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Chapter 1 Solid State

Numericals

Q.1. A compound is formed by two elements X and Y. The atoms of element X form hcp lattice and those of element Y occupy $1/4^{th}$ of the tetrahedral voids. What is the formula of the compound formed?

Soln-

Let, the total number of the atoms of element packed is n.

Then, the number of tetrahedral voids formed = 2n

Hence, the number of tetrahedral voids occupied by the atoms of element Y $= 1/4 \times 2n$

Therefore, ratio of the number of atoms of X and Y = $n : 1/4 \times 2n$

$= 2 : 1$

Hence, the formula of the compound formed is X_2Y.

Q.2. A cubic solid is made of two elements X and Y. Atoms Y are at the corners of the cube and X at the body centre. What is the formula of the compound?

Soln- Since, the atom at the body centre give their entire contribution to the same unit cell, while that present at the corner is equally shared by 8 cubes and give only their 1/8th contribution to the unit cell.

Thus,

The number of atoms of element Y per unit cell = $1/8 \times 8$ (atoms at corners)

$= 1$

Number of atoms of element X per unit cell = 1 (atoms at body centre)

$= 1$

Thus, the ratio of the number of atoms of X and Y = 1 : 1

Hence, the formula of the given compound will be XY.

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Q.3. Iron has a body-centred cubic unit cell with a cell edge of 286.65 pm. The density of iron is 7.87 g cm^{-3}. Use this information to calculate Avogadro's number. (At. Mass of Fe = 56 g mol^{-1}).

Soln- In a body-centred cubic unit cell, number of atoms present = 2

Atomic mass of iron = 56 g mol^{-1}

Density of iron = 7.87 g cm^{-3}

Mass of iron = 7.87 × Volume

Volume in BCC = $(a)^3$

$= (286.65)^3$ pm

$= 2.34 \times 10^{-23}$ cm

Mass = $7.87 \times 2.34 \times 10^{-23}$ g

$$\text{Avogadro's number} = \frac{\text{No.of atoms} \times \text{Molar mass}}{\text{Given mass}}$$

$$= \frac{2 \times 56}{7.87 \times 2.34 \times 10^{-23}}$$

$$= 6.022 \times 10^{23}$$

Q.4. An element has a body-centred cubic structure with a cell edge of 288 pm. The density of the element is 7.2 g cm^{-3} Calculate the number of atoms presents in 208 g of the element.

Soln- Cell edge (a) = 288 pm

Volume of unit cell = a^3

$= (288 \text{ pm})^3$

$= (288 \times 10^{-10} \text{ cm})^3$

$= 2.389 \times 10^{-23} \text{ cm}^3$

$$\text{Volume of 208 g of the element} = \frac{Mass}{Density}$$

$$= \frac{208 \text{ g}}{7.2 \text{ g cm}^{-3}}$$

$$= 28.89 \text{ cm}^3$$

$$\text{Number of unit cells} = \frac{Total\ volume}{Volume\ of\ one\ unit\ cell}$$

$$= \frac{28.89 \text{ cm}^3}{2.389 \times 10^{-23} \text{ cm}^3}$$

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$= 12.09 \times 10^{23}$

$\because$ In a bcc structure, the number of atoms per unit cell $= 2$

$\therefore$ number of atoms in 208 g of the given element $= 2 \times 12.09 \times 10^{23}$

$= 24.18 \times 10^{23}$

Q.5. Aluminium metal forms a cubic close-packed crystal structure. Its atomic radius is 125 x 10^{-12} m.

(a) Calculate the length of the side of the unit cell.

(b) How many such unit cells are there in 1.00 m^3 of aluminium ?

Soln-

(a) For a cubic close-packed crystal structure,

Radius of an atom,

$$r = \frac{a}{2\sqrt{2}}$$

Where, a = Edge length

$$= 2\sqrt{2}r$$

$$= 2\sqrt{2} \times 125 \times 10^{-12}$$

$$= 354 \times 10^{-12}\ \text{m}$$

Hence, the length of the side of the unit cell is 354×10^{-12} m.

(b) Volume of the unit cell $= a^3$

$$= (354 \times 10^{-12}\ \text{m})^3$$

$$= 4.436 \times 10^{-29}\ \text{m}^3$$

Therefore, the number of unit cells in 1.00 m^3 of aluminium $= \frac{1}{4.436 \times 10^{-29}\ \text{m}^3}$

$$= 2.25 \times 10^{28}$$

Q.6. The density of copper metal is 8.95 g cm^{-3}. If the radius of copper atom is 127.8 pm, is the copper unit cell a simple cubic, a body-centred cubic or a face centred cubic structure? (Given: At. Mass of Cu = 63.54 g mol^{-1} and $N_A = 6.02 \times 10^{23}\ mol^{-1}$).

Soln- Given that, M = 63.54 g mol^{-1}

$N_A = 6.02 \times 10^{23}\ mol^{-1}$

$r = 127.8 \times 10^{-10}$ cm

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Since. density,

$$d = \frac{z \times M}{a^3 \times N_A}$$

For SCC- $z = 1$ and $a = 2r$

For BCC- $z = 2$ and

$$a = \frac{4}{\sqrt{2}} r$$

For FCC- $z = 4$ and

$$a = \frac{4}{\sqrt{2}} r$$

Substituting the values of z and a in equation (i) we get,

$$\text{so } d = \frac{z \times 63.54}{a^3 \times 6.02 \times 10^{23}} = \frac{z \times 10.55 \times 10^{-23}}{a^3}$$

$$\text{For SCC: } d = \frac{(1) \times 10.55 \times 10^{-23}}{(2 \times 127.8 \times 10^{-10})^3}$$

$$= \frac{10.55}{1.67} = 6.31 \text{ g/cm}^3$$

$$\text{For BCC: } d = \frac{(2) \times 10.55 \times 10^{-23}}{\left(\frac{4}{\sqrt{3}} \times 127.8 \times 10^{-10}\right)^3}$$

$$= \frac{21.10}{2.57} = 8.2 \text{ g/cm}^3$$

$$\text{For FCC: } d = \frac{4 \times 10.55 \times 10^{-23}}{\left(\frac{4}{\sqrt{2}} \times 127.8 \times 10^{-10}\right)^3}$$

$$= \frac{42.2}{4.73} = 8.92 \text{ g/cm}^3$$

The given value of density is 8.95 g/cm^3. Hence, it can be concluded that the copper unit cell is face centred cubic or fcc.

Q.7. Calculate the density of silver which crystallizes in the face-centred cubic structure. The distance between the nearest silver atoms in this structure is 287 pm.

Soln- Given that, molar mass of Ag, $M = 107.87$ g mol^{-1}

$N_A = 6.02 \times 10^{23}$ mol^{-1}

For fcc lattice, number of atoms per unit cell, $z = 4$

Distance between two nearest Ag atoms = 287 pm

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Here,

$$\therefore \quad \frac{\text{Face diagonal}}{2} = 287\ \text{pm}$$

$$\text{or} \quad \frac{\sqrt{2}a}{2} = 287\ \text{pm}$$

$$\therefore \quad a = 287 \times \sqrt{2}\ \text{pm}$$

$$\text{or} \quad a = 287 \times 1.414\ \text{pm}$$

$$\text{or} \quad a = 405.818\ \text{pm}$$

Thus, edge length = 406 pm (approx)
$= 406 \times 10^{-12}$ m

Therefore, density of silver,

$$d = \frac{z.M}{a^3.\mathrm{N_A}}$$

$$= \frac{4 \times 107.87\ \text{g mol}^{-1}}{\left(406 \times 10^{-12}\ \text{m}\right)^3 \times \left(6.02 \times 10^{23}\ \text{mol}^{-1}\right)}$$

$$= 1.07 \times 10^{7}\ \text{g m}^{-3}$$

$$= 1.07 \times 10^{4}\ \text{kg m}^{-3}$$

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