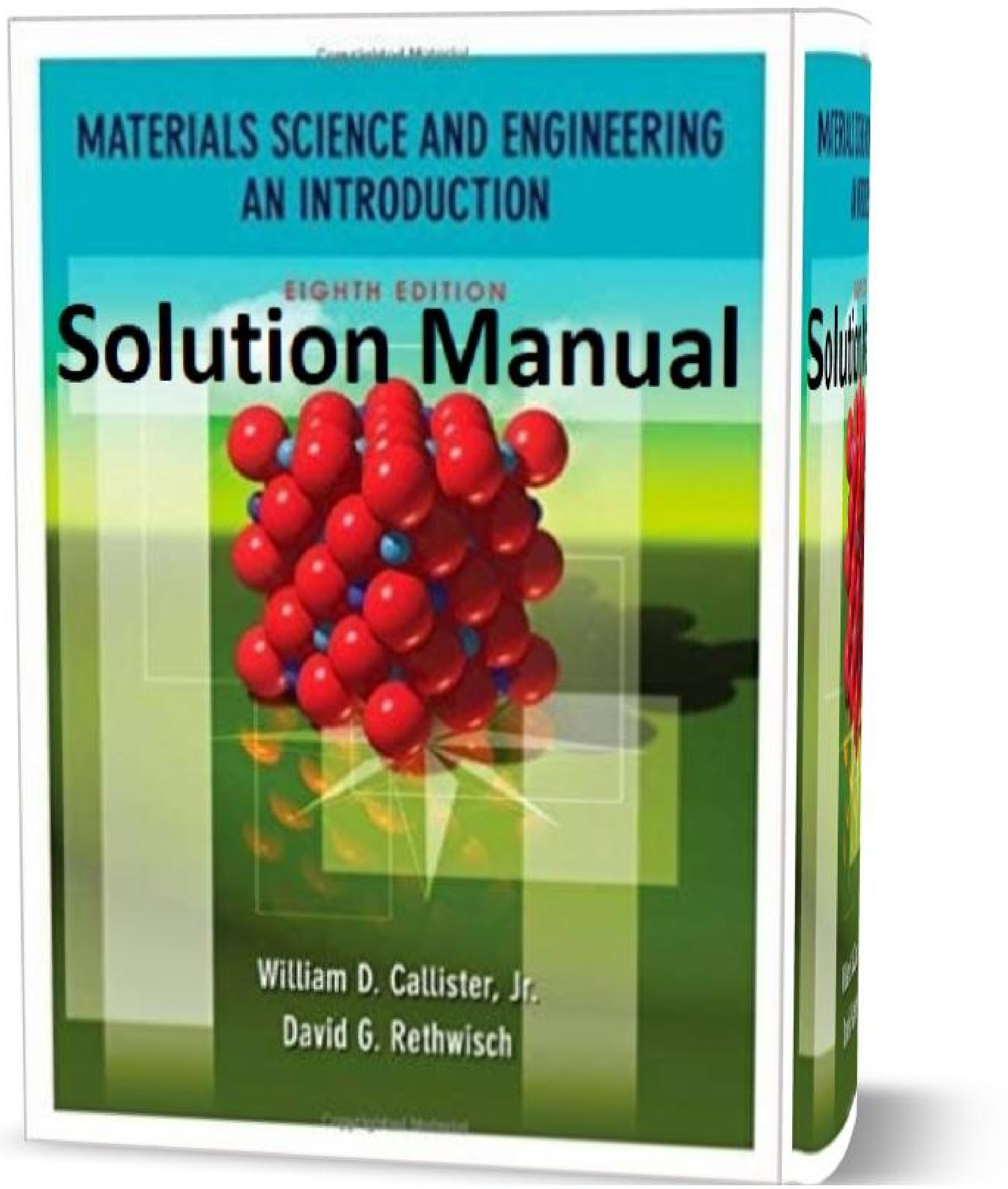
MATERIALS SCIENCE AND ENGINEERING
AN INTRODUCTION
EIGHTH EDITION
Solution Manual
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CHAPTER 2

ATOMIC STRUCTURE AND INTERATOMIC BONDING

PROBLEM SOLUTIONS

Fundamental Concepts
Electrons in Atoms

2.1 *Cite the difference between atomic mass and atomic weight.*

Solution

Atomic mass is the mass of an individual atom, whereas atomic weight is the average (weighted) of the atomic masses of an atom's naturally occurring isotopes.

2.2 *Chromium has four naturally-occurring isotopes: 4.34% of* ^{50}Cr*, with an atomic weight of 49.9460 amu, 83.79% of* ^{52}Cr*, with an atomic weight of 51.9405 amu, 9.50% of* ^{53}Cr*, with an atomic weight of 52.9407 amu, and 2.37% of* ^{54}Cr*, with an atomic weight of 53.9389 amu. On the basis of these data, confirm that the average atomic weight of Cr is 51.9963 amu.*

Solution

The average atomic weight of silicon $(\overline{A}_{Cr})$ is computed by adding fraction-of-occurrence/atomic weight products for the three isotopes. Thus

$$\overline{A}_{Cr} = f_{^{50}Cr} A_{^{50}Cr} + f_{^{52}Cr} A_{^{52}Cr} + f_{^{53}Cr} A_{^{53}Cr} + f_{^{54}Cr} A_{^{54}Cr}$$

$$= (0.0434)(49.9460 \text{ amu}) + (0.8379)(51.9405 \text{ amu}) + (0.0950)(52.9407 \text{ amu}) + (0.0237)(53.9389 \text{ amu}) = 51.9963 \text{ amu}$$

2.3 *(a) How many grams are there in one amu of a material?*

(b) Mole, in the context of this book, is taken in units of gram-mole. On this basis, how many atoms are there in a pound-mole of a substance?

Solution

(a) In order to determine the number of grams in one amu of material, appropriate manipulation of the amu/atom, g/mol, and atom/mol relationships is all that is necessary, as

$$\#\ \text{g/amu} = \left(\frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{atoms}}\right)\left(\frac{1\ \text{g/mol}}{1\ \text{amu/atom}}\right)$$

$$= 1.66 \times 10^{-24}\ \text{g/amu}$$

(b) Since there are 453.6 g/lb_m,

$$1\ \text{lb-mol} = (453.6\ \text{g/lb}_m)(6.022 \times 10^{23}\ \text{atoms/g-mol})$$

$$= 2.73 \times 10^{26}\ \text{atoms/lb-mol}$$

2.4 *(a) Cite two important quantum-mechanical concepts associated with the Bohr model of the atom.*

(b) Cite two important additional refinements that resulted from the wave-mechanical atomic model.

Solution

(a) Two important quantum-mechanical concepts associated with the Bohr model of the atom are (1) that electrons are particles moving in discrete orbitals, and (2) electron energy is quantized into shells.

(b) Two important refinements resulting from the wave-mechanical atomic model are (1) that electron position is described in terms of a probability distribution, and (2) electron energy is quantized into both shells and subshells--each electron is characterized by four quantum numbers.

2.5 *Relative to electrons and electron states, what does each of the four quantum numbers specify?*

Solution

The n quantum number designates the electron shell.

The l quantum number designates the electron subshell.

The m_l quantum number designates the number of electron states in each electron subshell.

The m_s quantum number designates the spin moment on each electron.

2.6 *Allowed values for the quantum numbers of electrons are as follows:*

$$n = 1, 2, 3, \ldots$$

$$l = 0, 1, 2, 3, \ldots, n-1$$

$$m_l = 0, \pm 1, \pm 2, \pm 3, \ldots, \pm l$$

$$m_s = \pm \frac{1}{2}$$

The relationships between n and the shell designations are noted in Table 2.1. Relative to the subshells,

l = 0 corresponds to an s subshell

l = 1 corresponds to a p subshell

l = 2 corresponds to a d subshell

l = 3 corresponds to an f subshell

For the K shell, the four quantum numbers for each of the two electrons in the 1s state, in the order of nlm_lm_s, *are* $100(\frac{1}{2})$ *and* $100(-\frac{1}{2})$*. Write the four quantum numbers for all of the electrons in the L and M shells, and note which correspond to the s, p, and d subshells.*

Solution

For the *L* state, $n = 2$, and eight electron states are possible. Possible l values are 0 and 1, while possible m_l values are 0 and ±1; and possible m_s values are $\pm\frac{1}{2}$. Therefore, for the *s* states, the quantum numbers are $200(\frac{1}{2})$ and $200(-\frac{1}{2})$. For the *p* states, the quantum numbers are $210(\frac{1}{2})$, $210(-\frac{1}{2})$, $211(\frac{1}{2})$, $211(-\frac{1}{2})$, $21(-1)(\frac{1}{2})$, and $21(-1)(-\frac{1}{2})$.

For the *M* state, $n = 3$, and 18 states are possible. Possible l values are 0, 1, and 2; possible m_l values are 0, ±1, and ±2; and possible m_s values are $\pm\frac{1}{2}$. Therefore, for the *s* states, the quantum numbers are $300(\frac{1}{2})$, $300(-\frac{1}{2})$, for the *p* states they are $310(\frac{1}{2})$, $310(-\frac{1}{2})$, $311(\frac{1}{2})$, $311(-\frac{1}{2})$, $31(-1)(\frac{1}{2})$, and $31(-1)(-\frac{1}{2})$; for the *d* states they are $320(\frac{1}{2})$, $320(-\frac{1}{2})$, $321(\frac{1}{2})$, $321(-\frac{1}{2})$, $32(-1)(\frac{1}{2})$, $32(-1)(-\frac{1}{2})$, $322(\frac{1}{2})$, $322(-\frac{1}{2})$, $32(-2)(\frac{1}{2})$, and $32(-2)(-\frac{1}{2})$.

2.7 *Give the electron configurations for the following ions: Fe^{2+}, Al^{3+}, Cu^{+}, Ba^{2+}, Br^{-}, and O^{2-}.*

Solution

The electron configurations for the ions are determined using Table 2.2 (and Figure 2.6).

Fe^{2+}: From Table 2.2, the electron configuration for an atom of iron is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$. In order to become an ion with a plus two charge, it must lose two electrons—in this case the two 4*s*. Thus, the electron configuration for an Fe^{2+} ion is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$.

Al^{3+}: From Table 2.2, the electron configuration for an atom of aluminum is $1s^2 2s^2 2p^6 3s^2 3p^1$. In order to become an ion with a plus three charge, it must lose three electrons—in this case two 3*s* and the one 3*p*. Thus, the electron configuration for an Al^{3+} ion is $1s^2 2s^2 2p^6$.

Cu^{+}: From Table 2.2, the electron configuration for an atom of copper is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$. In order to become an ion with a plus one charge, it must lose one electron—in this case the 4*s*. Thus, the electron configuration for a Cu^{+} ion is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$.

Ba^{2+}: The atomic number for barium is 56 (Figure 2.6), and inasmuch as it is not a transition element the electron configuration for one of its atoms is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^2$. In order to become an ion with a plus two charge, it must lose two electrons—in this case two the 6*s*. Thus, the electron configuration for a Ba^{2+} ion is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6$.

Br^{-}: From Table 2.2, the electron configuration for an atom of bromine is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$. In order to become an ion with a minus one charge, it must acquire one electron—in this case another 4*p*. Thus, the electron configuration for a Br^{-} ion is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$.

O^{2-}: From Table 2.2, the electron configuration for an atom of oxygen is $1s^2 2s^2 2p^4$. In order to become an ion with a minus two charge, it must acquire two electrons—in this case another two 2*p*. Thus, the electron configuration for an O^{2-} ion is $1s^2 2s^2 2p^6$.

2.8 *Sodium chloride (NaCl) exhibits predominantly ionic bonding. The* Na^+ *and* Cl^- *ions have electron structures that are identical to which two inert gases?*

Solution

The Na^+ ion is just a sodium atom that has lost one electron; therefore, it has an electron configuration the same as neon (Figure 2.6).

The Cl^- ion is a chlorine atom that has acquired one extra electron; therefore, it has an electron configuration the same as argon.

The Periodic Table

2.9 *With regard to electron configuration, what do all the elements in Group VIIA of the periodic table have in common?*

Solution

Each of the elements in Group VIIA has five *p* electrons.

2.10 *To what group in the periodic table would an element with atomic number 114 belong?*

Solution

From the periodic table (Figure 2.6) the element having atomic number 114 would belong to group IVA. According to Figure 2.6, Ds, having an atomic number of 110 lies below Pt in the periodic table and in the right-most column of group VIII. Moving four columns to the right puts element 114 under Pb and in group IVA.

2.11 *Without consulting Figure 2.6 or Table 2.2, determine whether each of the electron configurations given below is an inert gas, a halogen, an alkali metal, an alkaline earth metal, or a transition metal. Justify your choices.*

(a) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$

(b) $1s^2 2s^2 2p^6 3s^2 3p^6$

(c) $1s^2 2s^2 2p^5$

(d) $1s^2 2s^2 2p^6 3s^2$

(e) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$

(f) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

Solution

(a) The $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$ electron configuration is that of a transition metal because of an incomplete *d* subshell.

(b) The $1s^2 2s^2 2p^6 3s^2 3p^6$ electron configuration is that of an inert gas because of filled 3*s* and 3*p* subshells.

(c) The $1s^2 2s^2 2p^5$ electron configuration is that of a halogen because it is one electron deficient from having a filled *L* shell.

(d) The $1s^2 2s^2 2p^6 3s^2$ electron configuration is that of an alkaline earth metal because of two *s* electrons.

(e) The $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$ electron configuration is that of a transition metal because of an incomplete *d* subshell.

(f) The $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ electron configuration is that of an alkali metal because of a single *s* electron.

2.12 *(a) What electron subshell is being filled for the rare earth series of elements on the periodic table?*

(b) What electron subshell is being filled for the actinide series?

Solution

(a) The 4*f* subshell is being filled for the rare earth series of elements.

(b) The 5*f* subshell is being filled for the actinide series of elements.