Empirical Formulae:

1) Two ways of doing this, either one will give you the correct answer. The first is to assume 100 grams (this assumed 100 grams is an exact number so it doesn't affect sig. figs.) to convert from percent to mass. The second is to simply use the mass given and multiply by the percent to convert to mass. I will show both in the first example, but I think it is best to always make the 100g assumption.

Assuming 100.g:

$$0.317\times 100.\ \boldsymbol{g} = 31.7\ \boldsymbol{g}\ \mathbf{B} \qquad 0.683\times 100\ \boldsymbol{g} = 68.3\ \boldsymbol{g}\ \mathbf{N}$$

$$31.7\ \boldsymbol{g}\ \mathbf{B}\times\frac{1\ \boldsymbol{mol}\ \mathbf{B}}{10.811\ \boldsymbol{g}\ \mathbf{B}} = 2.93\ \boldsymbol{mol}\ \mathbf{B} \qquad 68.3\ \boldsymbol{g}\ \mathbf{N}\times\frac{1\ \boldsymbol{mol}\ \mathbf{N}}{14.0067\ \boldsymbol{g}\ \mathbf{N}} = 4.88\ \boldsymbol{mol}\ \mathbf{N}$$

$$\mathbf{B} = \frac{2.93}{2.93} = 1 \qquad \mathbf{N} = \frac{4.88}{2.93} = 1.67$$

Using mass provided:

$$0.317\times 57.3\ \boldsymbol{g} = 18.16\ \boldsymbol{g}\ \mathbf{B} \qquad 0.683\times 57.3\ \boldsymbol{g} = 39.14\ \boldsymbol{g}\ \mathbf{N}$$

$$18.16\ \boldsymbol{g}\ \mathbf{B}\times\frac{1\ \boldsymbol{mol}\ \mathbf{B}}{10.811\ \boldsymbol{g}\ \mathbf{B}} = 1.680\ \boldsymbol{mol}\ \mathbf{B} \qquad 39.14\ \boldsymbol{g}\ \mathbf{N}\times\frac{1\ \boldsymbol{mol}\ \mathbf{N}}{14.0067\ \boldsymbol{g}\ \mathbf{N}} = 2.794\ \boldsymbol{mol}\ \mathbf{N}$$

$$\mathbf{B} = \frac{1.680}{1.680} = 1 \qquad \mathbf{N} = \frac{2.794}{1.680} = 1.663$$

Either way, the third step results in a decimal that cannot be rounded to the nearest whole number, so you must multiply BOTH ratios by a number (3 in this case) to get rid of the decimal. This yields the empirical formula: $\mathbf{B_3N_5}$.

2) Start by assuming that the sample is 100.g.

$$69.94\% \Rightarrow 69.94\ \boldsymbol{g}\ \mathbf{Fe} \qquad 30.06\% \Rightarrow 30.06\ \boldsymbol{g}\ \mathbf{O}$$

$$69.94\ \boldsymbol{g}\ \mathbf{Fe}\times\frac{1\ \boldsymbol{mol}\ \mathbf{Fe}}{55.847\ \boldsymbol{g}\ \mathbf{Fe}} = 1.252\ \boldsymbol{mol}\ \mathbf{Fe} \qquad 30.06\ \boldsymbol{g}\ \mathbf{O}\times\frac{1\ \boldsymbol{mol}\ \mathbf{O}}{15.9994\ \boldsymbol{g}\ \mathbf{O}} = 1.879\ \boldsymbol{mol}\ \mathbf{O}$$

$$\mathbf{Fe} = \frac{1.252}{1.252} = 1 \qquad \mathbf{O} = \frac{1.879}{1.252} = 1.50$$

Again, the third step results in a decimal that cannot be rounded to the nearest whole number, so you must multiply BOTH ratios by a number (2 in this case) to get rid of the decimal. This yields the empirical formula: $\mathbf{Fe_2O_3}$.

3) Again, assume 100.g.

$$40.0\% \Rightarrow 40.0\ \boldsymbol{g}\ \mathbf{C} \qquad 6.70\% \Rightarrow 6.70\ \boldsymbol{g}\ \mathbf{H} \qquad 53.3\% \Rightarrow 53.3\ \boldsymbol{g}\ \mathbf{O}$$

$$40.0\ \boldsymbol{g}\ \mathbf{C}\times\frac{1\ \boldsymbol{mol}\ \mathbf{C}}{12.0112\ \boldsymbol{g}\ \mathbf{C}} = 3.330\ \boldsymbol{mol}\ \mathbf{C} \qquad 6.70\ \boldsymbol{g}\ \mathbf{H}\times\frac{1\ \boldsymbol{mol}\ \mathbf{H}}{1.00797\ \boldsymbol{g}\ \mathbf{H}} = 6.647\ \boldsymbol{mol}\ \mathbf{H} \qquad 53.3\ \boldsymbol{g}\ \mathbf{O}\times\frac{1\ \boldsymbol{mol}\ \mathbf{O}}{15.9994\ \boldsymbol{g}\ \mathbf{O}} = 3.331\ \boldsymbol{mol}\ \mathbf{O}$$

$$\mathbf{C} = \frac{3.330}{3.330} = 1 \qquad \mathbf{H} = \frac{6.647}{3.330} = 1.996 \approx 2 \qquad \mathbf{O} = \frac{3.331}{3.330} \approx 1$$

No decimals here, so the empirical formula is $\mathbf{CH_2O}$

4) $\mathbf{SnF_4}$
5) $\mathbf{NH_2}$
6) $\mathbf{Na_3AlF_6}$
7) $\mathbf{C_{12}H_{12}N_2O_3}$
8) $\mathbf{CSCl_2}$
9) $\mathbf{C_3OF_6}$
10) $\mathbf{C_5H_3O}$
11) You are told that you have 73.8241 g of the sample, 40.2489 g C and 26.8055 g O, so you can find the mass of hydrogen in the sample by applying the conservation of mass: **73.8241g = 40.2489 + 26.8055 + X, X = 6.7697 g H**

$40.2489\ \boldsymbol{g}\,\mathbf{C} \times \frac{1\,\boldsymbol{mol}\ \mathbf{C}}{12.01\,\boldsymbol{g}\ \mathbf{C}} = 3.351\,\boldsymbol{mol}\ \mathbf{C}$ $6.7697\ \boldsymbol{g}\,\mathbf{H} \times \frac{1\,\boldsymbol{mol}\ \mathbf{H}}{1.01\,\boldsymbol{g}\ \mathbf{H}} = 6.84\,\boldsymbol{mol}\ \mathbf{H}$ $26.8055\ \boldsymbol{g}\,\mathbf{O} \times \frac{1\,\boldsymbol{mol}\ \mathbf{O}}{16.00\,\boldsymbol{g}\ \mathbf{O}} = 1.675\,\boldsymbol{mol}\ \mathbf{O}$

$\mathbf{C} = \frac{3.351}{1.675} \approx 2$ $\mathbf{H} = \frac{6.84}{1.675} = 4.084 \approx 4$ $\mathbf{O} = \frac{1.675}{1.675} = 1$

No decimals here, so the empirical formula is $\mathbf{C_2H_4O}$

Molecular Formulae:

1) First find the molar mass of the empirical formula

B_3N_5 = 102.467 g/mol

Then divide the molar mass (given as 204.93 g/mol) by the empirical mass.

$$\frac{204.93}{102.467} = 2$$

This means that the molecular formula is twice as heavy as the empirical formula and therefore as twice as many moles of each element as the empirical formula. The molecular formula = $\mathbf{B_6N_{10}}$

2) The molar mass of the empirical formula (Fe_2O_3) is 159.692 g/mol. This is the same as the molar mass. Therefore the molecular formula is $\mathbf{Fe_2O_3}$.

3) $\mathbf{C_3H_6O_3}$
4) $\mathbf{Sn_2F_8}$
5) $\mathbf{N_2H_4}$
6) $\mathbf{Na_3AlF_6}$
7) $\mathbf{C_{36}H_{36}N_6O_9}$
8) $\mathbf{C_2S_2Cl_4}$
9) $\mathbf{C_3OF_6}$
10) $\mathbf{C_{20}H_{12}O_4}$
11) $\mathbf{C_{12}H_{24}O_6}$