Chapter 9 Solutions

6. A network solid is a lattice structure linked by covalent bonds and does not consist of discrete molecules. NaCl and $CaCO_3$ are ionic compounds and do not qualify. Au and $CuAl_2$ are metals and do not qualify. CO_2 is a covalent molecule, and does not qualify. Diamond, therefore, is the network solid consisting of a covalent network of carbons atoms.

17. Metals are malleable and ductile because the delocalized electrons do not belong to any pair of atoms; it is relatively easy, therefore, to move atoms across one another.

23. The largest number of induced dipole interactions (one atom influencing the electron density on adjacent atoms) occurs when there are more possible interactions. The larger the coordination number, the larger the number of interactions. This means either (c) or (d), both of which have a coordination number of 12.

36. In the solid state the ions are in a solid crystal lattice and are not free to move around. In a molten state, however, the ions are mobile.

37. The compounds must be ionic; thus, $MgCl_2$, KNO_3 and Ca_3P_2.

38. Any metal, or molten or dissolved ionic compound will be conducting; thus, solid NaCl is the only choice that will not conduct electricity.

58. The side of the unit cell is 0.5247 nm or 5.247×10^{-8} cm. For K, 6.022×10^{23} atoms weigh 39.098 g, or 6.49×10^{-23} g/atom; thus

$$d = 0.856\ \mathrm{g/cm^3} = \frac{\text{N atoms} \times (6.49 \times 10^{-23}\ \text{g/atom})}{(5.247 \times 10^{-8}\ \text{cm})^3}$$

Solving for N gives 1.9 or approximately 2. Two atoms per unit cell suggests body-centered cubic packing.

59. The side of the unit cell is 0.5582 nm or 5.582×10^{-8} cm. For Ca, 6.022×10^{23} atoms weigh 40.078 g, or 6.66×10^{-23} g/atom; thus

$$d = 1.55\ \mathrm{g/cm^3} = \frac{\text{N atoms} \times (6.66 \times 10^{-23}\ \text{g/atom})}{(5.582 \times 10^{-8}\ \text{cm})^3}$$

Solving for N gives 4.04 or approximately 4. Four atoms per unit cell suggests face-centered cubic packing.

64. Body-centered cubic packing means that there are 2 atoms per unit cell. Additionally, for body-centered cubic packing, the body-diagonal (opposite corners of a cube) has a length of 4r, where r is the radius of the atom. As derived in class and lab

$$4r = a\sqrt{3}$$

where a is the length of the unit cell. Knowing that r is 0.1321 nm gives a as 0.305 nm or 3.05×10^{-8} cm. The density, therefore, is

$$d = \frac{2\text{ atoms} \times (8.46 \times 10^{-23}\text{ g/atom})}{(3.05\times10^{-8}\text{ cm})^3} = 5.96\text{ g/cm}^3$$

66. Here we begin by finding the size of the unit cell; thus

$$d = 1.623\text{ g/cm}^3 = \frac{4\text{ atoms} \times (6.63 \times 10^{-23}\text{ g/atom})}{(a)^3}$$

Solving for a gives the size of the unit cell as 5.467×10^{-8} cm. For a face-centered cubic packing, as derived in class and lab, a face diagonal has the length 2r and

$$(4r)^2 = 2a^2$$

Solving for r gives the radius of a 1.93×10^{-8} cm or 0.193 nm.