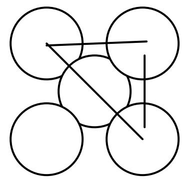
Density of Silver from the Unit Cell

Silver metal packs in the face-center cubic crystal structure. What is the density of silver metal?

Hints and helps: The atomic mass of silver is 107.87 g/mol; The atomic radius of silver is 165 pm; Avogadro's number is 6.02 x 1023 atoms/mol; Avogadro's number is also 6.02 x 10^{23} *u*/g. Density is mass/volume and it doesn't matter if it's mass/volume of a lot of silver or a unit cell.

Solution:



$$2L^2 = (4r)^2 = 16r^2$$

$$L = \sqrt{8} \cdot r$$

$$r = 165 \text{ pm} = 165 \times 10^{-12} \text{m} = 165 \times 10^{-10} \text{ cm}$$

$$V_{\text{unit cell}} = \left(\sqrt{8}\left(165 \times 10^{-10} \text{ cm}\right)\right)^3 = 1.0165 \times 10^{-22} \text{cm}^3$$

$$\text{occupancy} = 4 \text{ atoms}$$

$$\text{atomic mass} = \frac{107.87 \frac{\text{g}}{\text{mol}}}{6.02 \times 10^{23} \frac{\text{atoms}}{\text{mol}}} = 1.792 \times 10^{-22} \frac{\text{g}}{\text{atom}}$$

$$\text{mass in unit cell} = 1.792 \times 10^{-22} \frac{\text{g}}{\text{atom}} \times 4 \text{ atoms} = 7.167 \times 10^{-22} \text{g}$$

$$\text{density, } d = \frac{7.167 \times 10^{-22} \text{g}}{1.0165 \times 10^{-22} \text{cm}^3} = 7.05 \frac{\text{g}}{\text{cm}^3}$$