Practice questions for Ch. 3

1. A hypothetical element consists of two isotopes of masses 69.95 amu and 71.95 amu with abundances of 25.7% and 74.3%, respectively. What is the average atomic mass of this element?
   A) 70.95 amu
   B) 69.95 amu
   C) 70.5 amu
   D) 71.4 amu
   E) 71.95 amu

2. Naturally occurring copper exists in two isotopic forms: $^{63}$Cu and $^{65}$Cu. The atomic mass of copper is 63.55 amu. What is the approximate natural abundance of $^{63}$Cu?
   A) 63%
   B) 90%
   C) 70%
   D) 50%
   E) 30%

3. The average mass of a carbon atom is 12.011. Assuming you were able to pick up only one carbon unit, the chances that you would randomly get one with a mass of 12.011 is
   A) 0%
   B) 0.011%
   C) about 12%
   D) 12.011%
   E) greater than 50%

4. Iron is biologically important in the transport of oxygen by red blood cells from the lungs to the various organs of the body. In the blood of an adult human, there are approximately $2.64 \times 10^{13}$ red blood cells with a total of 2.90 g of iron. On the average, how many iron atoms are present in each red blood cell? (molar mass Fe = 55.85 g/mol)
   A) $8.44 \times 10^{-10}$
   B) $1.18 \times 10^9$
   C) $3.13 \times 10^{22}$
   D) $2.64 \times 10^{13}$
   E) $6.14 \times 10^{-2}$
5. A sample of ammonia has a mass of 43.5 g. How many molecules are in this sample?
   A) 2.55 molecules
   B) $2.62 \times 10^{25}$ molecules
   C) $2.36 \times 10^{23}$ molecules
   D) $1.54 \times 10^{24}$ molecules
   E) $8.63 \times 10^{-16}$ molecules

6. Phosphorus has the molecular formula $P_4$, and sulfur has the molecular formula $S_8$. How many grams of phosphorus contain the same number of molecules as 4.23 g of sulfur?
   A) 2.04 g
   B) 0.490 g
   C) 4.08 g
   D) 4.23 g
   E) none of these

7. A given sample of a xenon fluoride compound contains molecules of a single type $\text{XeF}_n$, where $n$ is some whole number. Given that $8.06 \times 10^{20}$ molecules of $\text{XeF}_n$ weigh 0.227 g, calculate $n$.
   A) 1
   B) 6
   C) 4
   D) 3
   E) 2

8. How many atoms of hydrogen are present in 7.63 g of ammonia?
   A) $2.70 \times 10^{23}$
   B) $1.52 \times 10^{24}$
   C) $1.38 \times 10^{25}$
   D) $8.09 \times 10^{23}$
   E) $1.12 \times 10^{20}$

9. One molecule of a compound weighs $2.93 \times 10^{-22}$ g. The molar mass of this compound is:
   A) 2.06 g/mol
   B) 567 g/mol
   C) 168 g/mol
   D) 176 g/mol
   E) none of these
10. A compound is composed of element X and hydrogen. Analysis shows the compound to be 80% X by mass, with three times as many hydrogen atoms as X atoms per molecule. Which element is element X?
   A) He
   B) C
   C) F
   D) S
   E) none of these

11. A substance contains 35.0 g nitrogen, 5.05 g hydrogen, and 60.0 g of oxygen. How many grams of hydrogen are there in a 153-g sample of this substance?
   A) 7.72 g
   B) 767 g
   C) 15.4 g
   D) 5.05 g
   E) 30.3 g

12. Chlorous acid, HClO₂, contains what percent hydrogen by mass?
   A) 1.92%
   B) 25.0%
   C) 23.4%
   D) 1.47%
   E) 5.18%

13. A substance, A₂B, has the composition by mass of 60% A and 40% B. What is the composition of AB₂ by mass?
   A) 40% A, 60% B
   B) 50% A, 50% B
   C) 27% A, 73% B
   D) 33% A, 67% B
   E) none of these

14. Which of the following compounds has the same percent composition by mass as styrene, C₈H₈?
   A) acetylene, C₂H₂
   B) benzene, C₆H₆
   C) cyclobutadiene, C₄H₄
   D) α-ethyl naphthalene, C₁₂H₁₂
   E) all of these
15. Suppose you are given the percent by mass of the elements in a compound and you wish to determine the empirical formula. Which of the following is true?

A) You must convert percent by mass to relative numbers of atoms.
B) You must assume exactly 100.0 g of the compound.
C) You must divide all of the percent by mass numbers by the smallest percent by mass.
D) You cannot solve for the empirical formula without the molar mass.
E) At least two of the above (A-D) are true.

16. A hydrocarbon (a compound consisting solely of carbon and hydrogen) is found to be 85.6% carbon by mass. What is the empirical formula for this compound?

A) CH
B) CH₂
C) C₂H
D) C₃H
E) CH₄

17. The empirical formula of a group of compounds is CHCl. Lindane, a powerful insecticide, is a member of this group. The molar mass of lindane is 290.8 g/mol. How many atoms of carbon does a molecule of lindane contain?

A) 2
B) 3
C) 4
D) 6
E) 8

18. Balanced chemical equations imply which of the following?

A) Numbers of molecules are conserved in chemical change.
B) Numbers of atoms are conserved in chemical change.
C) Volume is conserved in chemical change.
D) A and B
E) B and C

19. In balancing an equation, we change the __________ to make the number of atoms on each side of the equation balance.

A) formulas of compounds in the reactants
B) coefficients of compounds
C) formulas of compounds in the products
D) subscripts of compounds
E) none of these
20. What is the coefficient for oxygen when the following equation is balanced?
\[ \text{NH}_3(g) + \text{O}_2(g) \rightarrow \text{NO}_2(g) + \text{H}_2\text{O}(g) \]
A) 3  
B) 6  
C) 7  
D) 12  
E) 14

21. Potassium forms an oxide containing 1 oxygen atom for every 2 atoms of potassium. What is the coefficient of oxygen in the balanced equation for the reaction of potassium with oxygen to form this oxide?
A) 0  
B) 1  
C) 2  
D) 3  
E) 4

22. When the equation \( \text{C}_6\text{H}_{14} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \) is balanced with the smallest set of integers, the sum of the coefficients is
A) 4  
B) 47  
C) 15  
D) 27  
E) 34

23. You heat 3.869 g of a mixture of \( \text{Fe}_3\text{O}_4 \) and \( \text{FeO} \) to form 4.141 g \( \text{Fe}_2\text{O}_3 \). The mass of oxygen reacted is
A) 0.272 g  
B) 0.476 g  
C) 1.242 g  
D) 1.000 g  
E) none of these

24. When 233.1 g of ethylene \( (\text{C}_2\text{H}_4) \) burns in oxygen to give carbon dioxide and water, how many grams of \( \text{CO}_2 \) are formed?
A) 731.4 g  
B) 365.7 g  
C) 182.9 g  
D) 8.31 g  
E) 299.4 g
25. Consider the following reaction:
\[ \text{CH}_4(g) + 4\text{Cl}_2(g) \rightarrow \text{CCl}_4(g) + 4\text{HCl}(g) \]

What mass of CCl\(_4\) is formed by the reaction of 5.14 g of methane with an excess of chlorine?

A) 12.3 g  
B) 0.54 g  
C) 791 g  
D) 49.3 g  
E) none of these

26. Nitric oxide, NO, is made from the oxidation of NH\(_3\), and the reaction is represented by the equation:
\[ 4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O} \]

What mass of O\(_2\) would be required to react completely with 6.85 g of NH\(_3\)?

A) 4.56 g O\(_2\)  
B) 10.3 g O\(_2\)  
C) 8.04 g O\(_2\)  
D) 16.1 g O\(_2\)  
E) 12.9 g O\(_2\)

27. Phosphoric acid can be prepared by reaction of sulfuric acid with “phosphate rock” according to the equation:
\[ \text{Ca}_3(\text{PO}_4)_2 + 3\text{H}_2\text{SO}_4 \rightarrow 3\text{CaSO}_4 + 2\text{H}_3\text{PO}_4 \]

Suppose the reaction is carried out starting with 129 g of Ca\(_3\)(PO\(_4\))\(_2\) and 97.4 g of H\(_2\)SO\(_4\). Which substance is the limiting reactant?

A) Ca\(_3\)(PO\(_4\))\(_2\)  
B) H\(_2\)SO\(_4\)  
C) CaSO\(_4\)  
D) H\(_3\)PO\(_4\)  
E) none of these

28. SO\(_2\) reacts with H\(_2\)S as follows:
\[ 2\text{H}_2\text{S} + \text{SO}_2 \rightarrow 3\text{S} + 2\text{H}_2\text{O} \]

When 7.50 g of H\(_2\)S reacts with 12.75 g of SO\(_2\), which statement applies?

A) 6.38 g of sulfur are formed.  
B) 10.6 g of sulfur are formed.  
C) 0.0216 moles of H\(_2\)S remain.  
D) 1.13 g of H\(_2\)S remain.  
E) SO\(_2\) is the limiting reagent.
29. A 5.95-g sample of AgNO₃ is reacted with BaCl₂ according to the equation
   \[2 \text{AgNO}_3(aq) + \text{BaCl}_2(aq) \rightarrow 2 \text{AgCl}(s) + \text{Ba(NO}_3)_2(aq)\]
to give 3.17 g of AgCl. What is the percent yield of AgCl?
   A) 45.0%
   B) 53.3%
   C) 31.6%
   D) 63.1%
   E) 100%

30. Consider the following reaction:
   \[2A + B \rightarrow 3C + D\]
3.0 mol A and 2.0 mol B react to form 4.0 mol C. What is the percent yield of this reaction?
   A) 50%
   B) 67%
   C) 75%
   D) 89%
   E) 100%

31. For the reaction \(\text{N}_2(g) + 2\text{H}_2(g) \rightarrow \text{N}_2\text{H}_4(l)\), if the percent yield for this reaction is 82.0%, what is the actual mass of hydrazine \(\text{N}_2\text{H}_4\) produced when 25.57 g of nitrogen reacts with 4.45 g of hydrogen?
   A) 35.7 g \(\text{N}_2\text{H}_4\)
   B) 24.0 g \(\text{N}_2\text{H}_4\)
   C) 30.0 g \(\text{N}_2\text{H}_4\)
   D) 29.2 g \(\text{N}_2\text{H}_4\)
   E) 28.9 g \(\text{N}_2\text{H}_4\)

32. Tellurium consists of 3 common isotopes. Half of tellurium atoms have a mass of 127, and another 0.45 of tellurium atoms weigh 128. What is the mass of the remaining isotope? The atomic mass of tellurium is 127.6.
   A) 126
   B) 127
   C) 129
   D) 130
   E) 131
Practice questions for Ch. 3
Answer Section

1. ANS: D PTS: 1 DIF: Easy REF: 3.2
   KEY: Chemistry | general chemistry | early atomic theory | atomic theory of matter | atomic weight | mass spectroscopy
   MSC: Quantitative

2. ANS: C PTS: 1 DIF: Moderate REF: 3.2
   KEY: Chemistry | general chemistry | early atomic theory | atomic theory of matter | atomic weight | mass spectroscopy
   MSC: Quantitative

3. ANS: A PTS: 1 DIF: Easy REF: 3.2
   KEY: Chemistry | general chemistry | early atomic theory | atomic theory of matter | atomic weight
   MSC: Conceptual

4. ANS: B PTS: 1 DIF: Moderate REF: 3.3
   KEY: Chemistry | general chemistry | stoichiometry | mass and moles of substance | mole
   MSC: Quantitative

5. ANS: D PTS: 1 DIF: Moderate REF: 3.4
   KEY: Chemistry | general chemistry | stoichiometry | mass and moles of substance | mole
   MSC: Quantitative

6. ANS: A PTS: 1 DIF: Moderate REF: 3.4
   KEY: Chemistry | general chemistry | stoichiometry | mass and moles of substance | mole
   MSC: Quantitative

7. ANS: E PTS: 1 DIF: Moderate REF: 3.4
   KEY: Chemistry | general chemistry | stoichiometry | mass and moles of substance | mole
   MSC: Quantitative

8. ANS: D PTS: 1 DIF: Moderate REF: 3.4
   KEY: Chemistry | general chemistry | stoichiometry | mass and moles of substance | mole
   MSC: Quantitative

9. ANS: D PTS: 1 DIF: Moderate REF: 3.4
   KEY: Chemistry | general chemistry | stoichiometry | mass and moles of substance | mole
   MSC: Quantitative

10. ANS: B PTS: 1 DIF: Difficult REF: 3.6
    KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas | mass percentage
    MSC: Conceptual

11. ANS: A PTS: 1 DIF: Moderate REF: 3.6
    KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas
    MSC: Quantitative

12. ANS: D PTS: 1 DIF: Easy REF: 3.6
    KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas
    MSC: Quantitative

13. ANS: C PTS: 1 DIF: Moderate REF: 3.6
    KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas | mass percentage
    MSC: Conceptual
28. ANS: B  PTS: 1  DIF: Moderate  REF: 3.11
   KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | limiting reactant
   MSC: Quantitative
29. ANS: D  PTS: 1  DIF: Moderate  REF: 3.11
   KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | limiting reactant
   MSC: Quantitative
30. ANS: D  PTS: 1  DIF: Moderate  REF: 3.11
   KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | limiting reactant
   MSC: Conceptual
31. ANS: B  PTS: 1  DIF: Difficult  REF: 3.11
   KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | limiting reactant
   MSC: Quantitative
32. ANS: D  PTS: 1  DIF: Moderate  REF: 3.2
   KEY: Chemistry | general chemistry | early atomic theory | atomic theory of matter | atomic weight
   MSC: Quantitative
Solutions to the Ch. 3 practice problems

1. Atomic mass = \((0.257)(69.95) + (0.743)(71.95) = 71.4\text{ u}\)
   weighted average of isotopic masses

2. We already know the weighted average of the isotopic masses = 63.55 u
   When we have only 2 isotopes making up an element, we can find the abundance of one from the other because
   \((\text{abundance})_1 + (\text{abundance})_2 = 1\) (100%)
   So if \(x = (\text{abundance})_1\), then \((\text{abundance})_2 = 1 - x\)
   Also, we are not given the exact isotopic masses, but we know that mass number is a good approximation, so we have
   \(x(63) + (1-x)(65) = 63.55\)
   \(x = 0.725 \approx 70\%\)

3. Carbon atoms either have a mass of exactly 12 u, or 13.003355 u

4. \(29.0\text{ g Fe} \times \frac{1\text{ mol Fe}}{55.85\text{ g Fe}} \times \frac{6.022 \times 10^{23}\text{ Fe atoms}}{1\text{ mol Fe}} = 3.13 \times 10^{22}\text{ Fe atoms}\)
   no. of Fe atoms per blood cell = \(3.13 \times 10^{22}\text{ Fe atoms} / 2.64 \times 10^{13}\text{ blood cells} = 1.18 \times 10^{9}\text{ Fe atoms/blood cell}\)

5. \(43.5\text{ g NH}_3 \times \frac{1\text{ mol NH}_3}{17.04\text{ g NH}_3} \times \frac{6.022 \times 10^{23}\text{ NH}_3 molecules}{1\text{ mol NH}_3} = 1.54 \times 10^{24}\text{ NH}_3 molecules\)

6. \(4.23\text{ g S}_8 \times \frac{1\text{ mol S}_8}{256\text{ g S}_8} \times \frac{1\text{ mol P}_4}{1\text{ mol S}_8} \times \frac{124\text{ g P}_4}{1\text{ mol P}_4} = 2.04\text{ g P}_4\)
   "Same number of molecules" means "same number of moles" which means a 1:1 mole ratio.
7) If we know the molar mass of $\text{XeF}_n$, we can find $n$ because

$$131.3 + n(19.00) = \text{m.m. of } \text{XeF}_n \implies \text{solve for } n$$

We have grams per some number of molecules, which we can easily convert to grams per mole.

$$\frac{0.227\text{ g XeF}_n}{8.06 \times 10^{20} \text{ XeF}_n \text{ molecules}} \times \frac{6.022 \times 10^{23} \text{ XeF}_n \text{ molecules}}{1 \text{ mol XeF}_n} = 170 \text{ g/mol}$$

$$170 = 131.3 + n(19.00) \implies n \approx 2 \implies \text{XeF}_2$$

8) 

$$\frac{7.63\text{ g NH}_3}{17.04 \text{ g NH}_3} \times \frac{1 \text{ mol NH}_3}{1 \text{ mol NH}_3} \times \frac{3 \text{ mol H atoms}}{1 \text{ mol NH}_3} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H atoms}} = 8.09 \times 10^{23} \text{ H atoms}$$

9) 

$$\frac{2.93 \times 10^{-22} \text{ g}}{1 \text{ molecule}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 176 \text{ g/mol}$$

10) Since we are given % by mass values, let's use a 100 g sample

mass of X = (0.80)(100) = 80 g

mass of H = (0.20)(100) = 20 g

moles of H = $\frac{20}{1.01} = 19.8$ moles

(moles of H) = 3 (moles of X) "three times as many H atoms as X atoms"

$\Rightarrow$ moles of X = moles of H/3 = 6.6 moles

m.m. of X = $\frac{80\text{ g}}{6.6 \text{ moles}} \approx 12 \text{. } \implies \text{Carbon}
100 grams of $A_2B$ has 60 g of $A$ and 40 g of $B$.

To make $AB_2$, we would need only half the amount of $A$, but double the amount of $B$:

\[
\begin{align*}
\text{amount of } A \text{ in } AB_2 &= 30 g, \\
\text{amount of } B \text{ in } AB_2 &= 80 g.
\end{align*}
\]

\[\Rightarrow \text{total amount of } AB_2 = 110 g\]

\[
\%	ext{ of } A \text{ in } AB_2 = \frac{30}{110} \times 100 = 27\%.
\]

\[
\%	ext{ of } B \text{ in } AB_2 = \frac{80}{110} \times 100 = 73\%.
\]

Or, using the Law of Multiple proportions for a fixed amount of $A$:

\[
A_2B \\
AB_2 \Rightarrow \text{rewrite it as } A_2B_4 \text{ (same } \% \text{ composition as } AB_2).
\]

For a fixed amount of $A$:

\[
\frac{\text{amount of } B \text{ in } A_2B_4 \text{ (that is, } AB_2)}{\text{amount of } B \text{ in } A_2B} = 4
\]

If 100 g of $A_2B$ has 60 g of $A$ and 40 g of $B$:

for the same 60 g of $A$, a sample of $AB_2$ would have

\[
4 \times 40 = 160 \text{ g of } B
\]

Total mass of the sample would be $60 + 160 = 220 g$

\[
\Rightarrow \%	ext{ of } A = \frac{60}{220} \times 100 = 27\%
\]

\[
\%	ext{ of } B = \frac{160}{220} \times 100 = 73\%.
\]
(14) No need for calculations.
Same empirical formula \(\rightarrow\) same % composition
All have the same empirical formula as \(C_8H_8\) (CH).

(16) Take 100 g sample
mass of \(C\) = 85.6 g \(\rightarrow\) moles of \(C\) = \(\frac{85.6}{12.01}\) = 7.2 mol
mass of \(H\) = 14.4 g \(\rightarrow\) moles of \(H\) = \(\frac{14.4}{1.01}\) = 14.3 mol
Divide each mol amounts by the smallest one to get a preliminary number towards whole numbers

\[
\begin{align*}
C &: \frac{7.2}{7.2} = 1 \\
H &: \frac{14.3}{7.2} \approx 2
\end{align*}
\]

Already whole numbers \(\Rightarrow\) \(CH_2\).
We need the molecular formula. Empirical formula and molar mass can give us that.

\[
\text{m.m. (empirical formula)} = 48.7 \text{ g/mol}
\]
\[
\text{m.m. (molecular formula)} = \frac{290.8}{48.7} \approx 6
\]

1 C in empirical formula \(* 6 \rightarrow 6 \text{ C in molecular formula}

\begin{align*}
\text{NH}_3(g) + O_2(g) & \rightarrow \text{NO}_2(g) + \text{H}_2\text{O}(g) \\
\text{N occurs in one compound on each side, so we start with that, and proceed with H.} \\
\text{Let's get rid of } \frac{3}{2} \text{ by multiplying the equation by 2} \\
2 \text{NH}_3(g) + 2O_2(g) & \rightarrow 2\text{NO}_2(g) + 3\text{H}_2\text{O}(g)
\end{align*}

Oxygens are still not balanced (of course multiplying everything by 2 wouldn't balance anything). We balance oxygen by adjusting the coefficient of \(O_2\), as opposed to other molecules with oxygen, because that way we won't be destroying the balance of any other atoms. Reactant side has 4 oxygens, and the product side has 7. We need to add \(\frac{3}{2}\) \(O_2\) molecules to bring the reactant side to 7 oxygens. Coefficient of \(O_2\) becomes \(2 + \frac{3}{2} = \frac{7}{2}\). Before we write it down let's multiply everything by 2 to get rid of the fraction. We have:

\[
4\text{NH}_3(g) + 7O_2(g) \rightarrow 4\text{NO}_2(g) + 6\text{H}_2\text{O}
\]
23) We don't know the composition of the Fe$_3$O$_4$ and FeO mixture, or the exact reaction that gives Fe$_2$O$_3$ (although we can figure it out, since there is only one reaction that we can balance given the reactants and products). It doesn't matter. From the Law of Conservation of Mass:

\[
\text{mass of Fe}_3\text{O}_4 \text{ and FeO mixture} + \text{(mass of O}_2\text{ reacted)} = \text{(mass of Fe}_2\text{O}_3 \text{ formed)}
\]

\[
\text{mass of O}_2 \text{ reacted} = 4.141 - 3.869 = 0.272 \text{ g}
\]

24) While not clearly stated, "burns in oxygen" implies that O$_2$ is in excess, and C$_2$H$_4$ is limiting and can be used to find the answer.

\[
\frac{233.1 \text{ g} \text{ C}_2\text{H}_4}{28.06 \text{ g} \text{ C}_2\text{H}_4} \times \frac{1 \text{ mol} \text{ C}_2\text{H}_4}{1 \text{ mol} \text{ C}_2\text{H}_4} \times \frac{2 \text{ mol} \text{ CO}_2}{1 \text{ mol} \text{ C}_2\text{H}_4} \times \frac{44.01 \text{ g} \text{ CO}_2}{1 \text{ mol} \text{ CO}_2} = 731.2 \text{ g} \text{ CO}_2
\]

It's ok to have a slight difference from the correct choice in the last digit.

27) m.m. Ca$_3$(PO$_4$)$_2$ = 310.18 g/mol

m.m. H$_2$SO$_4$ = 98.09 g/mol

Find moles of reactants present, and find the limiting reactant by dividing each by their respective coefficient in the reaction:

Ca$_3$(PO$_4$)$_2$: no. of moles = \( \frac{129}{310.18} = 0.416 \)

H$_2$SO$_4$: no. of moles = \( \frac{97.4}{98.09} = 0.993 \)

0.993 \( \frac{1}{3} \leq \) smaller; so it's the limiting reactant

29) m.m. AgNO$_3$ = 169.9 g/mol, m.m. AgCl = 143.4

Theoretical yield = 5.95 g AgNO$_3$ \( \times \) \( \frac{1 \text{ mol AgNO}_3}{169.9 \text{ g}} \) \( \times \) \( \frac{2 \text{ mol AgCl}}{2 \text{ mol AgNO}_3} \) \( \times \) \( \frac{143.4 \text{ g AgCl}}{1 \text{ mol AgCl}} \) = 5.02 g AgCl

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = 63.1\%
\]
We are already given the moles of reactants. We find the limiting reactant by dividing the mole quantities by their respective coefficients:

A: \( \frac{3.0}{2} = 1.5 \) \(<\) smaller; A is limiting

B: \( \frac{2.0}{1} = 2.0 \)

Theoretical yield is calculated using the limiting reactant's amount:

\[ 3.0 \text{ mol A} \times \frac{2 \text{ mol C}}{2 \text{ mol A}} = 4.5 \text{ mol C} \]

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{4.0}{4.5} \times 100 = 89\%
\]

\[ 127.6 = (0.50)(127) + (0.45)(128) + (0.05) \times \]

\[ \Rightarrow x = 130 \]

\[ 1 - (0.50 + 0.45) \]