CH 222 Winter 2026:
“Solids” ***Lab*** *Instructions*

Step One:

Get a printed copy of this lab! You will need a printed (hard copy) version of pages I-6-2 through I-6-12 to complete this lab. If you do not turn in a printed copy of the lab, there will be a 2-point deduction.

Step Two:

Watch the **video introduction** for this lab **here: http://mhchem.org/y/6.htm**
The video introduction will help prepare you for the lab and assist you in completing the work before turning it in to the instructor.
Also **complete the PreLab questions** before starting the lab.

Step Three:

Bring the printed copy of the lab with you on Monday, February 9 (section L1), Wednesday, February 11 (section L2) *or* **Friday, February 13 (section L3).** During lab in room AC 2507, you will use these sheets (with the valuable instructions!) to gather data, all of which will be recorded in the printed pages below.

Step Four:

Complete the lab work and calculations on your own, then **turn it in (**pages I-6-7 through I-6-12 *only* to avoid a point penalty**) at the beginning of recitation to the instructor on Monday, February 16 (section L1), Wednesday, February 18 (section L2)** *or* **Friday, February 20 (section L3).** The graded lab will be returned to you the following week during recitation.

If you have any questions regarding this assignment, please email (mike.russell@mhcc.edu) the instructor! Good luck on this assignment!

Solids (The Crystal Structures of Solids)

Observing the **crystals** of an ordinary substance (such as table salt) using a magnifying glass, one sees many planes at right angles within the solid. This occurs in many common solids, and the regularity we see implies a deeper regularity in the arrangement of atoms or ions in the solid. The atomic nuclei are present in remarkably symmetrical arrays that continue for millions of units in three dimensions. Substances having a regular arrangement of atom-size particles in the solid are called **crystalline**, and the solid material consists of **crystals**. This lab deals with some of the simpler arrays in which atoms or ions occur in crystals and what these arrays can tell us about such properties as atomic sizes, densities of solids, and the efficiency of packing of particles.

Procedural Notes for the Crystal Structures of Solids Lab: Complete the handouts found at the end of this lab and turn it in (with all relevant work displayed on adjacent pages) to your instructor. Help on completing the "Solids" questions can be acquired in the Learning Success Center / AVID Center at MHCC.

Many crystals are unbelievably complex, and we will limit ourselves to the simplest crystals that have cubic structures. Cubic structures imply 90° angles and sides of equal length (hence, a cube.) We will also limit ourselves to the study of only one kind of system (namely metal elements), yet they will exhibit many of the interesting properties of more complicated structures.

The Simple Cubic (SC) Crystal

The simple cubic unit cell is a cube with an **edge length**, $\mathbf{d_0}$, equal to the distance from the center of one atom to the center of the next (see Figure One). The volume of the cube is equal to $(d_0)^3$, expressed as

$$\mathbf{V = (d_0)^3}$$

and is very small since d0 is on the order of 0.5 nm. Using x-ray diffraction we can measure the value of d0 easily to four significant figures. The number of atoms in a simple cubic unit cell is equal to one, for only 1/8 of each corner atom is actually inside the cell.

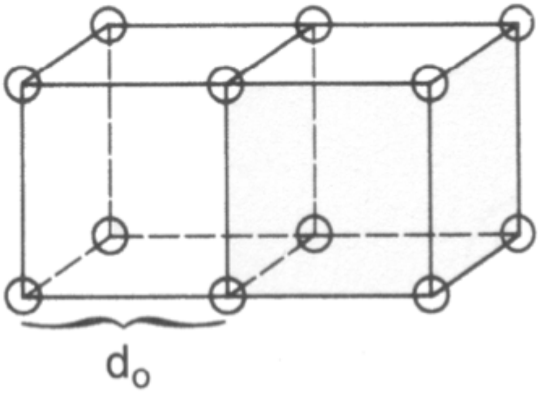


Figure One: *The Simple Cubic Crystal*

Each atom in the simple cubic unit cell is actually connected to six other atoms in the cubic lattice; hence, we say that the **coordination number** of the atoms in this structure is equal to six.

Many diagrams displaying the simple cubic unit cell show a gap between adjoining atoms. **In an actual crystal, we consider that the atoms that are closest are touching**. It is on this assumption that we determine **atomic radii, r**. In the SC crystal, if we know d_0, we can find the radius r of the atoms, since one side contains 2 atomic radii, or

$$\mathbf{d_0 = 2r}$$

for simple cubic crystals. Knowing the radius, we can calculate d0, and then we can calculate the volume of the unit cell. Knowing that one atom occupies the simple cubic cell, we can calculate the mass of the unit cell (using the molar mass and Avogadro's number), and from this we can determine the density using the volume of the cell.

Essentially no elements crystallize in the simple cubic structure, however, due to the inefficiency of the packing. The atoms in the simple cubic crystal are farther apart then they need to be, and inspection of the SC lattice will reveal a large hole in the center of the unit cell. Only about 52% of the cell volume is occupied by atoms, and more "empty space" means less stabilization for the crystal structure.

The Body Centered Cubic (BCC) Crystal

In a **body centered cubic crystal**, the unit cell still contains the corner atoms present in the SC structure, but the center of the cell now contains an additional atom. This means that every BCC crystal structure holds **two net atoms** (eight atoms are $^1/_8$ within the cell, and one whole atom within the center of the cell for two net atoms).

The edge length, d_0, can be determined using simple geometry from the **cube diagonal** (see Figure Two). The cube diagonal reaches across the cube, from an atom in the lower left front to an atom in the upper right back, or from any other appropriate combination. Geometry dictates the following relationship between the cube body diagonal and the edge length, d_0:

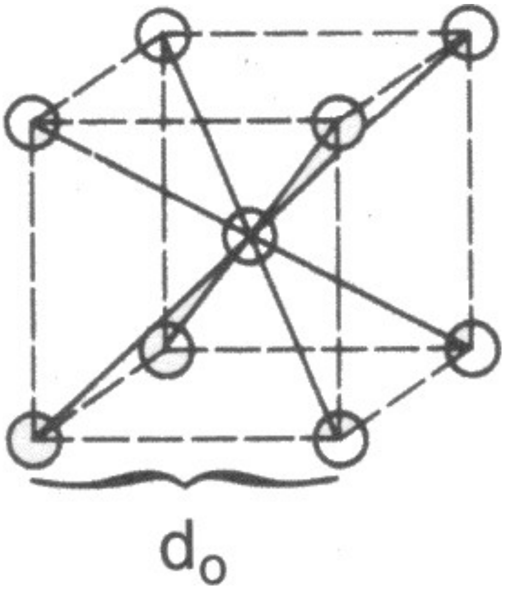


Figure Two: *Body Centered Cubic Crystal*

$$\text{cube diagonal} = \sqrt{3} \cdot d_0$$

The cube diagonal encompasses 4 radii lengths, and d_0 can be expressed in terms of the radius of the atom:

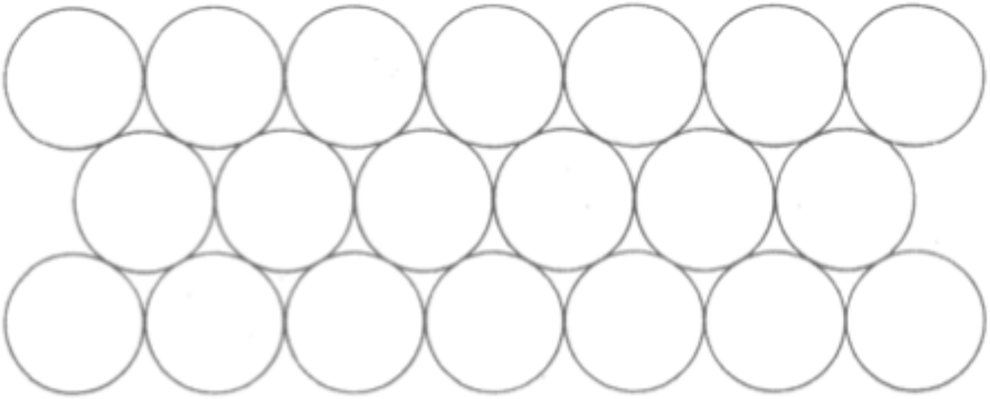
$$d_0 = \frac{4\text{r}}{\sqrt{3}}$$

The quantity d_0 can be used to find the volume of the cube; this is important for BCC cubic systems.

In a BCC lattice, each atom touches eight other atoms, and the **coordination number** is eight. The BCC lattice is much more stable than the SC structure, in part due to the higher coordination number. Many metals at room temperature display the BCC lattice, including sodium, chromium, tungsten and iron. Note that there are two atoms per unit cell in the BCC crystal. **BCC crystals are more efficient than SC crystals**, occupying approximately 68% of the total available volume.

Close Packed Structures

Although many elements prefer the BCC crystal arrangement, still more prefer structures in which the atoms are **close packed**. In close packed structures there are layers of atoms in which each atom is in contact with six others, as in the sketch below:



This is the way in which billiard balls lie in a rack or the honeycomb cells are arranged in a bees' nest. It is the most efficient way one can pack spheres, with about 74% of the volume in a close packed structure filled with atoms.

There is more than one way whereby close packed crystal structures can be stacked. One of the stacking methods is cubic and is called the **Face Centered Cubic (FCC)**. The other is called **Hexagonal Close-Packing**. We shall look at both close packed structures.

The Face Centered Cubic (FCC) Crystal

In the face centered cubic crystal unit cell there are atoms in each corner of the cell (as in the SC cell discussed earlier) and there is another atom at the center of each of the six faces. This means that FCC cubic systems consist of **four net atoms** per unit cell (eight atoms are $^1/_8$ within the cell, and six faces hold an atom which is $^1/_2$ within the cell for four net atoms). See Figure Three.

The edge length d_0 can be determined in an FCC crystal from the **face diagonal** which is defined as the distance across one face of the cube. Using geometry, we can find the edge length from the face diagonal using the following equation:

$$\text{face diagonal} = \sqrt{2} \cdot d_0$$

The face diagonal encompasses 4 radii lengths, and d_0 can be expressed in terms of the radius r:

$$d_0 = \frac{4\text{r}}{\sqrt{2}}$$

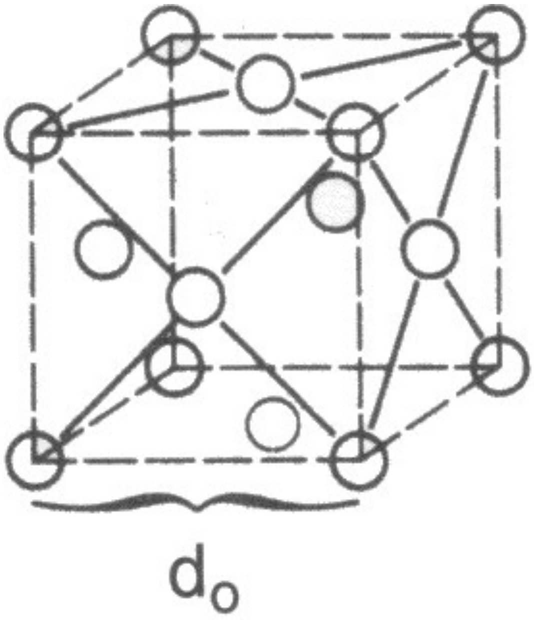


Figure Three: *Face Centered Cubic Crystal*

This expression can be used to find the volume of the cube; hence, this relationship is important for FCC cubic systems. The **coordination number** in an FCC lattice is 12, implying that FCC lattices are quite stable.

The close-packed layers of atoms in the FCC lattice are not parallel to the unit cell faces, but rather are perpendicular to the cell diagonal. If you look down the cell diagonal, you see six atoms in a close-packed triangle in the layer immediately behind the corner atom, and another layer of close-packed atoms below that, followed by another corner atom. The layers are indeed closely packed, and as one goes down the diagonal of this and succeeding cells, the layers repeat their positions in the order ABCABC…. This implies that atoms in every fourth layer lie below one another (see Figure Four (b)).

Hexagonal Close-Packing

There is another way to stack the layers as in the FCC lattice, above. The first and second layers will always be in the same relative positions, but the third layer could be below the first one if it were shifted properly. This results in a **close-packed structure** in which the order of the layers is ABABAB… (see Figure Four (a))

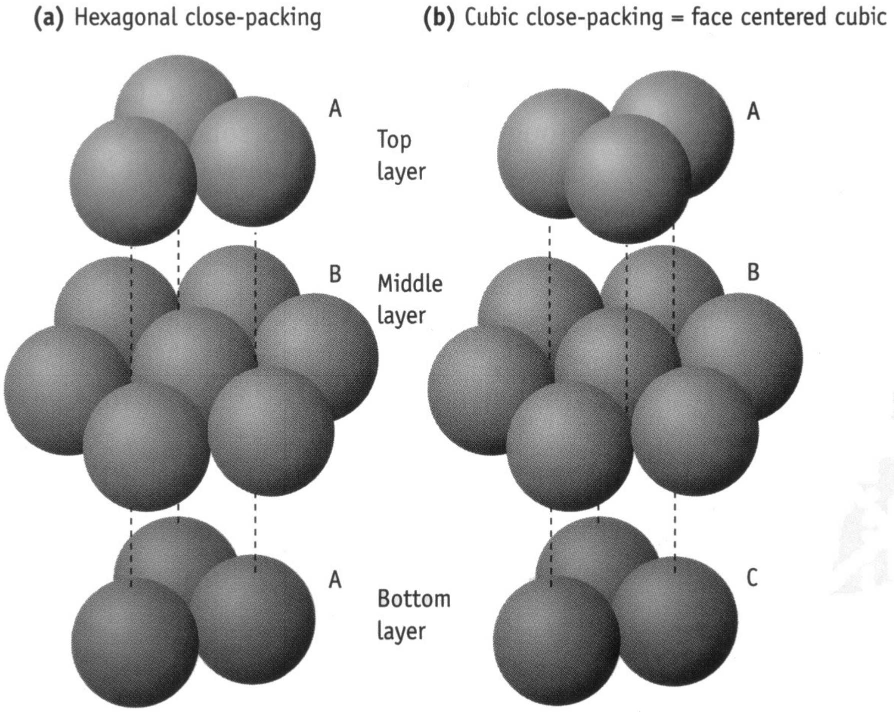


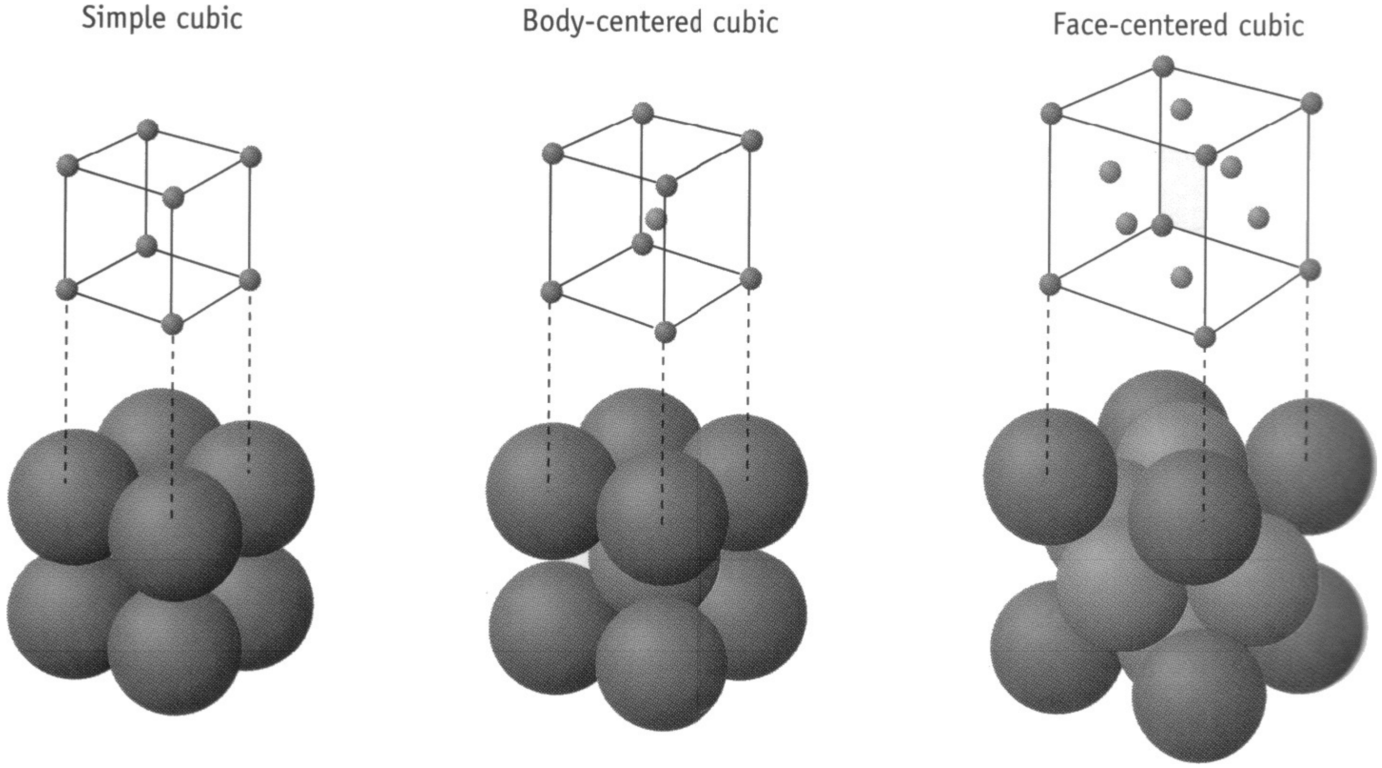
Figure Four: *Hexagonal Close Packing (left) and Cubic Close Packing (right)*

The crystal obtained from this arrangement of layers is not cubic but **hexagonal**. It is another common structure for metals. Cadmium, zinc and manganese have this structure. As you might expect, the stability of this structure is very similar to that of FCC crystals. We find that simply changing the temperature often converts a metal from one form to another. Calcium, for example, is FCC at room temperature, but if heated to 450 °C it converts to close-packed hexagonal.

In CH 222 (and CH 223), we will consider "hexagonal close-packing" structures to be identical to FCC lattices, but technically there are many differences between the two systems.

Summary of Crystal Lattice Types

Figure Five shows the three main cubic unit crystal types that we will explore in this lab. **Figure Six** show a helpful methodology to solve problems like these in this lab.


Simple cubic Body-centered cubic Face-centered cubic

Lattice Type	**Simple Cubic**	**Body Centered Cubic**	**Face Centered Cubic**
# net atoms per cell	1	2	4
d_0 (edge) in relation to r	$d_0 = 2r$	$d_0 = \frac{4r}{\sqrt{3}}$	$d_0 = \frac{4r}{\sqrt{2}}$

Figure Five: *Summary of the Three Cubic Unit Cell Types*

conversion *$V = \text{edge}^3$* *density* *molar mass (g/mol)* *Avogadro (6.022×10^{23})*

radius ↔ edge ↔ volume ↔ mass (g) ↔ moles ↔ atoms / molecules

1 pm = 10^{-12} m / 1 Å = 10^{-10} m / 1 cm = 10^{-2} m *4 atoms = 1 fcc cell, etc.*

Figure Six: *Helpful Conversion Methodology*

Solids Lab

YOUR NAME: ________________________
first and last name

Complete the five problems below and include detailed work showing all relevant calculations.

Problem 1: What element forms a face centered cubic cell, has a density of 8.92 g/cm^3, and a radius of 128 pm?

Element = __________ *Show relevant work below*

Problem 2: Chromium forms a body centered cubic crystal. If the length of an edge is 2.884 angstroms (Å), calculate the **density** (g/cm^3) and the **radius** of a chromium atom in angstroms.

density (g/cm^3) = __________________ radius (Å) = ____________

Show relevant work below.

Problem 3: Sodium (radius = 186 pm) forms a body centered cubic crystal. Calculate the **density** (g/cm^3) of sodium metal. **Propose a simple experiment to confirm your calculated density of sodium** in the lab.

density (g/cm^3) = ________________

Show relevant work below.

Proposed simple experiment:

Problem 4: Aluminum crystallizes in a face centered cubic unit cell. In addition, aluminum has an atomic radius of 143 pm. What is the density of aluminum?

density (g/cm^3) = ________________

Show relevant work below

Problem 5 *(Perform in the lab)*: Experimentally determine the density of an unknown metal solid to at least three significant figures using any equipment found in your lab drawer. Explain the process (and show *detailed* calculations) used to determine the density in three sentences or less **on this sheet**. *Hint:* use the displacement method. What liquid did you use?

density (g/cm^3) = __________________ **unknown letter used = ___________**

Relevant calculations and description of the process used to answer this question:

Solids PreLab Questions

Ideally you will complete these before performing the lab.
Include the completed PreLab Questions when you turn in your lab report.

1. What is the metric prefix for centi? (example: kilo is 10^3, etc.) _____________

2. How many centimeters are in 1 m? _____________

3. What is the metric prefix for pico?

4. How many picometers are in 1 m? _____________

5. An Angstrom (with the symbol Å) is a non-SI unit of length used often in crystallography. The Angstrom has been named after Anders Angstrom, a scientist who created an early chart of sunlight as a function of wavelength. How many Å can be found in exactly 1 m?

   answer: _____________

6. In geometry, a cube is a structure with all angles equaling 90° and all sides (s) being equal, so the volume = s^3 . If the length of the side of the cube (s) is 1.54 cm, what will be the volume of the cube (in cm^3)?

7. To convert a cube volume to a side, you will take the cubed root ($\sqrt[3]{}$) of the volume. Your calculator may have a cubed root ($\sqrt[3]{}$) button, but if it does not, you can also do a cubed root by taking the power of the number to the 1/3 power. Test this out yourself on your calculator: what is the cubed root of 27? (*Answer*: three, either through $\sqrt[3]{(27)}$ or $(27)^{1/3}$. Now try this process on this problem: **What is the length (cm) of a side on a cube with a volume of 36.64 cm^3?** *Show work.*

   Answer: _____________

8. The displacement method can be used to find the density of a solid by adding a non-dissolving / non-reactive solid to a known volume and mass of a liquid (water, etc.); the resulting difference in mass will be the mass of the solid, and the change in volume reflects the volume of the solid. If you add 18.82 g of an unknown solid to 8.50 mL of kerosene, the final volume is measured to be 15.30 mL, what is the density of the unknown solid? *Show work.*

   Answer: _____________