Solids

1. What are two broad categories of solids?

   a. Amorphous (short range order, e.g. glass)

   b. Crystalline (long range order)

2. Types of crystalline solids

   a. Ionic

   b. Molecular

   c. Atomic

   d. Covalent Network

   e. Metallic

3. What is X-ray diffraction used for?

   X-ray diffraction is used to determine the structures of crystalline solids.

4. What is the Bragg equation?

$$n\lambda = 2d \sin \theta$$

n = integer (order)
λ = wavelength used
d = distance between atoms
θ = angle of incidence/reflection

   This equation can be used to interpret the results of an diffraction experiment and determine the structure of a crystalline solid.

5. A topaz crystal has an interplanar spacing (d) of 1.36×10^{-10} m. Calculate the wavelength of the X-ray that should be used if $\theta = 15.0^{o}$

(assume n = 1).

These problems are typically just straight plug-ins. We know that we need to use the Bragg equation because we are dealing with X-ray diffraction.

We will start by organizing our data:

$d = 1.36 \times 10^{-10}$ m
$n = 1$
$\theta = 15.0^{o}$
$\lambda = ?$

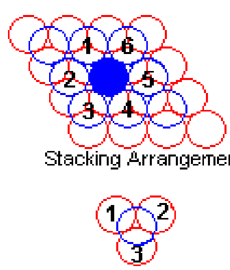
Plugging in:

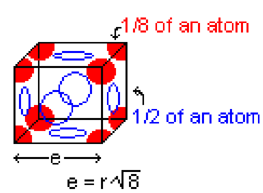
$$\lambda = \frac{2\,(1.36 \times 10^{-10}\text{ m})\,(\sin 15.0^{o})}{1} = 7.04 \times 10^{-11}\text{ m} = \boxed{70.4\text{ pm}}$$

6. What are 3 ways metals are typically arranged can be arranged?

a. Face Centered Cubic (fcc)

i. Illustrations





ii. Facts

1. Volume = e^3

r = radius of atom

2. # of nearest neighbors – 12 nearest neighbors
   3 above, 3 below and 6 on the same level.

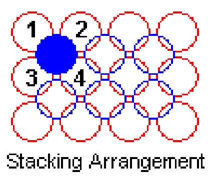
3. type of packing – Cubic Closest Packing

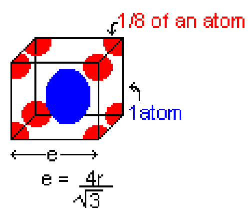
4. Total atoms within unit cell – 6(1/2) + 8(1/8) = 4 atoms

5. % of space used – 74%

b. Body Centered Cubic

i. Illustrations





ii. Facts

1. Volume = e^3

   r = radius of atom

2. # of nearest neighbors – 8 nearest neighbors
   4 above and 4 below

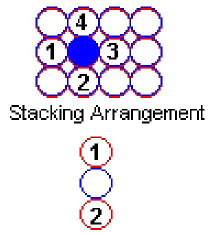
3. type of packing – You just need to know that it is *not* cubic closest packing

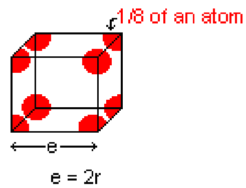
4. Total atoms within unit cell – 8(1/8) + 1 = 2 atoms

5. % of space used – 68%

c. Simple Cubic (aka Primitive)

i. Illustrations





ii. Facts

1. Volume – e^3

   r = radius of atom

2. # of nearest neighbors – 6 nearest neighbors
   1 above, 1 below and 4 on the same level.

3. Total atoms within unit cell – 8(1/8) = 1 atom

4. % of space used – 52.4%

7. A helpful formula for dealing with cubic structures and density:

$$\frac{(\text{\# of atoms in cubic})\left(\dfrac{1\text{ mol}}{6.022 \times 10^{23}\text{ atoms}}\right)(\text{molar mass of compound})}{e^3} = \text{density}$$

8. A certain form of lead has a cubic closest packed structure with an edge length of 492 pm. Calculate the value of the atomic radius and density of the lead.

   Because this is a cubic closest packed structure (meaning face centered cubic):

$$r = \frac{e}{\sqrt{8}}$$

Plugging in:

$$r = \frac{492\text{ pm}}{\sqrt{8}} = 174\text{ pm}$$

To solve for the density we can just plug into the formula from problem 23.

Because we are talking about a face-centered arrangement, we know that there are four total atoms per unit cell.

Plugging in:

$$\frac{(4\text{ atoms})\left(\frac{1\text{ mol}}{6.022\times 10^{23}\text{ atoms}}\right)\left(207.20\frac{\text{g}}{\text{mol}}\right)}{(492\times 10^{-10}\text{ cm})^3} = \boxed{11.6\frac{\text{g}}{\text{cm}^3}}$$

9. You are given a small bar of an unknown metal X. You find the density of the metal to be 10.5 g/cm^3. An X-ray diffraction experiment measures the edge of the face centered cubic unit cell as 4.09×10^{-10} m. Identify X.

The identity of X can be determined by calculating the molar mass. We will do this once again by using the formula established in question 23.

Once again, this is a face-centered cubic so we know that there are 4 total atoms per unit cell.

$$\frac{(4\text{ atoms})\left(\frac{1\text{ mol}}{6.022\times 10^{23}\text{ atoms}}\right)(\text{MM of X})}{(4.09\times 10^{-8}\text{ cm})^3} = 10.5\frac{\text{g}}{\text{cm}^3}$$

$$\text{MM of X} = 108\frac{\text{g}}{\text{mol}}$$

$$\rightarrow \boxed{\text{Ag}}$$