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 2.64×10^{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) 8.44×10^{-10}
   B) 1.18×10^{9}
   C) 3.13×10^{22}
   D) 2.64×10^{13}
   E) 6.14×10^{-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) 2.62×10^{25} molecules
   C) 2.36×10^{23} molecules
   D) 1.54×10^{24} molecules
   E) 8.63×10^{-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 XeF_n, where n is some whole number. Given that 8.06×10^{20} molecules of 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) 2.70×10^{23}
   B) 1.52×10^{24}
   C) 1.38×10^{25}
   D) 8.09×10^{23}
   E) 1.12×10^{20}

9. One molecule of a compound weighs 2.93×10^{-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, $HClO_2$, contains what percent hydrogen by mass?
    A) 1.92%
    B) 25.0%
    C) 23.4%
    D) 1.47%
    E) 5.18%

13. A substance, A_2B, has the composition by mass of 60% A and 40% B. What is the composition of AB_2 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, C_8H_8?
    A) acetylene, C_2H_2
    B) benzene, C_6H_6
    C) cyclobutadiene, C_4H_4
    D) α-ethyl naphthalene, $C_{12}H_{12}$
    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) CH_2
    C) C_2H
    D) C_3H
    E) CH_4

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?

$$NH_3(g) + O_2(g) \rightarrow NO_2(g) + H_2O(g)$$

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 $C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$ 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 Fe_3O_4 and FeO to form 4.141 g Fe_2O_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 (C_2H_4) burns in oxygen to give carbon dioxide and water, how many grams of 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:

$$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(g) + 4HCl(g)$$

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:
$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
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:

$$Ca_3(PO_4)_2 + 3H_2SO_4 \rightarrow 3CaSO_4 + 2H_3PO_4$$

Suppose the reaction is carried out starting with 129 g of $Ca_3(PO_4)_2$ and 97.4 g of H_2SO_4. Which substance is the limiting reactant?

A) $Ca_3(PO_4)_2$
B) H_2SO_4
C) $CaSO_4$
D) H_3PO_4
E) none of these

28. SO_2 reacts with H_2S as follows:

$$2H_2S + SO_2 \rightarrow 3S + 2H_2O$$

When 7.50 g of H_2S 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_2S remain.
D) 1.13 g of H_2S remain.
E) SO_2 is the limiting reagent.

29. A 5.95-g sample of $AgNO_3$ is reacted with $BaCl_2$ according to the equation

$$2AgNO_3(aq) + BaCl_2(aq) \rightarrow 2AgCl(s) + Ba(NO_3)_2(aq)$$

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 $N_2(g) + 2H_2(g) \rightarrow N_2H_4(l)$, if the percent yield for this reaction is 82.0%, what is the actual mass of hydrazine (N_2H_4) produced when 25.57 g of nitrogen reacts with 4.45 g of hydrogen?

A) 35.7 g N_2H_4
B) 24.0 g N_2H_4
C) 30.0 g N_2H_4
D) 29.2 g N_2H_4
E) 28.9 g N_2H_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

14. ANS: E PTS: 1 DIF: Easy REF: 3.6
KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas
MSC: Conceptual
15. ANS: A PTS: 1 DIF: Easy REF: 3.7
KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas | empirical formula MSC: Conceptual
16. ANS: B PTS: 1 DIF: Easy REF: 3.7
KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas | mass percentage
MSC: Quantitative
17. ANS: D PTS: 1 DIF: Moderate REF: 3.7
KEY: Chemistry | general chemistry | stoichiometry | determining chemical formulas | molecular formula MSC: Quantitative
18. ANS: B PTS: 1 DIF: Easy REF: 3.9
KEY: Chemistry | general chemistry | early atomic theory | chemical equation | balancing chemical equation MSC: Conceptual
19. ANS: B PTS: 1 DIF: Easy REF: 3.9
KEY: Chemistry | general chemistry | early atomic theory | chemical equation | balancing chemical equation MSC: Conceptual
20. ANS: C PTS: 1 DIF: Easy REF: 3.9
KEY: Chemistry | general chemistry | early atomic theory | chemical equation | balancing chemical equation MSC: Conceptual
21. ANS: B PTS: 1 DIF: Moderate REF: 3.9
KEY: Chemistry | general chemistry | early atomic theory | chemical equation | balancing chemical equation MSC: Conceptual
22. ANS: B PTS: 1 DIF: Easy REF: 3.9
KEY: Chemistry | general chemistry | early atomic theory | chemical equation | balancing chemical equation MSC: Conceptual
23. ANS: A PTS: 1 DIF: Easy REF: 3.1
KEY: Chemistry | general chemistry | general concepts | matter | Law of Conservation of Mass
MSC: Quantitative
24. ANS: A PTS: 1 DIF: Moderate REF: 3.1
KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | amounts of substances
MSC: Quantitative
25. ANS: D PTS: 1 DIF: Easy REF: 3.1
KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | amounts of substances
MSC: Quantitative
26. ANS: D PTS: 1 DIF: Easy REF: 3.1
KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | amounts of substances
MSC: Quantitative
27. ANS: B PTS: 1 DIF: Easy REF: 3.11
KEY: Chemistry | general chemistry | stoichiometry | stoichiometry calculation | limiting reactant
MSC: Quantitative

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 questions

(1) atomic mass $= \underbrace{(0.257)(69.95) + (0.743)(71.95)}_{\text{weighted average of isotopic masses}} = 71.4\ u$

(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

$$(\text{abundance})_1 + (\text{abundance})_2 = 1 \quad (100\%)$$

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

$$x(63) + (1-x)(65) = 63.55$$
$$x = 0.725 \simeq 70\%$$

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

(4) $290\ \cancel{g\ Fe} \times \frac{1\ \cancel{mol\ Fe}}{55.85\ \cancel{g\ Fe}} \times \frac{6.022 \times 10^{23}\ Fe\ atoms}{1\ \cancel{mol\ Fe}} = 3.13 \times 10^{22}\ Fe\ atoms$

no. of Fe atoms per blood cell $= \frac{3.13 \times 10^{22}\ Fe\ atoms}{2.64 \times 10^{13}\ blood\ cells} = 1.18 \times 10^{9}\ \frac{Fe\ atoms}{blood\ cell}$

(5) $43.5\ \cancel{g\ NH_3} \times \frac{1\ \cancel{mol\ NH_3}}{17.04\ \cancel{g\ NH_3}} \times \frac{6.022 \times 10^{23}\ NH_3\ molecules}{1\ \cancel{mol\ NH_3}} = 1.54 \times 10^{24}\ NH_3\ molecules$

(6) $4.23\ g\ S_8 \times \frac{1\ mol\ S_8}{256\ g} \times \underbrace{\frac{1\ mol\ P_4}{1\ mol\ S_8}}_{} \times \frac{124\ g\ P_4}{1\ mol\ P_4} = 2.04\ g\ P_4$

"same number of molecules" means "same number of moles", which means a 1:1 mole ratio

(7) If we know the molar mass of XeF_n, we can find n because

$$131.3 + n(19.00) = \text{m.m. of } XeF_n \longrightarrow \text{solve for } n$$

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

$$\frac{0.227\text{ g } XeF_n}{8.06\times10^{20}\ \cancel{XeF_n\text{ molecules}}} \times \frac{6.022\times10^{23}\ \cancel{XeF_n\text{ molecules}}}{1\text{ mol } XeF_n} = 170.\ \text{g/mol}$$

$$170 = 131.3 + n(19.00) \longrightarrow n \simeq 2 \implies \underline{\underline{XeF_2}}$$

(8) $$7.63\ \cancel{\text{g } NH_3} \times \frac{1\text{ mol } \cancel{NH_3}}{17.04\ \cancel{\text{g } NH_3}} \times \frac{3\ \cancel{\text{mol H atoms}}}{1\ \cancel{\text{mol } NH_3}} \times \frac{6.022\times10^{23}\text{ H atoms}}{1\ \cancel{\text{mol H atoms}}} = \underline{\underline{8.09\times10^{23}}}\text{ H atoms}$$

(9) $$\frac{2.93\times10^{-22}\text{ g}}{1\text{ molecule}} \times \frac{6.022\times10^{23}\text{ molecules}}{1\text{ mol}} = \underline{\underline{176\ \text{g/mol}}}$$

(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

Since there are only two kinds of atoms, we know % H by mass if we know % X by mass

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

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

$$\implies \text{moles of X} = \frac{\text{moles of H}}{3} = 6.6\text{ moles}$$

$$\text{m.m. of X} = \frac{80\text{g}}{6.6\text{ moles}} = 12. \implies \underline{\underline{\text{Carbon}}}$$

(11) In a way, this is a simple proportionality problem encountered since elementary school.

If we add up the masses of the elements making up the substance, we can use proportions, and do it in the dimensional analysis way

mass of the analyzed sample = 35.0 + 5.05 + 60.0 = 100.05 g

$$\text{mass of H in 153-g sample} = \frac{5.05\text{ g H}}{100.05\ \cancel{\text{g substance}}} \times 153\ \cancel{\text{g substance}} = 7.72\text{ g}$$

(12) Percent-by-mass can be calculated by taking one mole of the compound (expressed in grams, i.e. its molar mass). Percent by mass of the element is found by finding the number of moles (and then the corresponding mass) of the element in one mole of the compound, using the chemical formula of the compound.

mass of one mole of $HClO_2$ = 1.008 + 35.45 + (2)(16.00) = 68.46 g

moles of H in one mole of $HClO_2$ = 1 mole

mass of H in one mole of $HClO_2$ = (1)(1.008) = 1.008 g

$$\text{mass\% of H} = \frac{1.008}{68.48} \times 100 = 1.47\%$$

Or, by dimensional analysis:

$$\underbrace{\frac{1\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol HClO}_2}} \times \frac{1.008\text{ g H}}{1\ \cancel{\text{mol H}}} \times \frac{1\ \cancel{\text{mol HClO}_2}}{68.46\text{ g HClO}_2}}_{\text{grams of H per gram of HClO}_2} = 0.0147\ \frac{\text{g H}}{\text{g HClO}_2} \xrightarrow{\times 100} 1.47\%$$

per 1 gram

per 100 g

(13) 100 grams of A_2B has 60. g of A and 40. grams of B.

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

$$\left.\begin{array}{l}\text{amount of A in } AB_2 = 30\text{ g.}\\ \text{amount of B in } AB_2 = 80\text{ g.}\end{array}\right\} \Rightarrow \text{total amount of } AB_2 = 110.\text{ g}$$

$$\%\text{ of A in } AB_2 = \frac{30}{110} \times 100 = 27\%$$

$$\%\text{ of B in } AB_2 = \frac{80}{110} \times 100 = 73\%$$

Alternatively:

Using the relationships we derived when considering the Law of Multiple Proportions, we can write:

$$\frac{\dfrac{\text{mass of A in } AB_2}{\text{mass of B in } AB_2}}{\dfrac{\text{mass of A in } A_2B}{\text{mass of B in } A_2B}} = \frac{\dfrac{\text{atoms of A in } AB_2}{\text{atoms of B in } AB_2}}{\dfrac{\text{atoms of A in } A_2B}{\text{atoms of B in } A_2B}}$$

$$\frac{\text{mass of A in } AB_2}{\text{mass of B in } AB_2} = \left(\frac{60}{40}\right)\frac{\frac{1}{2}}{\frac{2}{1}} = \frac{60}{40}\cdot\frac{1}{4} = \frac{3}{8} = \frac{\text{mass of A in 100 g } AB_2}{\text{mass of B in 100 g } AB_2}$$

$$\Rightarrow \frac{3}{8} = \frac{x}{100-x}$$ where x = mass of A in 100 g AB_2 (i.e. mass% of A)

$x = 27\text{g}$

$100 - x = 73\text{g}$

mass % of A = 27%

mass % of B = 73%

(14) No need for calculations.

Same empirical formula → same % composition

All have the same empirical formula as C_8H_8 (CH).

(16) Take 100 g sample

mass of C = 85.6 g → moles of C = $\frac{85.6}{12.01}$ = 7.2 mol

mass of H = 14.4 g → 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

$$\left.\begin{array}{l} C: \frac{7.2}{7.2} = 1 \\ H: \frac{14.3}{7.2} \simeq 2 \end{array}\right\} \text{Already whole numbers} \Rightarrow \underline{\underline{CH_2}}$$

(17) We need the molecular formula. Empirical formula and molar mass can give us that.

m.m. (empirical formula) = 48.7 g/mol

$$\frac{\text{m.m. (molecular formula)}}{\text{m.m. (empirical formula)}} = \frac{290.8}{48.7} \approx 6$$

1 C in empirical formula × 6 → $\underline{\underline{6\ C}}$ in molecular formula

(20) $NH_3(g) + O_2(g) \rightarrow NO_2(g) + H_2O(g)$

[Not stated but apparently implied: We need minimum whole-number coefficents.]

N occurs in one compound on each side, so we start with that, and proceed with H.

$NH_3(g) + O_2(g) \rightarrow NO_2(g) + \frac{3}{2} H_2O(g)$

Let's get rid of $\frac{3}{2}$ by multiplying the equation by 2

$2 NH_3(g) + 2 O_2(g) \rightarrow 2 NO_2(g) + 3 H_2O(g)$

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:

$4NH_3(g) + 7 O_2(g) \rightarrow 4 NO_2(g) + 6 H_2O$

21) The question starts by stating something that we should be able to derive from our knowledge of periodic table. Potassium (a group 1A metal) forms K^+ ion, and oxygen (a group 6A nonmetal) forms O^{2-} ion when forming a binary ionic compound. The formula is K_2O (in order to have a neutral compound)

$$K(s) + O_2(g) \longrightarrow 2K_2O(s)$$

to have enough O atoms on the product side

$$4K(s) + O_2(g) \longrightarrow 2K_2O$$

to have enough K atoms on the reactant side after setting the coefficient of K_2O to 2

It is now balanced, with O_2 still having a coefficient of 1 (unwritten)

22) $C_6H_{14} + O_2 \longrightarrow 6CO_2 + 7H_2O$

to balance C atoms; to balance H atoms

Then we balance O atoms

$$C_6H_{14} + xO_2 \longrightarrow 6CO_2 + 7H_2O$$

$$2x = (2)(6) + (1)(7)$$
$$2x = 19$$
$$x = \frac{19}{2}$$

$$C_6H_{14} + \frac{19}{2}O_2 \longrightarrow 6CO_2 + 7H_2O$$

But we need the smallest set of integer coefficients, so we must multiply every coefficient by the denominator of the fractional coefficient $\frac{15}{2}$, which is 2

$$2C_6H_{14} + 19O_2 \longrightarrow 12CO_2 + 14H_2O$$

Sum of the coefficients = 2 + 19 + 12 + 14 = 47

(23) We don't know the composition of the Fe_3O_4 and FeO mixture, or the exact reaction that gives Fe_2O_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:

(mass of Fe_3O_4 and FeO mixture) + (mass of O_2 reacted) = (mass of Fe_2O_3 formed)

mass of O_2 reacted = 4.141 − 3.869 = $\underline{\underline{0.272\ g}}$

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(24) While not clearly stated, "burns in oxygen" implies that O_2 is in excess, and C_2H_4 is limiting and can be used to find the answer.

$$\underbrace{233.1\,\cancel{g\,C_2H_4} \times \frac{1\,\cancel{mol\,C_2H_4}}{28.06\,\cancel{g\,C_2H_4}}}_{\text{mass to mole}} \times \underbrace{\frac{2\,\cancel{mol\,CO_2}}{1\,\cancel{mol\,C_2H_4}}}_{\text{using the reaction stoichiometry}} \times \frac{44.01\,g\,CO_2}{1\,\cancel{mol\,CO_2}} = \underline{\underline{731.2\,g\,CO_2}}$$

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

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(27) m.m. $Ca_3(PO_4)_2$ = 310.18 g/mol

m.m H_2SO_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 $\quad \frac{0.416}{1}$ = 0.416

$\underline{\underline{H_2SO_4}}$: no. of moles = $\frac{97.4}{98.09}$ = 0.993 $\quad \frac{0.993}{3}$ = 0.331 ← smaller; so it's $\underline{\underline{\text{the limiting reactant}}}$

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(29) m.m. $AgNO_3$ = 169.9 g/mol , m.m. AgCl = 143.4

Theoretical yield = $5.95\,g\,AgNO_3 \times \frac{1\,mol\,AgNO_3}{169.9\,g} \times \frac{2\,mol\,AgCl}{2\,mol\,AgNO_3} \times \frac{143.4\,g\,AgCl}{1\,mol\,AgCl} = 5.02\,g\,AgCl$

% yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \underline{\underline{63.1\%}}$

(25) Since $Cl_2(g)$ is excess, methane (the other reactant, CH_4) is limiting, and the mass of CCl_4 formed is determined by the mass of CH_4.

$$5.14\,\cancel{g\,CH_4} \times \frac{1\,\cancel{mol\,CH_4}}{16.04\,\cancel{g\,CH_4}} \times \underbrace{\frac{1\,\cancel{mol\,CCl_4}}{1\,\cancel{mol\,CH_4}}}_{\text{reaction stoichiometry}} \times \frac{153.81\,g\,CCl_4}{1\,\cancel{mol\,CCl_4}} = 49.3\,g\,CCl_4$$

m.m. $CH_4 = 16.04\,g/mol$

m.m. $CCl_4 = 153.81\,g/mol$

(26) $$6.85\,\cancel{g\,NH_3} \times \frac{1\,\cancel{mol\,NH_3}}{17.03\,\cancel{g\,NH_3}} \times \frac{5\,\cancel{mol\,O_2}}{4\,\cancel{mol\,NH_3}} \times \frac{32.00\,g\,O_2}{1\,\cancel{mol\,O_2}} = 16.1\,g\,O_2$$

(28)

$$2\,H_2S + SO_2 \longrightarrow 3S + 2H_2O$$

7.50 g   12.75 g

First, determine which is the limiting reactant

H_2S: $7.50\,g \times \frac{1\ mol}{34.08\ g} = 0.2201\ mol \xrightarrow{\div 2} 0.110$ ← smaller, so H_2S is limiting

(÷2 ← coeff. in rxn.)

SO_2: $12.75\,g \times \frac{1\ mol}{64.066\ g} = 0.1990\ mol \xrightarrow{\div 1} 0.1990$

(÷1 ← coeff. in rxn.)

Since H_2S is limiting, none of it will remain when the reaction is complete. This eliminates options C, D, and E

We determine the mass of sulfur formed by 7.50 g of the limiting reactant, H_2S. We already calculated the number of moles of H_2S, so we start there.

$$0.2201\ mol\ H_2S \times \underbrace{\frac{3\ mol\ S}{2\ mol\ H_2S}}_{\text{reaction stoichiometry}} \times \frac{32.065\ g\ S}{1\ mol\ S} = 10.6\ g\ S$$

(30) 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

$$3.0 \text{ mol A} \times \frac{3 \text{ mol C}}{2 \text{mol A}} = 4.5 \text{ mol C}$$

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{4.0}{4.5} \times 100 = \underline{\underline{89\%}}$$

(31) First we need to determine the limiting reactant.
So we convert reactant masses to moles and divide them by their coefficients in the reaction equation:

$$N_2: \ 25.57 \text{ g} \times \frac{1 \text{ mol}}{28.02 \text{ g}} = 0.9126 \text{ mol} / 1 \Rightarrow 0.9126 \leftarrow \text{limiting}$$

$$N_2H_4: \ 4.45 \text{ g} \times \frac{1 \text{ mol}}{2.02 \text{ g}} = 2.20 \text{ mol} / 2 \Rightarrow 1.10$$

Now, based on the amount of limiting reactant N_2 and the reaction stoichiometry (i.e. the coefficients of the product, N_2H_4, and the limiting reactant, N_2) we calculate the ideal amount that would be produced:

$$0.9126 N_2 \times \frac{1 \text{ mol } N_2H_4}{1 \text{ mol } N_2} \times \frac{32.05 \text{ g } N_2H_4}{1 \text{ mol } N_2H_4} = 29.25 \text{ g } N_2H_4$$

Using the definition $\% \text{yield} = \frac{\text{actual yield}}{\text{ideal yield}} \times 100$

we solve for the actual yield

$$\text{actual yield} = (\text{ideal yield}) \frac{(\% \text{ yield})}{100} = (29.25) \underbrace{\frac{82.0}{100}}_{0.820} = 24.0 \text{ g}$$

(32) $$127.6 = (0.50)(127) + (0.45)(128) + \underbrace{(0.05)}_{1-(0.50+0.45)} x$$

$$\Rightarrow \underline{\underline{x = 130}}$$