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

**Answer 2.16:**

- (a) The higher the charge of the nucleus (more protons), the higher is 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.

**2.17** Explain the Pauli's exclusion principle.

**Answer 2.17:** 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.

**2.18** What can the values of spin quantum number be? Explain the meaning of these numbers.

**Answer 2.18:** The spin quantum number can take on either  $+1/2$  or  $-1/2$ . The electron can have only two directions of spin and no other position is allowed. According to Pauli's exclusion principle, no more than two electrons can occupy the same orbital of an atom, and the two electrons must have opposite spins, i.e.  $m_s=1/2$  for one, and  $m_s=-1/2$  for other.

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

**Answer 2.19:** 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  $\text{Cl}^-$  anions can pack around one  $\text{Na}^+$  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 can not be greater than four.

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

**Answer 2.20:**

**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** - is 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.

2.21 Describe a) Hess law, b) lattice energy and c) heat of formation.

**Answer 2.21:**

- (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:  $\Delta H^0 = \Delta H^1 + \Delta H^2 + \Delta H^3 + \Delta H^4 + \Delta H^5$ .
- (b) The total heat,  $\Delta H^0$ , is called the heat of formation.
- (c) Lattice energy,  $\Delta H^5$ , is the energy released when gaseous ions form solid ions due to electrostatic attraction forces.

2.22 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 2.22:**

- a) shared pair – the pair of electron 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).

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 –  $\Delta H^1$

Stage 2: Nonmetal molecules to nonmetal atoms – requires energy –  $\Delta H^2$

Stage 3: Metal atoms, removing outer electrons – requires energy –  $\Delta H^3$

Stage 4: Nonmetal atoms, adding outer electrons – produces energy –  $\Delta H^4$

Stage 5: Formation of ionic solid from gaseous ions – produces energy –  $\Delta H^5$

2.24 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 2.24:**

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), electrical insulators (in the solid state), and high melting temperature. Examples, 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), low thermal conductivity, and high melting temperature. Examples, quartz and diamond.

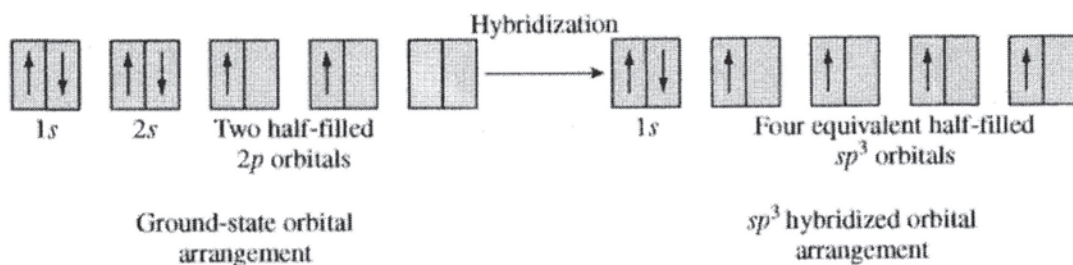
c) metallic solids – in a pure state, generally more malleable (soft and deformable) and less stiff than ionic or covalent networked materials. Strength can be increased through alloying. Highly conductive (both heat and electricity). Example, copper and aluminum.

2.25 What are primary bonds? What is the driving force for formation of such bonds? Give examples of materials in which such bonds exist.

**Answer 2.25:** The primary bonds are stronger than secondary bonds. Three different types of primary bonds are found in solids — ionic, covalent and metallic. The driving force arises from the tendency of the atoms to assume stable electron structure by completely filling the outermost electron shell.

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.



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

**Answer 2.27:**

- a) dipole moment – a moment produced to temporary or permanent separation of positive and negative charge centers in an atom or a molecule (moment =  $q \cdot d$ ).
- 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, 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.



**Application and Analysis Problems:**

\*2.28 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?

**Answer 2.28:**

2.28

Surface area of the moon =  $\frac{4}{3} \pi R^3$  : R = radius of moon

$$A_{s, \text{moon}} = \frac{4}{3} \pi \left( \frac{3.476}{2} \times 10^6 \text{ m} \right)^2 = 1.26 \times 10^{13} \text{ m}^2$$

Cross-sectional area of soccerball =  $\pi r^2$

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

$$\begin{aligned} \# \text{ of soccer balls} \\ \text{required to cover} \\ \text{the surface of} \\ \text{the moon} \end{aligned} = \frac{1.26 \times 10^{13} \text{ m}^2}{0.0611} = 2.07 \times 10^{14} \text{ balls}$$

Avogadro's number is  $\approx 3$  billion times larger.

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.

2.30 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.

Answer 2.30:

2.30

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

mass of spoon = 100 gr

mass of silver =  $M_{Ag}$

$$= 92.5 \text{ gr} \Rightarrow \# \text{ of silver atoms} = \frac{6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}}{107.9 \frac{\text{gr}}{\text{mol}}} \times 92.5 \text{ g}$$
$$= 5.2 \times 10^{23} \text{ atoms (Ag)}$$

mass of copper =  $M_{Cu}$

$$= 7.5 \text{ gr} \Rightarrow \# \text{ of atoms of copper} = \frac{6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}}{63.55 \frac{\text{gr}}{\text{mol}}} \times 7.5 \text{ g}$$
$$= 7.1 \times 10^{22} \text{ atoms (Cu)}$$

Total # of atoms in the spoon =  $5.93 \times 10^{23}$  atoms

note : it is smaller than  $N_A$ .

\*2.31 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. Compare your value with that presented in the periodic table.

Answer 2.31:

2.31

Boron isotopes :

$^{10}\text{B}$	- 10.0129 amu	19.91%
$^{11}\text{B}$	- 11.0093 amu	80.09%

$$\begin{aligned}\text{Average atomic mass} &= [(10.0129 \times 0.1991) + (11.0093 \times 0.8009)] \\ &= [1.993 + 8.817] = 10.81 \text{ amu}\end{aligned}$$

Relative atomic mass = 10.81 gr

note: Relative atomic mass will have the same numerical value as average atomic mass but with units of grams.

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

\*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? What percentage of the mass of a quarter is nickel and what percentage is copper?

Answer 2.32:

2.32

In each coin:

0.00740 moles of Ni

0.0886 moles of Cu

The masses of Ni & Cu are

$$0.0074 \text{ moles} \times \frac{6.02 \times 10^{23}}{\text{mole}} \times \frac{63.55 \text{ gr}}{\text{mole}} = 0.470 \text{ gr Ni}$$

$$0.0886 \text{ moles} \times \frac{58.96 \text{ gr}}{\text{mole}} = 5.22 \text{ gr Cu}$$

$$\text{Total mass} = 0.470 + 5.22 = 5.69 \text{ gr}$$

$$\% \text{ Ni} = 0. \frac{470}{5.69} \times 100 = 8\% \quad \% \text{ Cu} = 5. \frac{22}{5.69} \times 100 = 92\%$$

$$0.0074 \text{ moles} \times \frac{6.02 \times 10^{23}}{\text{mole}} \times \frac{63.55 \text{ gr}}{\text{mole}} = 0.470 \text{ gr Ni}$$

$$0.0886 \text{ moles} \times \frac{58.96 \text{ gr}}{\text{mole}} = 5.22 \text{ gr Cu} \quad \% \text{ Ni} = 0. \frac{470}{5.69} \times 100 = 8\%$$

$$\% \text{ Cu} = 5. \frac{22}{5.69} \times 100 = 92\%$$

2.33 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 ( $\lambda$ , of radiation of 1.20 cm) to achieve this, how many photons of the microwave radiation are required?

Answer 2.33:

2.33

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

$$\lambda_{\text{microwave}} = 1.20 \text{ cm} = 0.012 \text{ m}$$

$$E_{\text{mw photon}} = \frac{hc}{\lambda_{\text{mw photon}}} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/sec})}{0.012 \text{ m}}$$

$$E_{\text{mw photon}} = 1.66 \times 10^{-23} \text{ J}$$

$$\Rightarrow \# \text{ of photons needed} = \frac{33,440 \text{ J}}{1.66 \times 10^{-23} \text{ J}} = 2.0 \times 10^{27}$$

note: # is greater than  $N_A$  by  $\sim 10,000$  times.

2.34 For problem 2.33, determine the number of photons to achieve the same increase in temperature if (a) ultraviolet ( $\lambda = 1.0 \times 10^{-8}$  m), visible ( $\lambda = 5.0 \times 10^{-7}$  m), and infrared ( $\lambda = 1.0 \times 10^{-4}$  m) lights were used. What important conclusions can you draw from this exercise?

Answer 2.34:

2.34

Repeating prob 2.33 for UV and visible red lights

a) Ultraviolet  
 $\lambda = 1.0 \times 10^{-8}$  m

$$E_{uv \text{ photon}} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{1.0 \times 10^{-8} \text{ m}}$$

$$E_{uv \text{ photon}} = 2.0 \times 10^{-17} \text{ J}$$

$$\# \text{ of UV photons} = \frac{33,440 \text{ J}}{1.98 \times 10^{-17} \text{ J}} = 1.6 \times 10^{21}$$

b) Visible  
 $\lambda = 5.0 \times 10^{-7}$  m

$$E_{vis \text{ photon}} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{5.0 \times 10^{-7} \text{ m}}$$

$$E_{vis \text{ photon}} = 4.0 \times 10^{-19} \text{ J}$$

$$\# \text{ of vis photons} = \frac{33,440 \text{ J}}{4.0 \times 10^{-19} \text{ J}} = 8.4 \times 10^{22}$$

c) Infrared  
 $\lambda = 1.0 \times 10^{-4}$  m

$$E_{IR \text{ photon}} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{1.0 \times 10^{-4} \text{ m}}$$

$$E_{IR \text{ photon}} = 2.0 \times 10^{-21} \text{ J}$$

$$\# \text{ of IR photons} = \frac{33,440 \text{ J}}{2.0 \times 10^{-21} \text{ J}} = 1.6 \times 10^{25}$$

note  $E_{uv} > E_{vis} > E_{IR} > E_{mw}$

**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?

**Answer 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,

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

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

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

The atomic percentages may then be calculated as,

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

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

**\*2.36** What is the chemical formula of an intermetallic compound that consists of 37 wt % Ti and 63 wt % Al? Use a basis of 100 g of intermetallic compound.

**Answer 2.36:** The chemical formula,  $\text{Ti}_x\text{Al}_y$  may be determined based on the gram-mole fractions of titanium and aluminum.

$$\text{No. of gram - moles of Ti} = \frac{37 \text{ g}}{47.88 \text{ g/mol}} = 0.773 \text{ mol}$$

$$\text{No. of gram - moles of Al} = \frac{63 \text{ g}}{26.98 \text{ g/mol}} = \underline{2.335 \text{ mol}}$$

$$\text{Total gram - moles} = \underline{3.108 \text{ mol}}$$

$$x = \text{Gram - mole fraction of Ti} = \left[ \frac{0.773 \text{ mol}}{3.108 \text{ mol}} \right] = \underline{0.25}$$

$$y = \text{Gram - mole fraction of Al} = \left[ \frac{2.335 \text{ mol}}{3.108 \text{ mol}} \right] = \underline{0.75}$$

Thus we have  $\text{Ti}_{0.25}\text{Al}_{0.75}$  or, multiplying by 3,  $\text{TiAl}_3$ .

2.37 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 1 m (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.

Answer 2.37:

2.37

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

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

Heat lamp :  $\lambda \sim 750 \text{ nm to } 1 \text{ mm}$   
(infrared) (a bacterial cell to thickness of 10 sheets of paper)

Sun lamp :  $\lambda \sim \left\{ \begin{array}{l} 400 \text{ nm to } 10 \text{ nm} \\ \end{array} \right\}$  (size of a virus)  
(UV)

FM radio :  $\lambda \sim 1 \text{ m to } 10 \text{ m}$  (Baseball bat to a flag pole)  
(radio wave)



2.38 For the rays in problem 2.37, without any calculations, rank them in increasing order of the energy of the radiation.

Answer 2.38:

2.38 As wave length increases, energy decreases  
As wave length decreases, energy increases

FM radio → microwave → heat lamp → Sun lamp  
(Low energy)  
→ X-ray  
(high energy)

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

Answer 2.39:

2.39

a) Energy for detection of light by optic nerve =  $2.0 \times 10^{-17}$  J  
# of photons of red light needed ( $\lambda = 700$  nm)

$$E_{\text{red}} = \frac{(6.63 \times 10^{-34} \text{ J}\cdot\text{s})(3.00 \times 10^8 \text{ m/s})}{700 \times 10^{-9} \text{ m}} = 2.844 \times 10^{-19} \text{ J}$$

$$\# \text{ of red photons} = \frac{2.0 \times 10^{-17} \text{ J}}{2.84 \times 10^{-19} \text{ J}} \approx 72 \text{ photons.}$$

(very little energy is needed (72 photons)).

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.

**2.40** 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).

**Answer 2.40:**

(a) Photon energy emitted is:

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

(b) Photon frequency is found as:

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

(c) The wavelength is given as:

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

**\*2.41** 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.

**Answer 2.41:**

(a) Photon energy emitted is:

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

(b) Photon frequency is found as:

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

(c) The wavelength is given as:

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

**2.42** In a commercial x-ray generator, a stable metal such as Cu or 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 M atom struck by a high-energy electron may lose one of its K shell electrons. When this happens, another electron, probably from the W L shell will “fall” into the vacant site in the K shell. If such a  $2p \rightarrow 1s$  transition in W occurs, a W  $K_{\alpha}$  x-ray is emitted. A W  $K_{\alpha}$  x-ray has a wavelength  $\lambda$  of 0.2138 nm. What is its energy? What is its frequency?

**Answer 2.42:**

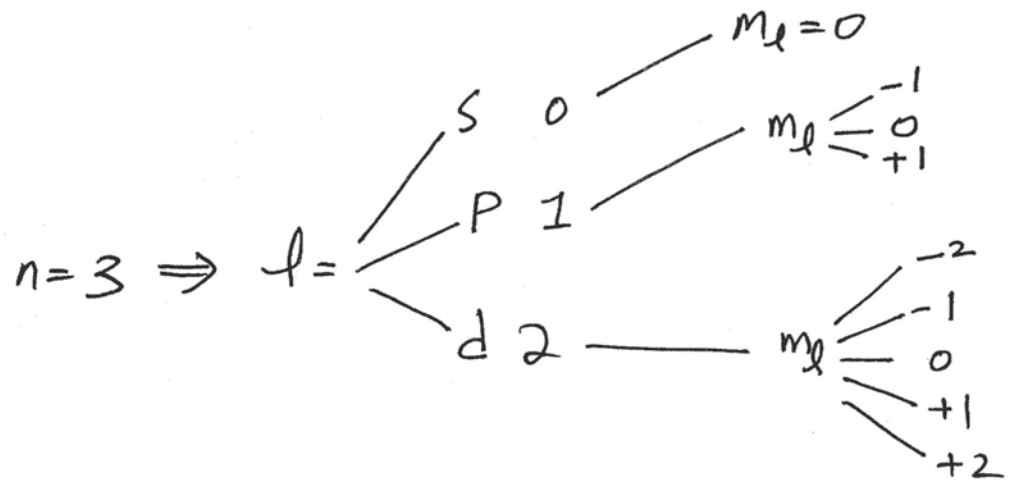
$$E = \frac{hc}{\lambda} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(0.2138 \text{ nm})(10^{-9} \text{ m/nm})} = 9.30 \times 10^{-15} \text{ J}$$

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

\*2.43 If the value of the principal quantum number,  $n$ , is 3, write down all possible quantum numbers for  $l$  and  $m_l$ .

Answer 2.43:

2.43



2.44 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?

Answer 2.44:

2.44

$$\begin{aligned}\text{speed of electron} &= 16.67\% \text{ of speed of light} \\ &= 0.1667 \times 3.0 \times 10^8 = 50.0 \times 10^6 \text{ m/s}\end{aligned}$$

$$\begin{aligned}1\% \text{ uncertainty in measurement of speed} \\ &= 0.01 \times 50.0 \times 10^6 = 500 \times 10^3\end{aligned}$$

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

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

note: diameter of atom is  $\approx 0.1 \text{ nm}$ .

Uncertainty in the position of electron will be close to the size of the atom.

2.45 Repeat problem 2.44 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 2.44. What is your conclusion?

Answer 2.45:

$$\begin{array}{l} \underline{2.45} \quad 2\% \text{ uncertainty associated with speed} \\ = 0.02 \times 50.0 \times 10^6 \text{ m/s} = 1000 \times 10^3 \text{ m/s} \end{array}$$

⇒ the corresponding uncertainty in position will be

$$\Delta x \gg \frac{6.62 \times 10^{-34} \frac{\text{kg} \cdot \text{m}^2}{\text{s}}}{4\pi (9.11 \times 10^{-31} \text{ kg}) (1000 \times 10^3 \text{ m/s})}$$

$$\Delta x \gg 5.78 \times 10^{-11} \text{ m} \approx 0.0578 \text{ nm}$$

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

2.46 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)  $n=2, l=1, m_l=+1$

Answer 2.46:

2.46

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)



2.47 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 <sub>l</sub>	Name
(a)	3	0	1	3s
(b)	2	1	-1	2s
(c)	3	1	+2	3d
(d)	3	3	0	4f

Answer 2.47:

2-47

(note: only one entry is wrong)

	n	l	m <sub>l</sub>	Name	
a)	3	0	1	3s	: m <sub>l</sub> → 0 since l = 0
b)	2	1	-1	2s	: name → 2p since l = 1
c)	3	1	+2	3d	: l → 2 since m <sub>l</sub> > +2 and name is 3d
d)	3	3	0	4f	: n → 4 since l = 3 and name is 4f

2.48 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)  $n = 2, l = 1$

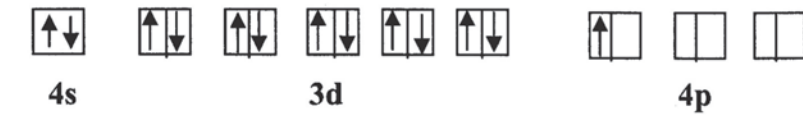
(c)  $n = 3, l = 2$

(d)  $n = 4, l = 3$

Answer 2.48:

<u>2-48</u>	sublevel name	$m_l$	# of orbitals
a) $n = 1, l = 0 \Rightarrow 1s$		0	1
b) $n = 2, l = 1 \Rightarrow 2p$		-1, 0, +1	3
c) $n = 3, l = 2 \Rightarrow 3d$		-2, -1, 0, +1, +2	5
d) $n = 4, l = 3 \Rightarrow 4f$		-3, -2, -1, 0, +1, +2, +3	7

2.49 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.



Answer 2.49:

2.49 The outer electron structure is  $4s^2 4p^1$   
This means 28 subvalent electrons.  
Total # of electrons will be 31.  
The element is Ga (gallium)  
(alternatively the element is in the 3<sup>rd</sup> column  
since 3 outer electrons and fourth period since  
 $4s^2 4p^1 \rightarrow Ga$ )

**2.50**

Write the electron configurations of the following elements by using *spdf* notation:

(a) Aluminum, (b) Titanium, (c) Iron, (d) Tungsten.

**Answer 2.50:**

(a) Al ( $Z=13$ ), [Ne]  $3s^23p^1$  (b) Ti ( $Z=22$ ), [Ar]  $4s^23d^2$

(c) Fe ( $Z=26$ ), [Ar]  $4s^23d^6$  (d) W ( $Z=74$ ), [Xe]  $4f^{14}5d^46s^2$

\*2.51 Determine the four quantum numbers for the 3<sup>rd</sup>, 15<sup>th</sup>, and 17<sup>th</sup> electrons of the Cl atom.

Answer 2.51:

2.51

Cl has 17 electrons

3<sup>rd</sup> electron of the Cl atom:  $n=2, l=0, s, m_l=0, m_s=+\frac{1}{2}$

15<sup>th</sup> electron of the Cl atom:  $n=3, l=1, P, m_l=+1, m_s=+\frac{1}{2}$

17<sup>th</sup> electron of the Cl atom:  $n=3, l=1, P, m_l=0, m_s=-\frac{1}{2}$

\*2.52 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
- ii) Sr, H, Ba, He, Mg, Cs

Answer 2.52:

2.52

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

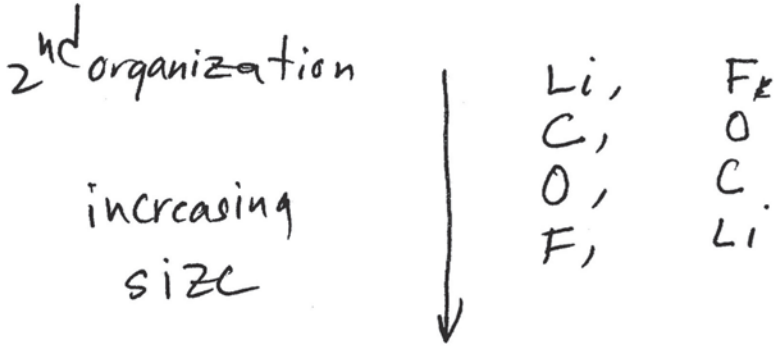
period 2 elements will be smaller than period 3, than period 6.

1<sup>st</sup> organization

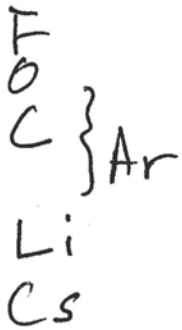
period 2      Li, F, O, C  
period 3      Ar  
period 6      Cs

↓ increasing size

within each period, size decreases moving to the right



Ar, last element in period 3, will be smaller than Li but larger than F



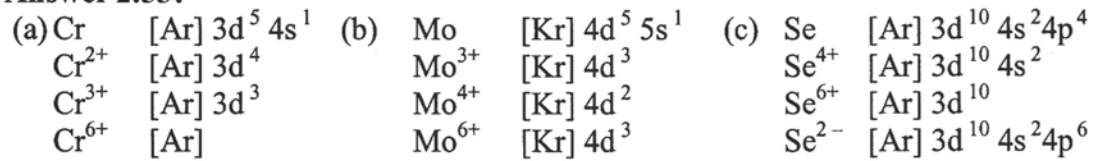
Cs, will be the largest since in period 6

2.53

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

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

**Answer 2.53:**





2.54 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

Answer 2.54:

2.54

a) Atomic Size (increasing)

i) as we move to left in a period, atomic size increases

Ga: period 4 group 3	} ⇒	Ga, Ca, K (increasing order)
Ca: period 4 group 2		
K: period 4 group 1		

b) Ionization Energy (decreasing)

as we move to left in a period, ionization energy increases

⇒ Ga, Ca, K (decreasing order)

a) Atomic Size (increasing)

ii)

Ca: period 4 group 2

Sr: period 5 group 2

Ba: period 6 group 2

as we move down in a group, the atomic size increases.

⇒ Ca, Sr, Ba (increasing size)

b) Ionization Energy (decreasing)

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

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

Ca, Sr, Ba

a) Atomic Size (increasing)

iii

I : period 5, group 7

Xe: period 5, group 8

Cs : period 6, group 1

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

Cs in period 6 will be larger than both.

Cs, I, Xe (increasing size)

b) Ionization Energy (decreasing)

iii)

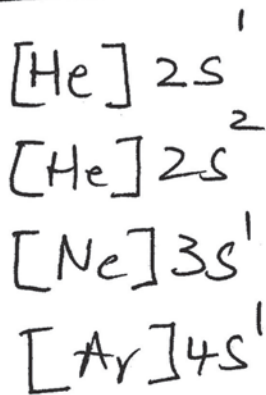
- in period 5, I will have a lower ionization energy than Xe
- Cs, in period 6, will have the lowest ionization energy.

Xe, I, Cs

2.55 The first ionization energies of three atoms with electronic configurations (a)  $[\text{He}]2s^2$ , (b)  $[\text{Ne}]3s^1$ , (c)  $[\text{Ar}]4s^1$  and (d)  $[\text{He}]2s^1$  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.

Answer 2.55:

2.55



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

$$\begin{array}{l} IE_1 [\text{Ar}]4s^1 = 419 \text{ kJ/mol} \\ IE_1 [\text{Ne}]3s^1 = 496 \text{ kJ/mol} \\ IE_1 [\text{He}]2s^1 = 520 \text{ kJ/mol} \end{array}$$

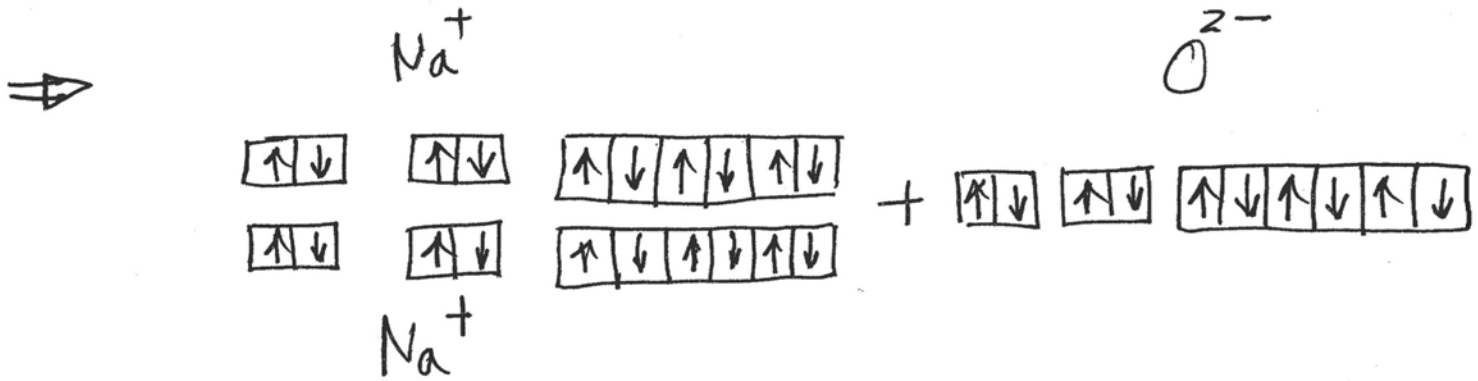
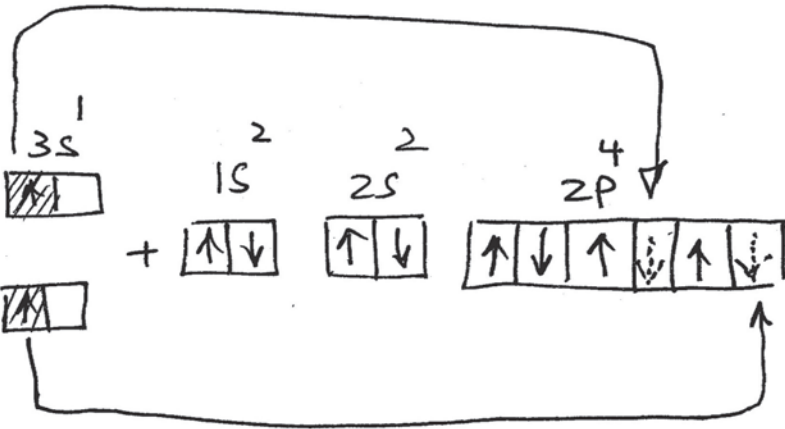
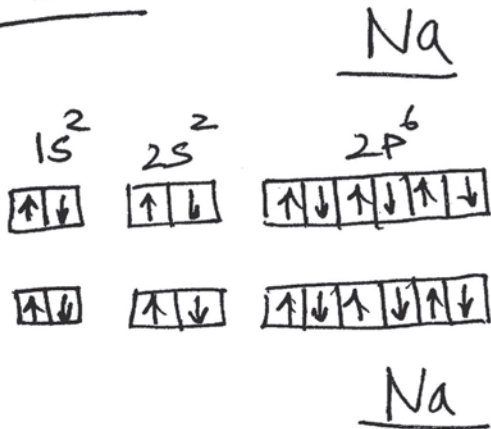
$[\text{He}]2s^2$  belongs to the second group in period II. Thus its  $IE_1$  must be higher than  $[\text{He}]2s^1$ .  
Thus,  $IE_1 [\text{He}]2s^2 = 899 \text{ kJ/mol}$

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

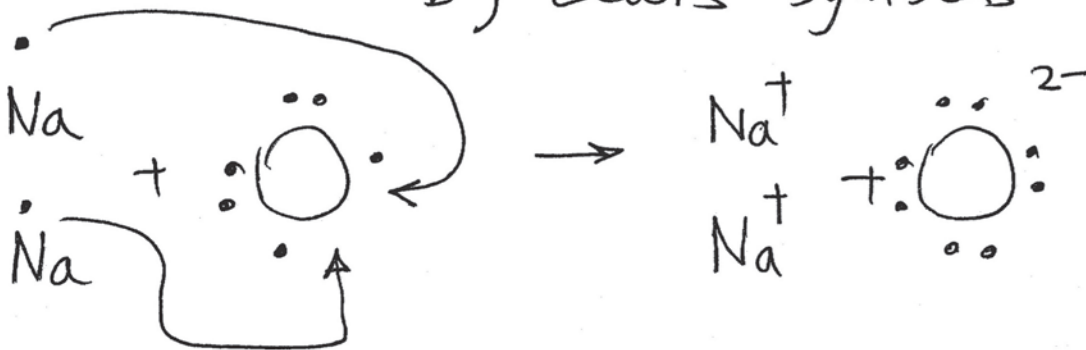
Answer 2.56:

2.56

a) orbital diagram



b) Lewis symbols



chemical formula =  $\text{Na}_2\text{O}$

2.57 The first ionization energies of two atoms with electronic configurations (a)  $1s^2 2s^2 2p^6$  and (b)  $1s^2 2s^2 2p^6 3s^1$  are given to be 2080 kJ/mol and 496 kJ/mol. Determine which IE1 belongs to which electronic structure and justify your answer.

Answer 2.57:

2.57

$1s^2 2s^2 2p^6$  : 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<sub>1</sub>.

$1s^2 2s^2 2p^6 3s^1$  : This atom belongs to group I, period III. There is a single electron in  $n=3$ . It will be easier to ionize. It must have the lower IE<sub>1</sub>.

2.58 Calculate the attractive force ( $\bullet \rightarrow \leftarrow \bullet$ ) between a pair of  $\text{Ba}^{2+}$  and  $\text{S}^{2-}$  ions that just touch each other. Assume the ionic radius of the  $\text{Ba}^{2+}$  ion to be 0.143 nm and that of the  $\text{S}^{2-}$  ion to be 0.174 nm.

**Answer 2.58:**

The attractive force between the ion pair is found by applying Coulomb's law,

$$F_{\text{attractive}} = \frac{-Z_1 Z_2 e^2}{4\pi\epsilon_0 a_o^2}$$

Where  $Z_1 = +2$  for  $\text{Ba}^{2+}$ ,  $Z_2 = -2$  for  $\text{S}^{2-}$ , and

$$a_o = r_{\text{Ba}^{2+}} + r_{\text{S}^{2-}} = 0.143 \text{ nm} + 0.174 \text{ nm} = 0.317 \text{ nm} = 3.17 \times 10^{-10} \text{ m}$$

Substituting,

$$F_{\text{attractive}} = \frac{-(+2)(-2)(1.60 \times 10^{-19} \text{ C})^2}{4\pi(8.85 \times 10^{-12} \text{ C}^2 / \text{N} \cdot \text{m}^2)(3.17 \times 10^{-10} \text{ m})^2} = 9.16 \times 10^{-9} \text{ N}$$

2.59 Calculate the net potential energy for a  $\text{Ba}^{2+} \text{S}^{2-}$  ion pair by using the  $b$  constant calculated from Prob. 2.58. Assume  $n = 10.5$ .

**Answer 2.59:**

The repulsive energy constant  $b$  is:

$$b = \frac{-F_{\text{Repulsive}} a_o^{n+1}}{n} = \frac{-(-F_{\text{Attractive}}) a_o^{n+1}}{n} = \frac{-(-9.16 \times 10^{-9} \text{ N})(3.17 \times 10^{-10} \text{ m})^{11.5}}{10.5}$$

$$= 5.045 \times 10^{-119} \text{ N} \cdot \text{m}^{10}$$

Thus the net potential energy between the ions is,

$$E_{\text{Ba}^{2+}\text{S}^{2-}} = \frac{+Z_1 Z_2 e^2}{4\pi\epsilon_0 a_o} + \frac{b}{a_o^n}$$

$$= \frac{+(+2)(-2)(1.60 \times 10^{-19} \text{ C})^2}{4\pi(8.85 \times 10^{-12} \text{ C}^2 / (\text{N} \cdot \text{m}^2))(3.17 \times 10^{-10} \text{ m})} + \frac{5.045 \times 10^{-119} \text{ N} \cdot \text{m}^{10}}{(3.17 \times 10^{-10} \text{ m})^{10.5}}$$

$$= (-2.905 \times 10^{-18} \text{ J}) + (2.765 \times 10^{-19} \text{ J}) = -2.63 \times 10^{-18} \text{ J}$$

**2.60** If the attractive force between a pair of  $\text{Cs}^+$  and  $\text{I}^-$  ions is  $2.83 \times 10^{-9}$  N and the ionic radius of the  $\text{Cs}^+$  ion is 0.165 nm, calculate the ionic radius of the  $\text{I}^-$  ion in nanometers.

**Answer 2.60:**

From Coulomb's law,

$$a_o = \sqrt{\frac{-Z_1 Z_2 e^2}{4\pi\epsilon_o F_{\text{Attractive}}}} = \sqrt{\frac{- (+1)(-1)(1.60 \times 10^{-19} \text{ C})^2}{4\pi(8.85 \times 10^{-12} \text{ C}^2/(\text{N} \cdot \text{m}^2))(2.83 \times 10^{-9} \text{ N})}}$$
$$= 2.852 \times 10^{-10} \text{ m} = 0.2852 \text{ nm}$$

The ionic radius of iodine is thus,

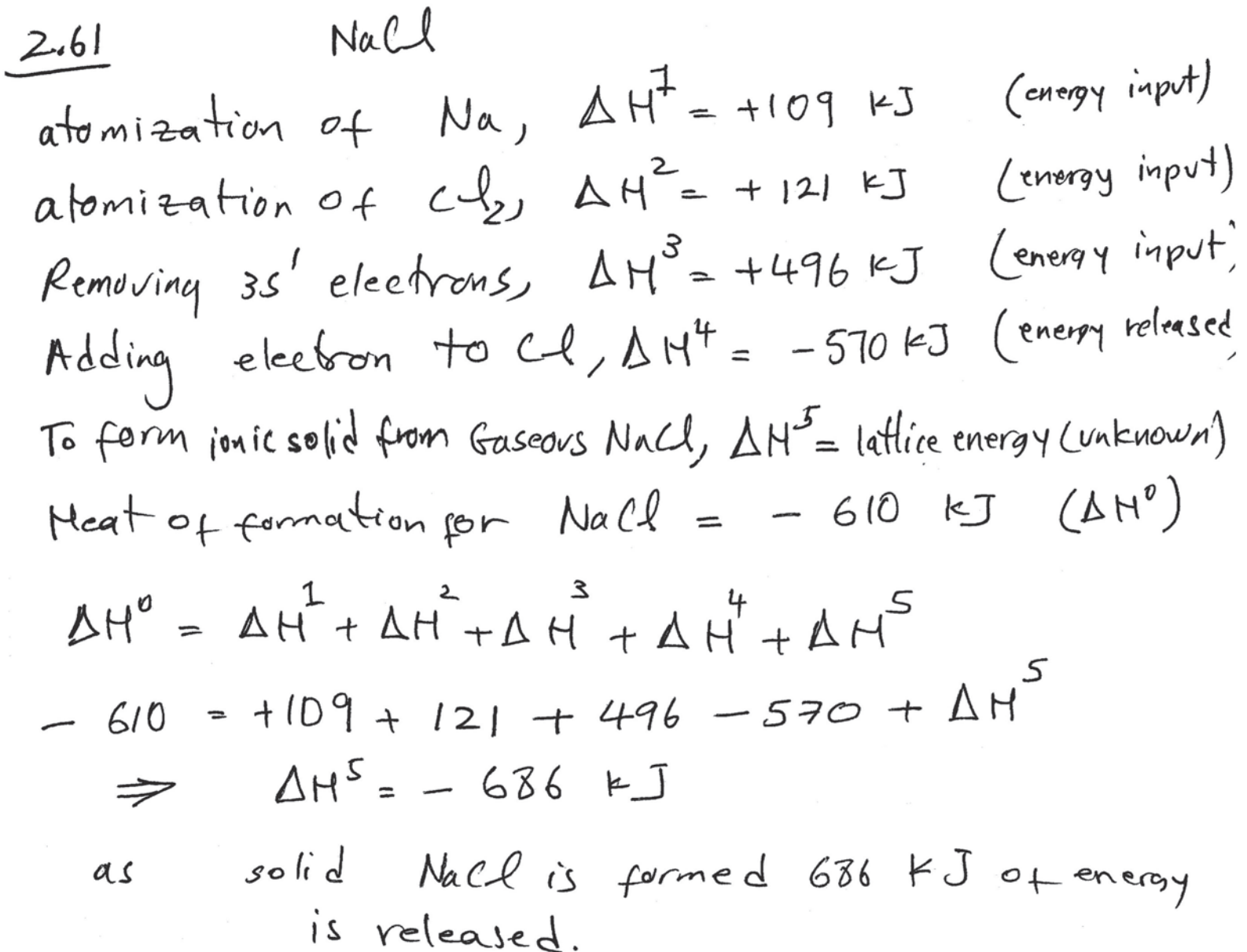
$$r_{\text{I}^-} = a_o - r_{\text{Cs}^+} = 0.285 \text{ nm} - 0.165 \text{ nm} = \mathbf{0.120 \text{ nm}}$$



2.61 Calculate the lattice energy for the formation of solid NaCl if the following information is given. What does the calculated lattice energy tell you about the material?

- 109 kJ is required to convert solid Na to gaseous Na
- 121 kJ is required to convert gaseous  $\text{Cl}_2$  to two monatomic Cl atoms
- 496 kJ is required to remove the  $3s^1$  electron of Na (form  $\text{Na}^+$  cation)
- 570 kJ of energy (energy is released) to add an electron to the Cl
- 610 kJ of energy to form gaseous NaCl (heat of formation of NaCl)

Answer 2.61:



\*2.62 For the each pair of compounds presented below, determine which has the higher lattice energy (more negative). Explain your answer. Also, which of the six ionic compound do you think has the highest melting temperature and why? Verify your answers by checking Table 2.5.

- (a) LiCl and CsCl
- (b) Cs Cl and RbCl
- (c) LiF and MgO
- (d) MgO and CaO

Answer 2.62:

2.62

Cl is common to both compounds.

a) LiCl CsCl : as the size of the cation increases ( $Cs > Li$ ), the lattice energy decreases  
Thus LiCl will have higher lattice energy than CsCl

b) CsCl RbCl : Cl is common to both compound  
Size of Cs is larger than Rb ( $Cs > Rb$ )  
Thus RbCl has a higher lattice energy than CsCl.

c) LiF MgO :  $Mg^{++}$  and  $O^{--}$  have higher ionic charge than  $Li^+$  and  $F^-$ , Thus MgO will have a higher lattice energy.

d)  $\text{MgO}$   $\text{CaO}$  : O is common to both compounds. All ions have a ionic charge of 2 :  $\text{Mg}^{++}$ ,  $\text{O}^{--}$ , and  $\text{Ca}^{++}$ . However,  $\text{Ca}^{++}$  is larger in radius than  $\text{Mg}^{++}$ . Thus,  $\text{MgO}$  will have higher lattice energy.

Overall,  $\text{MgO}$  will have the highest lattice energy and therefore highest melting point.

2.63 For each bond in the following series of bonds, determine the bond order, rank bond length, and rank bond strength. Use only the periodic table. Explain your answers.

(a) S-F; S-Br; S-Cl

(b) C-C; C=C; C≡C

Answer 2.63:

2.63

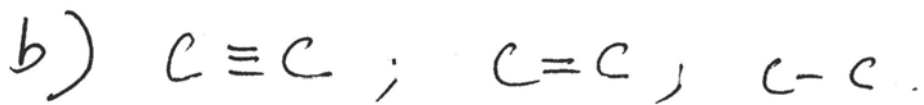
a) S-F, S-Br; S-Cl

Note S is the common element. All bonds have order "one" since all have one shared pair.

Since the diameter of Br is larger than Cl and Cl is larger than F, the bond length of S-Br will be greater than S-Cl, greater than S-F.

Bond strength will have an inverse relationship with bond length.

Thus, S-F stronger than S-Cl stronger than S-Br.



The bond order is directly related to number of shared pairs.

$C \equiv C$ , (bond order of 3

$C = C$ , bond order of 2

$C - C$  bond order of 1

As bond order increases, bond length decreases.

Thus,  $C \equiv C$  smaller bond length than  $C = C$  than  $C - C$ .  
Bond strength is inversely related to bond length.  $C \equiv C$  strongest  
 $C - C$  weakest.

\*2.64 Calculate the lattice energy for the formation of solid NaF if the following information is given. What does the calculated lattice energy tell you about the material?

109 kJ is required to convert solid Na to gaseous Na

243 kJ is required to convert gaseous  $F_2$  to two monatomic F atoms

496 kJ is required to remove the 3s<sup>1</sup> electron of Na (form Na<sup>+</sup> cation)

-349 kJ of energy (energy is released) to add an electron to the F (form Na<sup>-</sup> anion)

-411 kJ of energy to form gaseous NaF (heat of formation of NaF)

**Answer 2.64:**

2.64

NaF

atomization of Na,  $\Delta H^1 = +109$  kJ (energy input)

atomization of  $F_2$ ,  $\Delta H^2 = +243$  kJ (energy input)

Removing 3s<sup>1</sup> electrons,  $\Delta H^3 = +496$  kJ (energy input)

Adding electron to F,  $\Delta H^4 = -349$  kJ (energy released)

to form ionic solid from gaseous NaF,  $\Delta H^5 =$  lattice energy (unknown)

Heat of formation for NaF =  $-411$  kJ ( $\Delta H^0$ )

$$\Delta H^0 = \Delta H^1 + \Delta H^2 + \Delta H^3 + \Delta H^4 + \Delta H^5$$

$$-411 = +109 + 243 + 496 - 349 + \Delta H^5$$

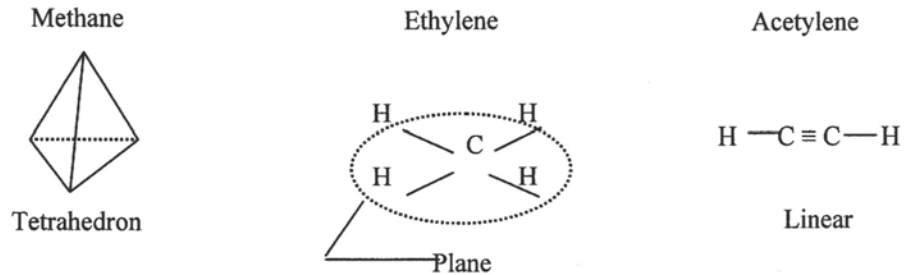
$$\Rightarrow \Delta H^5 = -910 \text{ kJ}$$

as solid NaF is formed 910 kJ is released.

**2.65** List the number of atoms bonded to a C atom that exhibits  $sp^3$ ,  $sp^2$ , and  $sp$  hybridization. For each, give the geometrical arrangement of the atoms in the molecule.

**Answer 2.65:**

- $sp^3$  hybridization: Four atoms are bonded to a central carbon atom in a tetrahedral arrangement. An example is methane,  $CH_4$ .
- $sp^2$  hybridization: Three atoms are bonded to a carbon atom in a planar arrangement. An example is ethylene,  $CH_2=CH_2$ .
- $sp$  hybridization: Two atoms are bonded to a carbon atom in a linear arrangement. An example is acetylene,  $CH\equiv CH$ .



**2.66** Is there a correlation between the electron configurations of the elements potassium ( $Z = 19$ ) through copper ( $Z = 29$ ) and their melting points? (See Table 2.7).

**Answer 2.66:**

A possible correlation between the melting points and the electron configurations of the elements from scandium ( $Z = 21$ ) to copper ( $Z = 29$ ) is that unpaired  $3d$  electrons cause covalent hybridized bonds, and hence give higher melting points to these transition metals.

2.67 Rank the following covalently bonded atoms according to the degree of polarity:  
C-Cl; C-Ge; C-B; C-C.

Answer 2.67:

2.67

C-Cl  
2.5 3.0

high polarity — electron shared more by Cl

C-Ge  
2.5 1.8

high polarity — electron shared more by C

C-B  
2.5 2.0

high polarity — electron shared more by C

C-C  
2.5 2.5

nonpolar, no difference in electronegativity

therefore,

$C-Ge > C-B > C-Cl > C-C$



**\*2.68** Most modern scanning electron microscopes (SEMs) are equipped with energy dispersive x-ray detectors for the purpose of chemical analysis of the specimens. This x-ray analysis is a natural extension of the capability of the SEM because the electrons that are used to form the image are also capable of creating characteristic x-rays in the sample. When the electron beam hits the specimen, x-rays specific to the elements in the specimen are created. These can be detected and used to deduce the composition of the specimen from the well-known wavelengths of the characteristic x-rays of the elements. For example:

Element	Wavelength of $K_{\alpha}$ x-rays
Cr	0.2291 nm
Mn	0.2103 nm
Fe	0.1937 nm
Co	0.1790 nm
Ni	0.1659 nm
Cu	0.1542 nm
Zn	0.1436 nm

Suppose a metallic alloy is examined in an SEM and three different x-ray energies are detected. If the three energies are 7492, 5426, and 6417 eV, what elements are present in the sample? What would you call such an alloy? (Look ahead to Chap. 9 in the textbook.)

**Answer 2.68:**

The elements may be identified by calculating their respective wavelengths.

$$(a) \quad \lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(7492 \text{ eV})(1.60 \times 10^{-19} \text{ J/eV})}$$

$$= 1.659 \times 10^{-10} \text{ m} = 0.1659 \text{ nm} \Rightarrow \text{Ni}$$

From the table provided, a wavelength of 0.1659 nm corresponds to nickel, Ni.

$$(b) \quad \lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(5426 \text{ eV})(1.60 \times 10^{-19} \text{ J/eV})}$$

$$= 2.291 \times 10^{-10} \text{ m} = 0.2291 \text{ nm} \Rightarrow \text{Cr}$$

From the table provided, a wavelength of 0.2291 nm corresponds to chromium, Cr.

$$(c) \quad \lambda = \frac{hc}{E} = \frac{(6.63 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(6417 \text{ eV})(1.60 \times 10^{-19} \text{ J/eV})}$$

$$= 1.937 \times 10^{-10} \text{ m} = 0.1937 \text{ nm} \Rightarrow \text{Fe}$$

From the table provided, a wavelength of 0.1937 nm corresponds to iron, Fe.

The elements present, nickel, chromium, and iron, are the primary constituents of the austenitic stainless steels.

**\*2.69** Compare the percentage ionic character in the semiconducting compounds CdTe and InP.

**Answer 2.69:**

Applying Pauling's equation to CdTe and InP compounds,

$$\text{For CdTe (2 - 6), \% Ionic character} = (1 - e^{-0.25(1.5-2.0)^2})(100\%) = 6.1\%$$

$$\text{For InP (3 - 5), \% Ionic character} = (1 - e^{-0.25(1.5-2.1)^2})(100\%) = 8.6\%$$

While a 2 – 6 compound typically has a higher ionic character than a 3 – 5, the relatively high electronegativity of phosphorous causes InP to be more ionic in nature.

**\*2.70** Both K and Ca are group 2A metals. Why is the melt temperature of Ca higher than the melt temperature of K?

**Answer 2.70:**

In metals, as the number of valence electrons increases, the attraction force between the positive core and the electron cloud also increases. In group 2A metals, for example, K with one valence electron has a melt temperature of 63.5°C compared to considerably higher 851°C for Ca with two valence electrons.

**2.71** Of the noble gases Ne, Ar, Kr, and Xi, which should be the most chemically reactive?

**Answer 2.71:**

Xenon should be most reactive since its outermost electrons ( $5s^26p^6$ ) are further away from the nucleus than the other noble gases, and thus easier to remove.

**Synthesis and Evaluation Problems:**

2.72  $^{39}\text{K}$ ,  $^{40}\text{K}$ , and  $^{41}\text{K}$  are the three isotopes of potassium. If  $^{40}\text{K}$  has the lowest %abundance, which other isotope has the highest?

Answer 2.72:

2.72

Note from the periodic table that the relative atomic mass of K is 39.10 grams. This number is much closer to  $^{39}\text{K}$  than to  $^{40}\text{K}$ . Thus, given that  $^{41}\text{K}$  is the least abundant,  $^{39}\text{K}$  must be the most abundant.

2.73 According to section 2.5.1, in order form monatomic ions from metals and nonmetals, energy must be added. However, we know that primary bonds form because the involved atoms want to lower their energies. Why then do ionic compounds form?

**Answer 2.73:**

2.73

Although energy is added during atomization of metals and nonmetals, when solid ionic materials are formed, a significant amount of energy is released (lattice energy). The activity will result in net release of energy. Thus, the energy of the atoms involved is lowered.

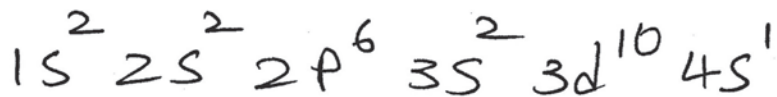
2.74 Cartridge brass is an alloy of two metals: 70 wt% copper and 30 wt% zinc. Discuss the nature of the bonds between copper and zinc in this alloy.

Answer 2.74:

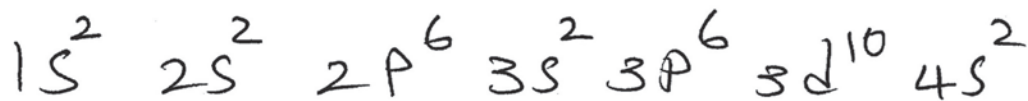
2.74

Both copper and zinc are in period IV of the periodic table.

The electronic structure of copper is



and that of zinc is



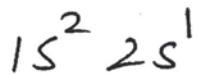
The 3d and 4s orbitals of zinc are full. Thus it has significantly less tendency for covalent bonding. When Cu is replaced by Zn, the overall percentage of covalent bonds drops. Thus brass (70 wt% Cu - 30 wt% Zn) will have a higher percentage metallic bonds when compared to Cu. Therefore Brass will have a lower melt temp than Cu and higher than Zn.

\*2.75 The melt temperature of Li ( $180^{\circ}\text{C}$ ) is significantly lower than the melt temperature of its neighbor Be ( $1287^{\circ}\text{C}$ ). Can you explain this in terms of the differences in electronic structure?

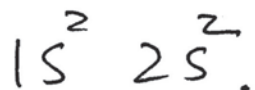
**Answer 2.75:**

2.75

The electronic structure of Li is



and that of Be is



There are more  $+$  particles in the nucleus of Be than in Li. Thus, the attraction force between the valence electrons and the nucleus is larger in Be than in Li.

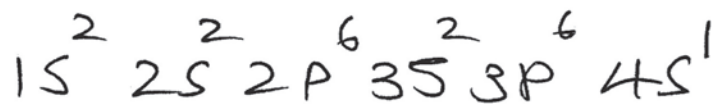
Therefore, the melt temperature of Be is higher. due to its higher nucleus charge.

2.76 The melting point of the metal potassium is  $63.5^{\circ}\text{C}$ , while that of titanium is  $1660^{\circ}\text{C}$ . What explanation can be given for this great difference in melting temperatures?

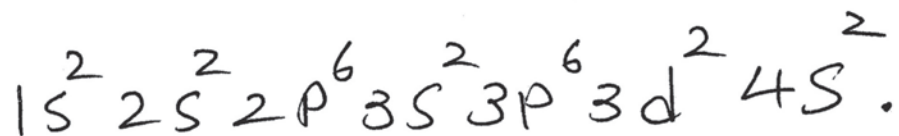
**Answer 2.76:**

2.76

Both K & Ti are located in period IV of the periodic table. The electronic structure of K is



and that of Ti is



First, as the number of valence electrons increases such as in Ti, the attraction force between the electron and the nucleus also increases.

As 3d electrons are introduced (in Ti) the melt temperature further increases.

This is due to more covalent bonding being introduced.



**2.77** Regardless of the type of primary bond, why does this tendency exist for atoms to bond?

**Answer 2.77:**

The tendency of an atom to bond with other atoms lowers its energy level and enables it to exist in the most stable state after bonding.

\*2.78 Pure aluminum is a ductile metal with low tensile strength and hardness. Its oxide  $\text{Al}_2\text{O}_3$  (Alumina) is extremely strong, hard, and brittle. Can you explain this difference from an atomic bonding point of view?

Answer 2.78:

2.78

pure aluminum is a metal with mostly metallic bonds.

Alumina ( $\text{Al}_2\text{O}_3$ ) is a ceramic with mostly ionic bonds in a 3-d network.

The strength of the ionic bonds is significantly higher than the metallic bonds.

Thus, more energy is required to break the bonds between Al and O than between Al and Al.

**2.79** After ionization, why is the sodium ion smaller than the sodium atom? After ionization, why is the chloride ion larger than the chlorine atom?

**Answer 2.79:**

After ionization to the  $\text{Na}^+$ , the Na atom becomes smaller because the electron-to-proton ratio of the Na atom is decreased when the  $\text{Na}^+$  ion forms. Also, the outer third shell no longer exists once the  $3s^1$  electron is lost by the Na atom.

After ionization, the  $\text{Cl}^-$  ion is larger because the electron-to-proton ratio of the chlorine atom is decreased by the ionization process.

2.80 Silicon is extensively used in the manufacturing of integrated circuit devices such as transistors and light emitting diodes. It is often necessary to develop a thin oxide layer ( $\text{SiO}_2$ ) on silicon wafers. a) What are the differences in properties between the silicon substrate and the oxide layer? b) Design a process that produces the oxide layer on a silicon wafer. c) Design a process that forms the oxide layer only in certain desired areas.

Answer 2.80:

2.80

a) Si is a metalloid (refer to the periodic table,  
page 32.

It has a structure similar to diamond (Fig. 2.2

It has excellent semiconductive characteristics.

$\text{SiO}_2$  is a ceramics. It has excellent insulative characteristics. The bonding is mostly ionic.

b) Exposing Si to an oxygen atmosphere at high temperatures will create a thin oxide layer on the surface of Si.

c) After oxidation, we can expose certain regions to chemical etchants and remove the oxide material. This is often done by applying a mask with a specific pattern to the surface of Si (see figure 14.44).

**2.81** Explain, on average, the melting points of pure metals are lower than ceramics. In general, alkali earth metals have higher melting points than alkali metals. Why?

**Answer 2.81:**

The pure metals do not require to break the bond between the ionic core and the electron cloud. Thus, on average, ceramics (most are ionic and covalent bond materials) have higher melting temperatures because they require breakage of the bonds for melting. In general, alkali metals (group 1A) have only one valence electron and possess almost exclusively metallic bonds. As a result, these metals have lower melting temperatures than alkali earth metals (group 2A) that have two valence electrons and a higher percentage of covalent bonding.

2.82 Graphite and diamond are both made from carbon atoms. a) List some of the physical characteristics of each. b) Give one application for graphite and one for diamond. c) If both materials are made of carbon, why does such a difference in properties exist?

Answer 2.82:

2.82 a)  
Graphite: dark color material that can have powder (flake-like) form or fiber form. It is an electrical conductor. strength is high in the transverse plane, but not high in other directions (anisotropic). It has a layered hexagonal structure. The bond inside the layer are covalent and the bond between the layers are of secondary (electrostatic) nature.

Diamond: Transparent, hard, strong, very high melt temperature. It is an electrical insulator. It has a 3-D networked structure. all bonds are covalent.

b)

Graphite is used as a lubricant because various layers are weakly bonded and can slide on top of each other.

Diamond (synthetic type) is used as abrasive and cutting tool materials.

c) The bonding structure for diamond is 3-D covalent network (very strong) but graphite is layered with very weak forces between layers.



**2.83** Methane ( $\text{CH}_4$ ) has a much lower boiling temperature than does water ( $\text{H}_2\text{O}$ ). Explain why this is true in terms of the bonding between molecules in each of these two substances.

**Answer 2.83:**

The methane molecules are bonded together by weak  $\text{C-H}$  dipoles. The water molecules are bonded together by the much stronger  $\text{O-H}$  hydrogen bonded dipoles.

**2.84** Describe fluctuating dipole bonding among the atoms of the noble gas neon. Of a choice between the noble gases krypton and xenon, which noble gas would be expected to have the strongest dipole bonding and why?

**Answer 2.84:**

A fluctuating electric dipole exists in the atoms of noble gases, such as neon, because there is, at any instant, an asymmetrical distribution of electrical charge among their electrons. The noble gas xenon would be expected to have a stronger fluctuating dipole moment than krypton since it has an additional electron shell; the krypton atom has four electron shells while the xenon atom has five. The electrons of this fifth shell, being further away from the xenon nucleus, and are able to fluctuate more and thus create greater asymmetry of charge.

**2.85** Carbon tetrachloride ( $\text{CCl}_4$ ) has a zero dipole moment. What does this tell us about the  $\text{C-Cl}$  bonding arrangement in this molecule?

**Answer 2.85:**

Since the molecule  $\text{CCl}_4$  has a zero dipole moment, the  $\text{C-Cl}$  bonding arrangement must be symmetrical about the carbon nucleus.

2.86 Stainless steel is a corrosion resistant metal because it contains large amounts of chromium in its composition. How does chromium protect the metal from corrosion?

Answer 2.86:

2.86

Chromium easily reacts with oxygen in the atmosphere and form chromium oxide  $\text{Cr}_2\text{O}_3$ . Chromium oxide is a ceramic with high hardness. The chromium on the surface forms a thin protective ceramic layer that is chemically stable and protects the material inside. This is called a passivity layer.

**2.87** For each of the following compounds, state whether the bonding is essentially metallic, covalent, ionic, van der Waals, or hydrogen: (a) Ni, (b) ZrO<sub>2</sub>, (c) graphite, (d) solid Kr, (e) Si, (f) BN, (g) SiC, (h) Fe<sub>2</sub>O<sub>3</sub>, (i) MgO, (j) W, (k) H<sub>2</sub>O within the molecules, (l) H<sub>2</sub>O between the molecules. If ionic and covalent bonding are involved in the bonding of any of the compounds listed, calculate the percentage ionic character in the compound.

**Answer 2.87:**

- |   |  |
|---|--|
| (a) Ni:                                     | Nickel bonding is primarily metallic.  |
| (b) ZrO <sub>2</sub> :                      | From Pauling's equation, the Zr–O bond is 73.4% ionic and 26.6% covalent, where $x_A$ and $x_B$ are the electronegativities of zirconium and oxygen, respectively. |
| (c) Graphite:                               | The bonding is covalent within the layers and secondary between the layers.  |
| (d) Solid Kr:                               | The bonding represents van der Waals due to fluctuating dipoles.   |
| (e) Si:                                     | Silicon bonding is covalent.   |
| (f) BN:                                     | The B–N bond, from Pauling's equation, is 26.1% ionic and 73.9% covalent.  |
| (g) SiC:                                    | The Si–C bond is 11% ionic and 89% covalent.   |
| (h) Fe <sub>2</sub> O <sub>3</sub> :        | The Fe–O bond is 55.5% ionic and 44.5% covalent.   |
| (i) MgO:                                    | The Mg–O bond is 70.2% ionic and 29.8% covalent.   |
| (j) W:                                      | Tungsten bonding primarily consists of metallic bonding with some covalent character.  |
| (k) H <sub>2</sub> O within the molecules:  | The H–O bond is 38.7% ionic and 61.3% covalent.  |
| (l) H <sub>2</sub> O between the molecules: | Hydrogen bonding exists between H <sub>2</sub> O molecules.  |

\*2.88 In the manufacturing of a light bulb, the bulb atmosphere is evacuated and then filled with argon gas. What is the purpose of this?

Answer 2.88:

2-88

Because, the tungsten filament in a light bulb becomes hot, it can react (form bonds) with impurities in the atmosphere of a light bulb.

This will weaken and damage the filament.

Argon is a noble gas and is non-reactive (inert). It does not react with tungsten and keeps the integrity of the wire.

2.89 A certain application requires a material that is lightweight, an electrical insulator, and has some flexibility. a) Which class of materials would you search for this selection? b) Explain your answer from a bonding point of view.

Answer 2.89:

2.89

- a) The characteristics required are found in polymers.
- b) polymers are insulators because all valence electrons are involved in covalent bonds within the polymeric molecule.
- The bonding between molecules however are of the secondary type (electrostatic attraction). This allows the material to be flexible.
  - polymers are also low density because there exist large empty spaces (gaps) in the structure between molecules.

2.90 A certain application requires a material that is electrically non-conductive (insulator), extremely stiff, and lightweight. Which classes of materials would you search for this selection? b) Explain your answer from a bonding point of view.

Answer 2.90:

- 2.90
- a) The only class of materials that possess the required properties is structural ceramics.
- b) structural ceramics are mostly made of ionically bonded atoms in a 3-D network. They are insulators because the valence electrons are tightly held in ionic bonds. Ionic bonds are very strong and require high temperatures for melting.

2.91 Robots are used in auto industries to weld two components at specific locations. Clearly, the end-position of the arm must be determined accurately in order to weld the components at the exact position. a) In selecting the material for the arm of such robots, what factors must be considered? b) Select a proper material for this application.

**Answer 2.91:**

2.91

- a) - one would have to choose materials that are very stiff (little elastic deformation under load).  
- The stronger the bonds are between the atoms in the material, the stiffer the material will behave.
- b) Thus ceramics, advanced composites, and some metals will be good choices. (economic considerations are important).