Calculate Density

Cr metal crystallizes as a bcc with an edge length of 2.884 Å. What is the density of Cr? (AW: 51.9961) (1 Å = 1 x 10^{-8} cm) (N_A = 6.02 x 10^{23} atoms/mol)

You want Density: g/cm^3

Start with something on the right-hand side that has one of these units in the correct place. There's two things you could start with, g/mol or uc/cm^3 (inverse of the volume of the u.c.). Dimensional analysis is the most convenient way to do this and as you go through the setup cross out things given as you use them so you know which remain.

ℓ = 2.884 Å = 2.884 x 10^{-8} cm

V = ℓ^3 = (2.884 x $10^{-8})^3$ cm^3/uc (write the unit this way)

molar mass = 51.9961 g/mol (use AW, MW or FW as molar mass)

N_A = 6.02 x 10^{23} atoms/mol

bcc ==> 2 atoms/uc

$$? \frac{\text{g}}{\text{cm}^3} = \frac{51.9961 \text{ g}}{\text{mol}} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{2 \text{ atoms}}{\text{uc}} \times \frac{1 \text{ uc}}{(2.884 \times 10^{-8})^3 \text{ cm}^3}$$

$$= 7.2\underline{\mathbf{0}}142 \text{ g/cm}^3$$

$$= 7.20 \text{ g/cm}^3$$

or

$$? \frac{\text{g}}{\text{cm}^3} = \frac{1 \text{ uc}}{(2.884 \times 10^{-8})^3 \text{ cm}^3} \times \frac{2 \text{ atoms}}{\text{uc}} \times \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}} \times \frac{51.9961 \text{ g}}{\text{mol}}$$

$$= 7.2\underline{\mathbf{0}}142 \text{ g/cm}^3$$

$$= 7.20 \text{ g/cm}^3$$

Calculate Edge Length

Ag metal has a fcc unit cell with a density of 10.6 g/cm^3. (AW: 107.8682)
(1 Å = 1 x 10^{-8} cm) (N_A = 6.02 x 10^{23} atoms/mol)

You want Volume of the u.c.: cm^3/uc (When asked for volume, edge length or radius, must get volume first)

Start with something on the right-hand side that has one of these units in the correct place. There's two things you could start with, atoms/uc or cm^3/g (inverse of the density of the u.c.). Dimensional analysis is the most convenient way to do this and as you go through the setup cross out things given as you use them so you know which remain.

density = 10.6 g/cm^3

molar mass = 107.8682 g/mol (use AW, MW or FW as molar mass)

N_A = 6.02 x 10^{23} atoms/mol

fcc ==> 4 atoms/uc

$$? \frac{\text{cm}^3}{\text{uc}} = \frac{1\ \text{cm}^3}{10.6\ \text{g}} \times \frac{107.8682\ \text{g}}{\text{mol}} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{atoms}} \times \frac{4\ \text{atoms}}{\text{uc}}$$

$= 6.7\underline{\mathbf{6}}16 \times 10^{-23}\ \text{cm}^3/\text{uc}$ (**NOTE**: The volume is a SMALL number)

$= 6.76 \times 10^{-23}\ \text{cm}^3/\text{uc}$

or

$$? \frac{\text{cm}^3}{\text{uc}} = \frac{4\ \text{atoms}}{\text{uc}} \times \frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{atoms}} \times \frac{107.8682\ \text{g}}{\text{mol}} \times \frac{1\ \text{cm}^3}{10.6\ \text{g}}$$

$= 6.7\underline{\mathbf{6}}16 \times 10^{-23}\ \text{cm}^3/\text{uc} = 6.76 \times 10^{-23}\ \text{cm}^3/\text{uc}$

Edge length: $\ell = V^{1/3} = (6.76 \times 10^{-23}\ \text{cm}^3)^{1/3} = 4.0\underline{7}3 \times 10^{-8}\ \text{cm} = 4.07$ Å

Radius: $\ell = 4r/(2)^{1/2}$ $r = \{\ell * (2)^{1/2}\}/4 = 1.44$ Å

Calculate Avogadro's Number, N_A

Li metal is bcc with an edge length of 3.509 Å and a density of 0.534 g/cm^3. Calculate Avogadro's Number, N_A. (AW: 6.941) (1 Å = 1 x 10^{-8} cm)

You want N_A: atoms/mol

Start with something on the right-hand side that has one of these units in the correct place. There's two things you could start with, g/mol or atoms/uc. Dimensional analysis is the most convenient way to do this and as you go through the setup cross out things given as you use them so you know which remain.

ℓ = 3.509 Å = 3.509 x 10^{-8} cm

V = ℓ^3 = (3.509 x 10^{-8})^3 cm^3/uc (write the unit this way)

density = 0.534 g/cm^3

molar mass = 6.941 g/mol (use AW, MW or FW as molar mass)

bcc ==> 2 atoms/uc

$$? \frac{\text{atoms}}{\text{mol}} = \frac{6.941\text{ g}}{\text{mol}} \times \frac{1\text{ cm}^3}{0.534\text{ g}} \times \frac{1\text{ uc}}{(3.509 \times 10^{-8})^3\text{ cm}^3} \times \frac{2\text{ atoms}}{\text{uc}}$$

$= 6.0\underline{\mathbf{1}}673 \times 10^{23}$ atoms/mol

$= 6.02 \times 10^{23}$ atoms/mol

or

$$? \frac{\text{atoms}}{\text{mol}} = \frac{2\text{ atoms}}{\text{uc}} \times \frac{1\text{ uc}}{(3.509 \times 10^{-8})^3\text{ cm}^3} \times \frac{1\text{ cm}^3}{0.534\text{ g}} \times \frac{6.941\text{ g}}{\text{mol}}$$

$= 6.0\underline{\mathbf{1}}673 \times 10^{23}$ atoms/mol

$= 6.02 \times 10^{23}$ atoms/mol