**Chapter 1 Fundamental Concepts of Chemistry**

**Solutions to Problems in Chapter 1**

* 1. You must commit to memory the correspondence between names and symbols of various elements, but remembering them is simplified by the fact that most English names and elemental symbols are related: (a) H; (b) He; (c) Hf; (d) N; (e) Ne; and (f) Nb.

1.2 You must commit to memory the correspondence between names and symbols of various elements, with special attention to those elements whose symbols come from Latin words rather than their English names: (a) K; (b) Pt; (c) Pu; (d) Pb; (e) Pd; and (f)P.

1.3 Associating an element’s name with its symbol requires memorization of both names and symbols. The examples in this problem all begin with the letter “A” but their names do not necessarily begin with “A” too: (a) arsenic; (b) argon; (c) aluminum; (d) americium; (e) silver; (f) gold; (g) astatine; and (h) actinium.

1.4 Associating an element’s name with its symbol requires memorization of both names and symbols. The examples in this problem all begin with the letter “B:” (a) bromine; (b) beryllium; (c) boron; (d) berkelium; (e) barium; and (f) bismuth.

1.5 To convert a molecular picture into a molecular formula, count the atoms of each type and consult (or recall) the colour scheme used for the elements. See Figure 1-3 of your textbook for the colour scheme used in this and many other texts: (a) Br2; (b) HCl; (c) C2H5I; and (d) C3H6O.

1.6 To convert a molecular picture into a molecular formula, count atoms of each type and consult (or recall) the colour scheme used for the elements. See Figure 1-3 of your textbook for the colour scheme used in this and many other texts: (a) PCl3; (b) SF4; (c) N2O4; (d) C3H8O; and (e) H2S.

1.7 In writing a chemical formula, remember to use elemental symbols and subscripts for the number of atoms: (a) CCl4; (b) H2O2; (c) P4O10; and (d) Fe2S3.

1.8 In writing a chemical formula, remember to use elemental symbols and subscripts for the number of atoms: (a) C5H12; (b) SiF4; (c) N2O5; and (d) FeCl3.

1.9 Scientific notation expresses any number as a value between 1 and 10 times a power of ten. Trailing zeros are retained only when they are significant: (a) 1.00000 × 105; (b) 1.0 × 104; (c) 4.00 × 10-4; (d) 3 × 10-4; and (e) 2.753 × 102.

1.10 Scientific notation expresses any number as a value between 1 and 10 times a power of ten. Trailing zeros are retained only when they are significant: (a) 1.75906 × 105; (b) 6.05 × 10−5; (c) 2.5000 × 106; and (d) 2.500 × 109.

1.11 To do unit conversions, multiply by a ratio that cancels the unwanted unit(s). Refer to Table 1-2 for the SI base units:

(a) 432 kg = 4.32 × 102 kg







1.12 To do unit conversions, multiply by a ratio that cancels the unwanted unit(s). Refer to Table 1-2 for the SI base units:



1.13 This is a unit-conversion problem involving summation and unusual units. First convert all masses into kg, then put them into the same power of ten and add the masses:





1.14 When values are added, they must all be in the same units and power of ten. The mass of water must first be converted to grams:

*m*(sodium chloride) = 0.0065 × 103 g

*m*(sugar) = 0.047546 × 103 g

Sum and round: (2.000 + 0.0065 + 0.047546) × 103 g = 2.054 × 103 g

1.15 The question asks for the density of water expressed in SI units (kg/m3). Begin by analyzing the given information. The mass of the container is given before and after the water has been added. Thus, the mass of water can be obtained from the difference between the masses of the filled and empty container:

*m* = 270.064 g – 93.054 g = 177.010 g H2O



Convert the units from inches to metres before computing the volume:



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*V =* π *r2h =* (3.1416)(0.0222 m)2(0.1143 m) = 1.77 × 10-4 m3

Calculate the density by dividing the mass of water by the volume:

(Round to three significant figures because the radius of the cylinder is known to only three significant figures.)

1.16 The question asks for the density of an organic liquid expressed in g/cm3. The mass of the container is given before and after the organic liquid has been added. Thus, the mass of liquid is the difference between the mass of the filled container and the empty container:

*m* = 0.4827 g – 0.4763 g = 6.4 × 10−3 g

Divide the mass of liquid by the volume to obtain density:



(The result has two significant figures because the mass of the liquid has only two significant figures.)

1.17 The question asks for a comparison of the masses of two objects of different densities and shapes. For each object, *m =* *******V*:

Au sphere: ****** = 19.3 g/cm3,  and *r* = diameter/2 = 1.00 cm

 and 





The silver cube has more mass than the gold sphere.

1.18 The question asks for a comparison of the masses of two objects of different densities and shapes. For each object, *m =* *******V*:

Zn sphere: ***ρ*** = 7.14 g/cm3, *V* = (4/3)π*r*3, and *r* = diameter/2 = 1.00 cm

*V* =(4/3)π(1.00 cm)3 = 4.19 cm3 and *m*(zinc) = 4.19 cm3(7.14 g/cm3) = 29.9 g

Al cube: *ρ* = 2.70 g/cm3, *V= l w h*, and *l=w=h=*2.00 cm

*V* = (2.00 cm)3 = 8.00 cm3 and *m*(aluminum) =8.00 cm3 (2.70 g/cm3) = 21.6 g

The zinc sphere has more mass than the aluminum cube.

1.19 Volume and mass are related through density:



1.20 Mass and volume are related through density, 



1.21 The molar mass of a naturally occurring element can be calculated by summing the product of the fractional abundance of each isotope times its isotopic molar mass:







Elemental molar mass = 0.121 g/mol + 0.024 g/mol + 39.803 g/mol = 39.948 g/mol

1.22 The molar mass of a naturally occurring element can be calculated by summing the product of the fractional abundance of each isotope times its isotopic molar mass:



Elemental molar mass = 25.80 g/mol + 0.898 g/mol + 1.400 g/mol = 28.10 g/mol

1.23 Mass–mole conversions require the use of masses in grams and molar masses in grams per mole. To determine the number of moles, convert the mass into grams and divide by the molar mass:



(a) 

(b) 

(c) 

(d) *n* = 1.46 ton

1.24 Mass–mole conversions require the use of masses in grams and molar masses in grams per mole. To determine the number of moles, convert the mass into grams and divide by the molar mass:

(a) *n* = 3.67 kg (103 g/kg)(1 mol/47.87 g) = 76.7 mol

(b) *n* = 7.9 mg (10–3 g/mg)(1 mol/40.08 g) = 2.0 × 10–4 mol

(c) *n* = 1.56 g (1 mol/101.07 g) = 1.54 × 10–2 mol

(d) *n* = 9.63 pg (10–12 g/pg)(1 mol/98.9 g) = 9.7 × 10–14 mol

1.25 To calculate the number of atoms in a mass, convert the mass to grams, divide by molar mass to obtain moles and multiply by Avogadro’s number to obtain the number of atoms:

(a)

(b) 

(c) 

(d) 

1.26 To calculate the number of atoms in a mass, convert the mass to grams, divide by molar mass to obtain moles and multiply by Avogadro’s number to obtain the number of atoms:



1.27 To calculate the molar mass of a compound, multiply each elemental molar mass by the number of atoms in the formula and sum over the elements:

(a) CCl4: *M* = 12.01 g/mol C + 4(35.45 g/mol Cl) = 153.81 g/mol

(b) K2S: *M* = 2(39.10 g/mol K) + 32.07 g/mol S = 110.27 g/mol

(c) O3: *M* = 3(16.00 g/mol O) = 48.00 g/mol

(d) LiBr: *M* = 6.94 g/mol Li + 79.90 g/mol Br = 86.84 g/mol

(e) GaAs: *M* = 69.72 g/mol Ga + 74.92 g/mol As = 144.64 g/mol

(f) AgNO3: *M* = 107.87 g/mol Ag + 14.01 g/mol N + 3(16.00 g/mol O) = 169.88 g/mol

1.28 To calculate the molar mass of a compound, multiply each elemental molar mass by the number of atoms in the formula and sum over the elements:

(a) (NH4)2CO3:

*M* = 2[(14.01 g/mol N) + 4(1.008 g/mol H)] + 12.01 g/mol C + 3(16.00 g/mol O) = 96.09 g/mol

(b) N2O: *M* = 2(14.01 g/mol N) + 16.00 g/mol O = 44.02 g/mol

(c) CaCO3: *M* = 40.08 g/mol Ca + 12.01 g/mol C + 3(16.00 g/mol O) = 100.09 g/mol

(d) NH3: *M* = 14.01 g/mol N + 3(1.008 g/mol H) = 17.03 g/mol

(e) Na2SO4:

*M* = 2(22.99 g/mol Na) + 32.07 g/mol S + 4(16.00 g/mol O) = 142.05 g/mol

(f) C4H10: *M* = 4(12.01 g/mol C) + 10(1.008 g/mol H) = 58.12 g/mol

1.29 Determine the molecular formula from the line drawing, taking into account the “missing” carbon atoms at the ends and vertices of lines and the “missing” hydrogen atoms attached to carbon atoms. To calculate the molar mass, multiply each elemental molar mass by the number of atoms in the formula and sum over the elements:

(a) Tyrosine: 9 C atoms, 7 missing H atoms, 4 shown H atoms, 1 N atom, 3 O atoms.

C9H11NO3: *M* = 9(12.01 g/mol) + 11(1.008 g/mol) + 1(14.01 g/mol) + 3(16.00 g/mol) = 181.19 g/mol

(b) Tryptophan: 11 C atoms, 8 missing H atoms, 4 shown H atoms, 2 N atoms, 2 O atoms.

C11H12N2O2: *M* =11(12.01 g/mol) + 12(1.008 g/mol) + 2(14.01 g/mol) + 2(16.00 g/mol) = 204.23 g/mol

(c) Glutamic acid: 5 C atoms, 9 H atoms, 1 N atom, 4 O atoms.;

C5H9NO4: *M* = 5(12.01 g/mol) + 9(1.008 g/mol) + 1(14.01 g/mol) + 4(16.00 g/mol) = 147.13 g/mol

(d) Lysine: 6 C atoms, 14 shown H atoms, 2 N atoms, 2 O atoms.

C6H14N2O2: *M* = 6(12.01 g/mol) + 14(1.008 g/mol) + 2(14.01 g/mol) + 2(16.00 g/mol) = 146.19 g/mol

1.30 Determine the molecular formula from the line drawing, taking into account the “missing” carbon atoms at the ends and vertices of lines and the “missing” hydrogen atoms attached to carbon atoms. To calculate the molar mass, multiply each elemental molar mass by the number of atoms in the formula and sum over the elements.

Biotin: 10 C atoms, 13 missing H atoms, 3 shown H atoms, 2 N atoms, 3 O atoms,1 S atom.

C10H16N2O3S: *M* = 10(12.01 g/mol) + 16(1.008 g/mol) + 2(14.01 g/mol) + 3(16.00 g/mol) + 1(32.07 g/mol) = 244.32 g/mol

(b) Nicotinamide: 6 C atoms, 4 missing H atoms, 2 shown H atoms, 2 N atoms, 1 O atom.

C6H6N2O: *M* = 6(12.01 g/mol) + 6(1.008 g/mol) + 2(14.01 g/mol)

+ 1(16.00 g/mol) = 122.13 g/mol

(c) Pyridoxamine: 8 C atoms, 8 missing H atoms, 4 shown H atoms, 2 N atoms, 2 O atoms.

C8H12N2O2: *M* = 8(12.01 g/mol) + 12(1.008 g/mol) + 2(14.01 g/mol)

+ 2(16.00 g/mol) = 168.20 g/mol

(d) Pantothenic acid: 9 C atoms, 13 missing H atoms, 4 shown H atoms, 1 N atom,

5 O atoms.

C9H17NO5: *M* = 9(12.01 g/mol) + 17(1.008 g/mol) + 1(14.01 g/mol)

+ 5(16.00 g/mol) = 219.24 g/mol

1.31 To calculate the number of atoms in a mass, convert the mass to grams, divide by molar mass to obtain moles, and multiply by Avogadro’s number to obtain the number of atoms:

(a) CH4: *M* = 1(12.01 g/mol) + 4(1.008 g/mol) = 16.04 g/mol



(b) Phosphorus trichloride is PCl3:

*M* = 1(30.974 g/mol) + 3(35.453 g/mol) = 137.33 g/mol



(c) C2H6O: *M* = 2(12.01 g/mol) + 6(1.008 g/mol) + 1(15.999 g/mol) = 46.07 g/mol



(d) Uranium hexafluoride is UF6:

*M* = 1(238.03 g/mol) + 6(18.998 g/mol) = 352.02 g/mol



1.32 To calculate the number of molecules in a mass, convert the mass to grams, divide by molar mass to obtain moles and multiply by Avogadro’s number to obtain the number of molecules:



1.33 To calculate the mass of some number of molecules of a substance, divide by Avogadro’s number to obtain moles and multiply by molar mass to obtain grams:

(a) CH4: *M* = 1(12.01 g/mol) + 4(1.008 g/mol) = 16.04 g/mol



(b) C9H13NO3:

*M* =9(12.01 g/mol) + 13(1.008 g/mol) + 1(14.01 g/mol) + 3(16.00g/mol) = 183.2 g/mol



(c) C55H72MgN4O5:

*M* =55(12.01 g/mol) + 72(1.008 g/mol) + 1(24.305 g/mol) + 4(14.01 g/mol) + 5(16.00 g/mol) = 893.5 g/mol



1.34 To calculate the mass of some number of molecules of a substance, divide by Avogadro’s number to obtain moles and multiply by molar mass to obtain grams:



1.35 All parts of this question involve mass–mole–number conversions. Moles and mass in grams are related through the equation, . Use Avogadro’s number to convert between number and moles. When masses are not given in grams, unit conversions must be made. The chemical formula states the number of atoms of each element per molecule of substance:

*M*vitamin A *=* 20(12.01 g/mol) + 30(1.008 g/mol) + 1(16.00 g/mol) = 286.44 g/mol





There are 30 H atoms for every molecule of vitamin A:

# atoms = (atoms/molecule)(# molecules)





1.36 All parts of this question involve mass–mole–number conversions. Moles and mass in grams are related through the equation, . Use Avogadro’s number to convert between number and moles. When masses are not given in grams, unit conversions must be made. The chemical formula states the number of atoms of each element per molecule of substance:



(An additional significant figure is carried here to avoid rounding errors later on.)



(There are only two significant figures in the final result because the mass of the compound is only known to 2 significant figures.)

1.37 The solution process is MgCl2 (*s*) → Mg2+ (*aq*) + 2 Cl− (*aq*). Each mole of solid generates 1 mole of magnesium cations and 2 moles of chloride anions:

(a) Molarity is found using the equations,  and 

*M* = 1(24.31 g/mol) + 2(35.45 g/mol) = 95.21 g/mol





Divide by the volume in litres to obtain the concentrations:

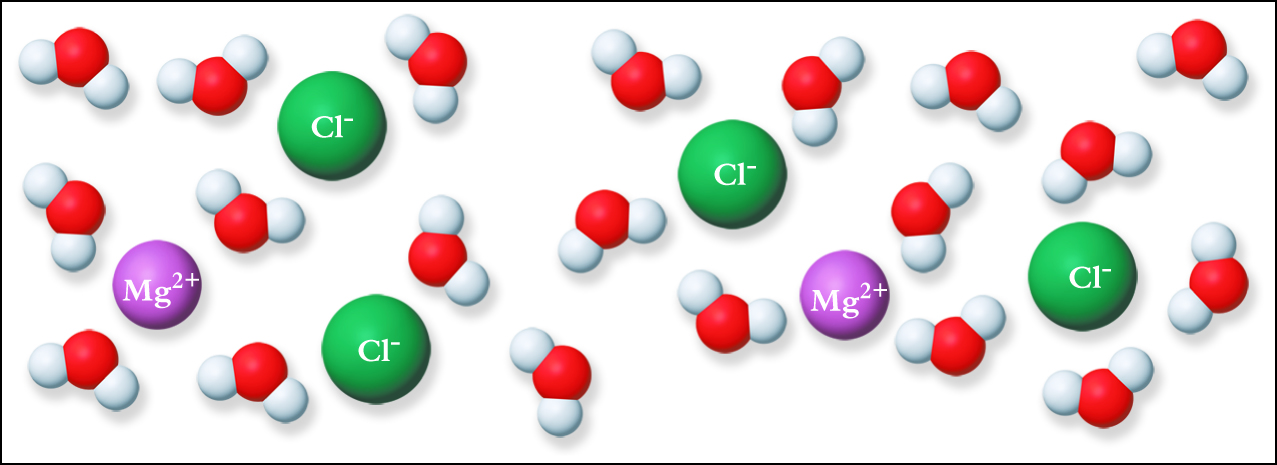






(Answers have three significant figures because the mass and volume are known to three significant figures.)

(b) Molecular pictures of solutions must illustrate the relative numbers of ions of each type present in the solution. Your picture must show twice as many chloride ions as magnesium ions. There are many more molecules of water than of either of the ions:



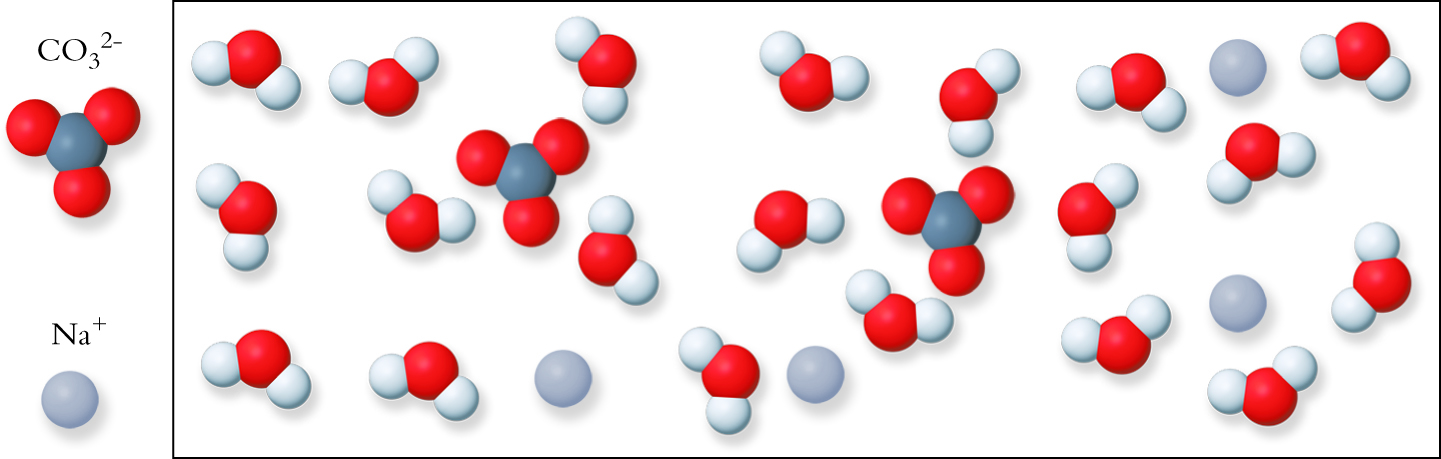
1.38 The solution process is Na2CO3 (*s*) → 2 Na+ (*aq*) + CO32− (*aq*). Each mole of solid generates 1 mole of carbonate anions and 2 moles of sodium cations:

(a) Molarity is found using the equations, 



(Answers have three significant figures because the mass and volume are known to three significant figures.)

(b) Molecular pictures of solutions must illustrate the relative numbers of ions of each type present in the solution. Your picture must show twice as many sodium ions as carbonate ions. There are many more molecules of water than of either of the ions:



1.39 The solution process is KOH (*s*) → K+ (*aq*) + OH– (*aq*). Each mole of solid generates 1 mole of each ion:

(a) Molarity is found using the equations,  and 

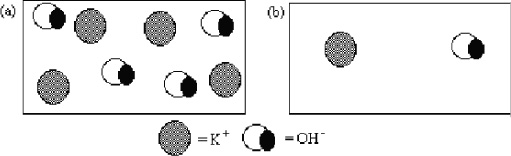
*M* = 39.098 g/mol + 15.999 g/mol + 1.0079 g/mol = 56.11 g/mol



(b) In a dilution, the number of moles of solute remains constant whereas volume increases, so *c*f*V*f*= c*i*V*i. The starting volume is 25.00 mL and the final volume is 100.00 mL:



(c) Molecular pictures of solutions must illustrate the relative numbers of ions of each type present in the solution. Because of the four-fold dilution, the solution in (b) is 1/4 as concentrated as the solution in (a). We omit solvent molecules for clarity, but remember that there are many more molecules of water than of either of the ions:



1.40 The solution process is NaCl (*s*) → Na+ (*aq*) + Cl− (*aq*). Each mole of solid generates 1 mole of each ion:

(a) Molarity is found using the equations, 

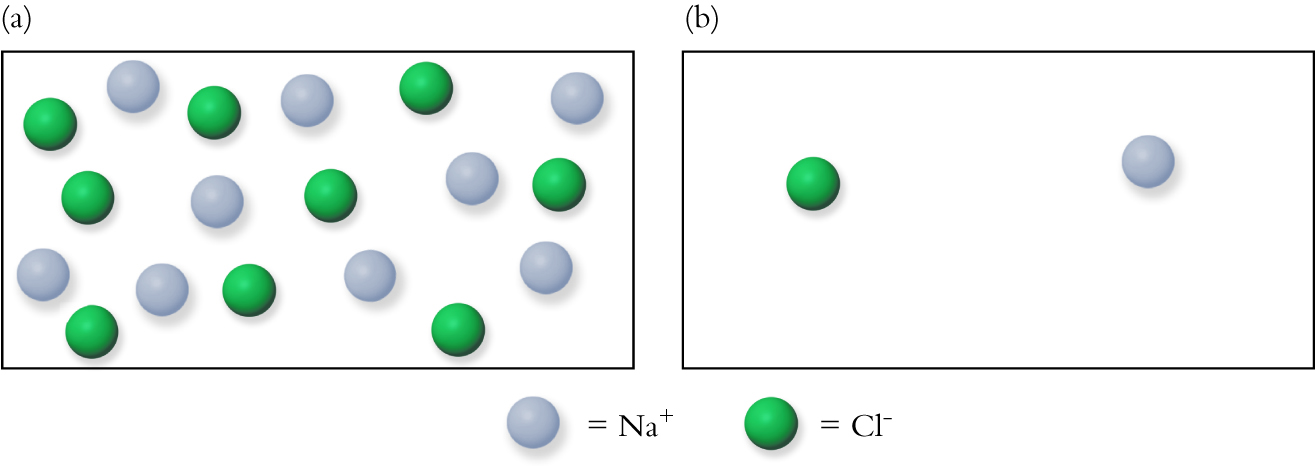
*M =* 22.900 g/mol + 35.45 g/mol = 58.44 g/mol



(b) In a dilution, the number of moles of solute remains constant whereas volume increases, so *c*f*V*f = *c*i*V*i. The starting volume is 50.0 mL and the final volume is 450. mL:



(c) Molecular pictures of solutions must illustrate the relative numbers of ions of each type present in the solution. Because of the nine-fold dilution, the solution in (b) is 1/9 as concentrated as the solution in (a). We omit solvent molecules for clarity, but remember that there are many more molecules of water than of either of the ions:



1.41 This is a dilution type problem. Rearrange the dilution equation to give an expression for the initial volume:





1.42 This is a dilution type problem. Rearrange the dilution equation to give an expression for the initial volume:



1.43 In each of the following, remember that concentration is  Begin each part by identifying the major ionic species present in solution.

(a) Na2CO3 contains a Group 1 metal ion, Na+, and a polyatomic anion, CO32−, which are the major ionic species. Determine the number of moles in 4.55 g of Na2CO3 and use stoichiometric ratios to determine the number of moles of each ion:

 = 2(22.99g/mol) + 1(12.01 g/mol) + 3(16.00g/mol) = 105.99 g/mol





There are 2 moles of Na+ per mole of Na2CO3:



Now determine the molarities by dividing the moles by the volume of the solution:







(b) NH4Cl contains a polyatomic ion and is therefore an ionic compound with two major ionic species, NH4+ and Cl−. Determine the number of moles in 27.45 mg of the salt. The stoichiometric ratio is 1:1, so the number of moles of each ion is equal to the number of moles of the salt.

 = 1(14.01 g/mol) + 4(1.008 g/mol) + 1(35.45 g/mol) = 53.49 g/mol



Determine the molarities of each ion by dividing moles by the volume of the solution:



(c) Potassium sulfate (K2SO4) contains a Group 1 metal ion, K+, and a polyatomic anion, SO42−, which are the major ionic species. Determine the number of moles in 1.85 kg of the salt and use stoichiometric ratios to determine the number of moles of each ion:

= 2(39.01 g/mol) + 1(32.06 g/mol) + 4(16.00 g/mol) = 174.08 g/mol







Obtain the molarities by dividing the moles of each ion by the volume of the solution:



Similarly to part (a), K+ has twice the concentration as the anion:



1.44 In each of the following, remember that concentration is . Begin each part by identifying the major ionic species present:

(a) Sodium hydrogen carbonate (NaHCO3) contains both a Group 1 metal ion and a polyatomic ion. The major ionic species are Na+ and HCO3−.

= 22.99 g/mol + 1.008 g/mol + 12.01 g/mol + 3(16.00g/mol) = 84.01 g/mol

The stoichiometric ratio is 1:1, so the number of moles of each ion is equal to the number of moles of the salt.



Obtain the molarities by dividing the number of moles by the volume:





(b) Major ionic species are Fe3+ and Cl–. Determine the number of moles in 1.44 mg of the salt and use stoichiometric ratios to determine the number of moles of each ion: = 55.85 g/mol + 3(35.45 g/mol) = 162.2 g/mol



Obtain the molarities by dividing the number of moles of each ion by the volume:





(c) KNO3 contains both a Group 1 metal ion and a polyatomic ion. The major ionic species are K+ and NO3−.

The stoichiometric ratio is 1:1, so the number of moles of each ion is equal to the number of moles of the salt.

*M* = 39.10 g/mol + 14.01 g/mol + 3(16.00 g/mol) = 101.1 g/mol





Determine the molarities by dividing the number of moles by the volume:



1.45 A balanced chemical equation must have equal numbers of atoms of each element on each side of the arrow. Balance each element in turn, beginning with those that appear in only one reactant and product, by adjusting stoichiometric coefficients. Generally, H and O are balanced last. In each case, we start by determining the number of atoms on each side of the chemical equation:

(a) NH4NO3 → N2O + H2O

2 N + 3 O + 4 H → 2 N + 2 O + 2 H

There are two nitrogen atoms in the reactant and two in the products, thus nitrogen is already balanced. Balance H by making the stoichiometric coefficient of water 2:

NH4NO3 → N2O + **2** H2O

2 N + 3 O + 4 H → 2 N + 3 O + 4 H (all elements balanced)

(b) P4O10 + H2O → H3PO4

4 P + 11 O + 2 H → P + 4 O + 3 H

Start by balancing P, which occurs in only one reactant and one product. There are 4 P on the reactant side and 1 P on the product side. Hence, balance P by giving H3PO4 a coefficient of 4:

P4O10 + H2O → **4** H3PO4

4 P + 11 O + 2 H → 4 P + 16 O + 12 H

Next, balance H by giving H2O a coefficient of 6, which also balances O:

P4O10 + **6** H2O → 4 H3PO4

4 P + 16 O + 12 H → 4 P + 16 O + 12 H (all elements balanced)

(c) HIO3 → I2O5 + H2O

I + H + 3 O → 2 I + 2 H + 6 O

Notice that there are half as many of each atom on the reactant side as on the product side. Therefore, this equation can be balanced by simply increasing the HIO3 coefficient from 1 to 2.

**2** HIO3 → I2O5 + H2O

2 I + 2 H + 6 O → 2 I + 2 H + 6 O (all elements balanced)

(d) As + Cl2 → AsCl5

1 As + 2 Cl → 1 As + 5 Cl

The As is already balanced. There are 2 Cl on the reactant side and 5 on the product. To balance Cl we need to change the Cl2 coefficient from 1 to :

As + Cl2 → AsCl5

1 As + 5 Cl → 1 As + 5 Cl

Notice that both As and Cl are now balanced. However, we do not want fractions in a chemical equation. Therefore, multiply **all** coefficients by 2:

**2**(As + Cl2 → AsCl5)

**2** As + **5** Cl2 → **2** AsCl5

2 As + 10 Cl → 2 As + 10 Cl (all elements balanced)

1.46 A balanced chemical equation must have equal numbers of atoms of each element on each side of the arrow. Balance each element in turn, beginning with those that appear in only one reactant and product, by adjusting stoichiometric coefficients. Generally, H and O are balanced last. In each case, start by determining the number of atoms on each side of the chemical equation:

(a) N2O5 + H2O → HNO3

2 N + 6 O + 2 H → N + 3 O + 1 H

There are two nitrogen atoms in the reactants and one in the product, so nitrogen can be balanced by increasing the HNO3 coefficient from 1 to 2:

N2O5 + H2O → 2 HNO3

2 N + 6 O + 2 H → 2 N + 6 O + 2 H

Note that this balances hydrogen and oxygen as well.

(b) KClO3 → KCl + O2

1 K + 1 Cl + 3 O → 1 K + 1 Cl + 2 O

Both K and Cl are already balanced. However, oxygen is not. Since there are 3 O on the reactant side and 2 O on the product side, change the O2 coefficient to :

KClO3 → KCl + O2

1 K + 1 Cl + 3 O → 1 K + 1 Cl + 3 O

Note that all atoms are now balanced. However, we do not want to leave fractions in a chemical equation. Therefore, multiply **all** coefficients by 2:

**2**(KClO3 → KCl +  O2)

**2** KClO3 → **2** KCl + **3** O2

(c) Fe + O2 + H2O → Fe(OH)2

1 Fe + 3 O + 2 H → 1 Fe + 2 O + 2 H

Fe and H are already balanced. Change the coefficient of O2 (which will not affect the number of Fe or H) to :

Fe +  O2 + H2O → Fe(OH)2

1 Fe + 2 O + 2 H → 1 Fe + 2 O + 2 H

All atoms are now balanced, but we do not want to leave a fraction in the equation. Therefore, multiply **all** coefficients by 2:

**2**(Fe +  O2 + H2O → Fe(OH)2)

**2** Fe + O2 + **2** H2O → **2** Fe(OH)2

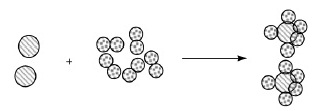
(d) P4 + Cl2 → PCl3

4 P + 2 Cl → 1 P + 3 Cl

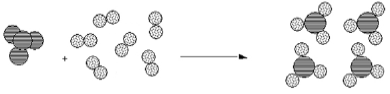
Balance P by giving PCl3 a coefficient of 4, and balance Cl by giving Cl2 a coefficient of 6:

P4 + **6** Cl2 → 4 PCl3

1.47 Molecular pictures must show the correct number of molecules undergoing the reaction. In Problem 1.45(d), two atoms of As react with five molecules of Cl2 to form two molecules of AsCl5. Remember that when drawing molecular pictures you must differentiate between the different atom types by colour, labelling, or shading (here we use shading):



1.48 Molecular pictures must show the correct number of molecules undergoing the reaction. In 1.46(d), one molecule of P4 reacts with six molecules of Cl2 to form four molecules of PCl3. Remember that when drawing molecular pictures, you must differentiate between the different atom types by colour, labelling, or shading (here we use shading):



1.49 A balanced chemical equation must have equal numbers of atoms of each element on each side of the arrow. Balance each element in turn, beginning with those that appear in only one reactant and product, by adjusting stoichiometric coefficients. Generally, H and O are balanced last. When balancing an equation, start by determining the number of atoms on each side of the chemical equation:

(a) Start by determining the chemical formula of each compound. Molecular hydrogen: H2; carbon monoxide: CO; methanol: CH3OH

The chemical equation (unbalanced) is

H2 + CO → CH3OH

2 H + 1 C + 1 O → 4 H + 1 C + 1 O

Both carbon and oxygen are balanced. There are 4 H on the product side and 2 H on the reactant. Change the coefficient of H2 from 1 to 2:

**2** H2 + CO → CH3OH

4 H + 1 C + 1 O → 4 H + 1 C + 1 O (all elements balanced)

(b) CaO + C → CO + CaC2

1 Ca + 1 O + 1 C → 1 Ca + 1 O + 3 C

Note that both Ca and O are balanced. Multiply the coefficient of C by 3:

CaO + **3** C → CO + CaC2

1 Ca + 1 O + 3 C →1 Ca + 1 O + 3 C (all elements balanced)

(c) C2H4 + O2 + HCl → C2H4Cl2 + H2O

2 C + 5 H + 2 O + Cl → 2 C + 6 H + 1 O + 2 Cl

Carbon is already balanced. Both O and Cl are found in only one reactant and one product, so we can start with either. We choose Cl. Since there is one Cl on the reactant side and 2 on the product side, change the HCl coefficient to 2:

C2H4 + O2 + **2** HCl → C2H4Cl2 + H2O

2 C + 6 H + 2 O + 2 Cl → 2 C + 6 H + 1 O + 2 Cl

This balances both Cl and H while leaving C balanced. To balance O, change the coefficient on O2 from 1 to :

C2H4 + O2 + 2 HCl → C2H4Cl2 + H2O

2 C + 6 H + 1 O + 2 Cl → 2 C + 6 H + 1 O + 2 Cl

The reaction is now balanced. Multiply **all** coefficients by 2 to eliminate fractions from the chemical equation:

**2**(C2H4 + O2 + 2 HCl → C2H4Cl2 + H2O)

**2** C2H4 + O2 + **4** HCl → **2** C2H4Cl2 + **2** H2O

1.50 A balanced chemical equation must have equal numbers of atoms of each element on each side of the arrow. Balance each element in turn, beginning with those that appear in only one reactant and product, by adjusting stoichiometric coefficients. Generally, H and O are balanced last. When balancing an equation, start by determining the number of atoms on each side of the chemical equation.

(a) NH3 + O2 →NO + H2O

1 N + 3 H + 2 O →1 N + 2 H + 2 O

To balance H, change the coefficient of H2O from 1 to 3 and NH3 from 1 to 2:

**2** NH3 + O2 →NO + **3** H2O

2 N + 6 H + 2 O →1 N + 6 H + 4 O

Next balance nitrogen by giving NO a coefficient of 2:

2 NH3 + O2 →**2** NO + 3 H2O

2 N + 6 H + 2 O →2 N + 6 H + 5 O

Finally balance O by giving O2 a coefficient of  , and multiply **all** coefficients by 2 to remove the fractions from the balanced equation:

**2**(2 NH3 + O2 →2 NO + 3 H2O)

**4** NH3 + **5** O2 →**4** NO + **6** H2O

(b) Start by determining the chemical formula of each compound:

Nitrogen oxide: NO; Molecular oxygen: O2; Nitrogen dioxide: NO2

Hence the chemical reaction is (unbalanced):

NO + O2 →NO2

1 N + 3 O →1 N + 2 O

Nitrogen is already balanced. Give O2 a coefficient of  to balance O, and multiply **all** coefficients by 2 to yield the final balanced equation:

**2**(NO + O2 →NO2)

**2** NO + O2 →**2** NO2

(c) Start by determining the chemical formula of each compound:

nitrogen dioxide: NO2; water: H2O; nitric acid: HNO3; nitrogen oxide: NO

The (unbalanced) chemical reaction is:

NO2 + H2O →HNO3 + NO

1 N + 3 O + 2 H →2 N + 4 O + 1 H

Since H only occurs in one reactant and one product, balance H first, by multiplying the nitric acid coefficient by 2:

NO2 + H2O →**2** HNO3 + NO

1 N + 3 O + 2 H →3 N + 7 O + 2 H

Now change the coefficient of NO2 from 1 to 3 to balance nitrogen:

**3** NO2 + H2O →2 HNO3 + NO

3 N + 7 O + 2 H →3 N + 7 O + 2 H

Since all elements are balanced, this is the balanced equation.

(d) NH3 + O2 →N2 + H2O

1 N + 3 H + 2 O →2 N + 2 H + 1 O

Since N is only in one reactant and one product, balance that first by multiplying the coefficient of NH3 by 2:

**2** NH3 + O2 →N2 + H2O

2 N + 6 H + 2 O →2 N + 2 H + 1 O

Hydrogen can be balanced by multiplying the water coefficient by 3:

2 NH3 + O2 →N2 + **3** H2O

2 N + 6 H + 2 O →2 N + 6 H + 3 O

Finally, balance O by multiplying the O2 coefficient by :

2 NH3 +  O2 →N2 + 3 H2O

2 N + 6 H + 3 O →2 N + 6 H + 3 O

This balances the equation; multiply **all** coefficients by 2 to yield the final balanced equation:

**2**(2 NH3 +  O2 →N2 + 3 H2O)

**4** NH3 + **3** O2 →**2** N2 + **6** H2O

(e) NH3 + NO →N2 + H2O

2 N + 3 H + 1 O → 2 N + 2 H + 1 O

To balance H, multiply NH3 by 2 and H2O by 3:

**2** NH3 + NO →N2 + **3** H2O

3 N + 6 H + 1 O →2 N + 6 H + 3 O

To balance oxygen, we need to multiply the NO coefficient by 3:

2 NH3 + **3** NO →N2 + 3 H2O

5 N + 6 H + 3 O →2 N + 6 H + 3 O

Finally, balance N by giving N2 a coefficient of  , and multiply **all** coefficients by 2 to yield the final balanced equation:

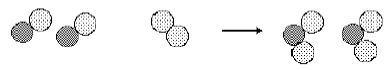
**2**(2 NH3 + 3NO →N2 + 3 H2O)

**4** NH3 + **6** NO →**5** N2 + **6** H2O

1.51 Molecular pictures must show the correct number of molecules undergoing the reaction. In Problem 1.49(a), two molecules of hydrogen react with a molecule of CO to form one molecule of CH3OH.

Olmsted_SolutionsArt_1

1.52 Molecular pictures of reactions must show the correct number of molecules undergoing reaction. In 1.50(b), two molecules of NO react with one molecule of O2 to give two molecules of NO2:



1.53 A balanced chemical equation must have equal numbers of atoms of each element on each side of the arrow. Balance each element in turn, beginning with those that appear in only one reactant and product, by adjusting stoichiometric coefficients. Start by determining the number of atoms on each side of the chemical equation.

(a) Ca(OH)2 + H3PO4 → H2O + Ca3(PO4)2

1 Ca + 6 O + 5 H + 1 P → 3 Ca + 9 O + 2 H + 2 P

Start by balancing Ca by changing the coefficient of Ca(OH)2 from 1 to 3:

**3** Ca(OH)2 + H3PO4 → H2O + Ca3(PO4)2

3 Ca + 10 O + 9 H + 1 P → 3 Ca + 9 O + 2 H + 2 P

Next, balance P by multiplying the H3PO4 coefficient by 2:

3 Ca(OH)2 + **2** H3PO4 → H2O + Ca3(PO4)2

3 Ca + 14 O + 12 H + 2 P → 3 Ca + 9 O + 2 H + 2 P

Now balance H by multiplying the water coefficient by 6

3 Ca(OH)2 + 2 H3PO4 → **6** H2O + Ca3(PO4)2

3 Ca + 14 O + 12 H + 2 P → 3 Ca + 14 O + 12 H + 2 P (all elements balanced)

(b) Na2O2 + H2O → NaOH + H2O2

2 Na + 3 O + 2 H → 1 Na + 3 O + 3 H

Start by balancing Na by changing the NaOH coefficient from 1 to 2:

Na2O2 + H2O → **2** NaOH + H2O2

2 Na + 3 O + 2 H → 2 Na + 4 O + 4 H

Since H occurs in only one reactant (O occurs in both), balance H next. Balance H by multiplying the water coefficient by 2:

Na2O2 + **2** H2O → 2 NaOH + H2O2

2 Na + 4 O + 4 H → 2 Na + 4 O + 4 H (all elements balanced)

(c) BF3 + H2O → HF + H3BO3

1 B + 3 F + 2 H + 1 O → 1 B + 1 F + 4 H + 3 O

Boron is already balanced. Balance F by multiplying the HF coefficient by 3:

BF3 + H2O → **3** HF + H3BO3

1 B + 3 F + 2 H + 1 O → 1 B + 3 F + 6 H + 3 O

Next, balance H by multiplying the water coefficient by 3:

BF3 + **3** H2O → 3 HF + H3BO3

1 B + 3 F + 6 H + 3 O → 1 B + 3 F + 6 H + 3 O (all elements balanced)

(d) NH3 + CuO → Cu + N2 + H2O

1 N + 3 H + 1 Cu + 1 O → 2 N + 2 H + 1 Cu + 1 O

Cu is balanced. To balance H, multiply the NH3 coefficient by 2 and H2O by 3:

**2** NH3 + CuO → Cu + N2 + **3** H2O

2 N + 6 H + 1 Cu + 1 O → 2 N + 6 H + 1 Cu + 3 O

This also balances N. Finally, balance O without unbalancing Cu by giving *both* CuO *and* Cu coefficients of 3:

2 NH3 + **3** CuO → **3** Cu + N2 + 3 H2O

2 N + 6 H + 3 Cu + 3 O → 2 N + 6 H + 3 Cu + 3 O

1.54 A balanced chemical equation must have equal numbers of atoms of each element on each side of the arrow. Balance each element in turn, beginning with those that appear in only one reactant and product, by adjusting stoichiometric coefficients. Start by determining the number of atoms on each side of the chemical equation.

(a) P4 + Na → Na3P

4 P + 1 Na → 1 P + 3 Na

Start by multiplying Na3P by 4 to balance phosphorus:

P4 + Na → **4** Na3P

4 P + 1 Na → 4 P + 12 Na

Now multiply Na by 12 to balance sodium:

P4 + **12** Na → 4 Na3P

4 P + 12 Na → 4 P + 12 Na (all elements balanced)

(b) Na3P + H2O → PH3 + NaOH

3 Na + P + 2 H + 1 O → 1 Na + P + 4 H + 1 O

Note that P and O are already balanced. Balance Na by multiplying the NaOH coefficient by 3:

Na3P + H2O → PH3 + **3** NaOH

3 Na + P + 2 H + 1 O → 3 Na + P + 6 H + 3 O

Next balance H by multiplying the water coefficient by 3:

Na3P + **3** H2O → PH3 + 3 NaOH

3 Na + P + 6 H + 3 O → 3 Na + P + 6 H + 3 O (all elements balanced)

(c) PH3 + O2 → P4O10 + H2O

1 P + 3 H + 2 O → 4 P + 2 H + 11 O

First balance P by changing the PH3 coefficient from 1 to 4:

**4** PH3 + O2 → P4O10 + H2O

4 P + 12 H + 2 O → 4 P + 2 H + 11 O

Now balance H by multiplying the water coefficient by 6:

4 PH3 + O2 → P4O10 + **6** H2O

4 P + 12 H + 2 O → 4 P + 12 H + 16 O

Finally, balance O by multiplying the O2 coefficient by 8:

4 PH3 + **8** O2 → P4O10 + 6 H2O

4 P + 12 H + 16 O → 4 P + 12 H + 16 O (all elements balanced)

(d) P4O10 + H2O → H3PO4

4 P + 11 O + 2 H → 1 P + 4 O + 3 H

Start by balancing P by multiplying the phosphoric acid coefficient by 4:

P4O10 + H2O → **4** H3PO4

4 P + 11 O + 2 H → 4 P + 16 O + 12 H

Now balance O and H by multiplying the water coefficient by 6 to yield the balanced equation:

P4O10 + **6** H2O → 4 H3PO4

4 P + 16 O + 12 H → 4 P + 16 O + 12 H (all elements balanced)

1.55 We must calculate the mass of the second reactant that will completely react with 5.00 g of the first. Remember that calculations of amounts in chemistry always centre on the mole. Thus, we need first to determine how many moles there are in the first reactant and then determine how many moles of the second are required to react completely:

(a) Balanced equation: 2 H2 + CO → CH3OH



(b) Balanced equation: CaO + 3 C → CO + CaC2



(c) Balanced equation: 2 C2H4 + O2 + 4 HCl → 2 C2H4Cl2 + 2 H2O



1.56 We must calculate the mass of the first reactant (in kg) that will completely react with 875 kg of the second. This is a mass–mole–mass problem. Remember that calculations of amounts in chemistry always centre on the mole:

(Note that the kg to g mass unit conversions will cancel and thus are not shown.)







1.57 You are asked to calculate the mass of sodium iodide required to produce 1.50 kg of iodine. The balanced equation is given in the problem. Remember to convert the mass into grams before dividing by the molar mass:



1.58 You are asked to calculate the mass of ammonia required to produce 3.50 metric tons of ammonium sulphate using the chemical equation: 2 NH3 + H2SO4 → (NH4)2SO4. The result is asked for in kilograms; thus, the kg to g mass unit conversions will cancel and do not need to be shown:



1.59 This problem tells you how much of the reactant you have (in kg) and asks you to calculate the amount of product formed. Remember that calculations of amounts in chemistry always centre on the mole:



1.60 This problem tells you how much of the reactant you have (in kg) and asks you to calculate the amount of product formed. Remember that calculations of amounts in chemistry always centre on the mole:



1.61 To calculate masses for a chemical reaction, balance the equation and then do appropriate mass–mole–mass conversions. The reaction is

CCl4 + HF → CCl2F2 + HCl

1 C + 4 Cl + 1 H + 1 F → 1 C + 3 Cl + 2 F + 1 H

Carbon is balanced. Fluorine occurs in only 1 reactant and 1 product, and therefore, balance it next. Fluorine can be balanced by giving HF a coefficient of 2:

CCl4 + **2** HF → CCl2F2 + HCl

1 C + 4 Cl + 2 H + 2 F → 1 C + 3 Cl + 2 F + 1 H

Finally, balance H and Cl by giving HCl a coefficient of 2:

CCl4 + 2 HF → CCl2F2 + **2** HCl

1 C + 4 Cl + 2 H + 2 F → 1 C + 4 Cl + 2 F + 2 H (all elements balanced)

To determine the amount of HF required to completely react with the CCl4, convert to moles and do the appropriate conversions:



Since this reaction is 100% efficient, all of the reactants will be converted to products. Mass of products formed:





1.62 To calculate masses for a chemical reaction, balance the equation, and then do appropriate mass–mole–mass conversions. The mass unit for starting material is metric tons, but because the results are asked for in metric tons also, the mass unit conversions will cancel and thus are not shown. The first step of the problem is to balance the chemical equation. The reaction (unbalanced) is:

SO2 + CaCO3 + O2 → CaSO4 + CO2

1 S + 7 O + 1 Ca + 1 C → 1 S + 6 O + 1 Ca + 1 C

All elements except O are already balanced. Balance O by giving O2 a coefficient of :

SO2 + CaCO3 +O2→ CaSO4 +CO2

Next, multiply through by 2 to clear fractions and obtain the balanced equation:

2 SO2 + 2 CaCO3 + O2 → 2 CaSO4 + 2CO2

Next, determine the mass of CaCO3 required for the reaction. Since we are doing a mass–mole–mass conversion, we will need the molar masses of SO2 and CaCO3.

 = 32.07 g/mol + 2(16.00 g/mol) = 64.07 g/mol SO2

 = 40.08 g/mol + 12.01 g/mol + 3(16.00 g/mol) = 100.1 g/mol CaCO3

2.00 × 104 ton SO2 

Finally, determine the mass of CaSO4 formed. Remember that the reaction is 100% efficient, thus all of the reactants will form the product:

= 40.08 g/mol + 32.07 g/mol + 4(16.00 g/mol) = 136.1 g/mol



1.63 In a yield problem, we compare actual amounts with theoretical amounts. Carry out standard mole–mass conversions to find the theoretical amount.

The reaction (balanced) is

Ca5(PO4)3F + 5 H2SO4 + 10 H2O → 3 H3PO4 + 5 CaSO4•2H2O + HF



The problem gives the actual yield (400 g), but we need to calculate the theoretical yield. Use stoichiometry to calculate the theoretical yield (FA = fluoroapatite):



Now divide the actual yield by the theoretical yield to obtain the percent yield:



1.64 This is a yield problem, where we want to compare the actual amounts with the theoretical amounts. In this problem, we are given the actual yield (60.0 g); therefore, we must carry out standard mole–mass conversions to find the theoretical amount.

The reaction (unbalanced) is HgO → Hg + O2

Note that Hg is already balanced. To balance O, give O2 a coefficient of  and then multiply through by 2 to clear fractions: 2 HgO → 2 Hg + O2

Now use stoichiometry to obtain the theoretical yield:



To obtain the percent yield, divide the actual yield by the theoretical yield:



1.65 In a multiple-step synthesis, the yield of each step must be multiplied to determine the overall yield. In this case, there are eight steps, each with a yield of 88%, so the overall fractional yield is (0.88)(0.88)(0.88)(0.88)(0.88)(0.88)(0.88)(0.88) = (0.88)8 = 0.36. Use this value and the desired amount of phenobarbital in moles to determine how many moles of toluene are required. Do the usual mole–mass conversions:

For phenobarbital, *M* = 12(12.01 g/mol) + 12(1.01g/mol) + 2(14.01 g/mol) + 3(16.00 g/mol) = 232.3 g/mol





Mass required = 299 mol (92.13 g/mol)(10–3 kg/g) = 28 kg

(The result has two significant figures because the mass and yield are only known to two significant figures.)

1.66 In a multiple-step synthesis, multiply the yields of all steps to determine the overall yield. In this case, there are three steps, each with a yield of 94.5%, so the overall fractional yield is (0.945)(0.945)(0.945) = (0.945)3 = 0.844. Use this value and the starting amount in moles to determine how many moles of product will form. Do the usual mole–mass conversions:

We are told that every mole of ammonia theoretically produces 1 mole of nitric acid.

Moles of NH3 = 7.50 × 102 kg

To determine the amount of nitric acid actually produced, multiply the fractional yield by the theoretical yield:

Moles of nitric acid produced = (4.404 × 104 mol)(0.844) = 3.716 × 104 mol

Finally, convert from moles to kilograms to obtain the mass of nitric acid produced:

Mass produced = 3.716 × 104 mol 

1.67 The calculation of a yield requires a theoretical amount and an actual amount. The theoretical amount is calculated as in Problem 1.61:



(Carry an additional significant figure until the calculation is complete.) The actual amount (given) is 105 kg.



To find how much of each reactant is required, divide the desired amount by the percent yield to obtain the theoretical yield and then do the usual stoichiometric calculations:







1.68 To calculate amounts required for a synthesis that is less than 100% efficient, first convert the actual yield into a theoretical yield and then do the usual stoichiometric calculations (adipic acid = AA):

Theoretical yield = 3.50 kg 

Amount required: 4.575 kg AA

1.69 To determine which ingredient will run out first, calculate how many cheeseburgers could be made with each ingredient:

|  |  |  |  |
| --- | --- | --- | --- |
| **Ingredient** | **Inventory** | **Amount/Burger** | **# Burgers** |
| Roll | (12 dozen)(12/dozen) | 1 | 144 |
| Beef | 40 lb | 2(1/4 lb) | 80 |
| Cheese | (2 pkg)(65/pkg) | 1 | 130 |
| Tomato | 40 | 1/4 | 160 |
| Lettuce | (1 kg)(103 g/kg) | 15 g | 66.7 |

The lettuce will run out first, after 66 burgers have been made.

1.70 To determine which type of screw will run out first, calculate how many assortments could be made with each type of screw:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Type | *m/*Screw | Screws/Box | #/Assort. | # of Assorts. |
| #8, 1/2" | 12.0 g | 454/12.0 = 38 | 10 | 2(38)/10 = 7.6 |
| #8, 1" | 21.5 | 454/21.5 = 21 | 8 | 2(21)/8 = 5.3 |
| #10, 1/2" | 14.5 | 454/14.5 = 31 | 8 | 2(31)/8 = 7.8 |
| #10, 1 ½" | 31.0 | 454/31.0 = 15 | 5 | 2(15)/5 = 6 |
| #12, 2" | 45.0 | 454/45.0 = 10 | 4 | 2(10)/4 = 5 |

The store can prepare five assortments from two boxes each of these screws.

1.71 The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the mass of the product formed. Starting amounts are in kilograms (1 kg = 103 g), so it will be convenient to work with 103 mol amounts. The balanced equation is given in the problem.

Begin by calculating the initial amounts:





Next, using the balanced chemical equation, construct an amounts table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Reaction** | **N2 +** | **3 H 2 →** | **2 NH 3** |
| Initial amount (103 mol) | 2.68 | 37.1 | 0 |
| 103 mol/coeff | 2.68 (LR) | 12.4 |  |
| Change in amount (103 mol) | −2.68 | −8.04 | +5.36 |
| Final amount (103 mol) | 0 | 29.06 | 5.36 |

Do a mole–mass conversion to determine the mass of ammonia that could be produced:



1.72 The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the mass of the product formed. Starting amounts are in kilograms, so it will be convenient to work with 103 mol amounts. The balanced equation is given in the problem.

Calculate the initial amounts:



Next, construct an amounts table:

|  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Reaction | 2 C3H6 | + | 2 NH3 | + | 3 O2 | → | | 2 C3H3N | + 6 H2O |
| Initial amount (103 mol) | 35.6 |  | 39.9 |  | 60.0 |  | | 0 | 0 |
| kmol/coeff | 17.8 (LR) | | 19.95 |  | 20.0 |  | |  |  |
| Change (103 mol) | –35.6 | | –35.6 |  | – | | (35.6) | +35.6 | +3(35.6) |
| Final amount (103 mol) | 0.0 | | 4.3 |  | 6.6 | | | 35.6 | 106.8 |

Do a mole–mass conversion to determine the mass of C3H3N that could be produced:

35.6 × 103 mol 

1.73 The problem gives information about the amounts of starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the final product masses. Starting amounts are in metric tons (1 metric ton = 106 g), so it will be convenient to work with 106 mol amounts. See the answers to Problem 1.49 for the balanced equations:

(a) Calculate the initial amounts:





Now set up an amounts table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Reaction** | **2 H2 +** | **CO →** | **CH3OH** |
| Initial amount (106 mol) | 0.495 | 0.0357 | 0 |
| 106 mol/coeff | 0.248 | 0.0357 (LR) |  |
| Change (106 mol) | −2(0.0357) | −0.0357 | +0.0357 |
| Final amount (106 mol) | 0.424 | 0.00 | 0.0357 |

The mass that could be produced is



(b) Calculate the initial amounts:





Now set up an amounts table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Reaction** | **CaO +** | **3 C →** | **CO +** | **CaC2** |
| Initial amount (106 mol) | 0.0178 | 0.0833 | 0 | 0 |
| 106 mol/coeff | 0.0178 (LR) | 0.0278 |  |  |
| Change (106 mol) | −0.0178 | −3(0.0178) | +0.0178 | +0.0178 |
| Final amount (106 mol) | 0 | 0.0299 | +0.0178 | +0.0178 |

The masses that could be produced are





(c) Calculate the initial amounts:







Now set up an amounts table:

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Reaction** | **2 C2H4 +** | **O2 +** | **4 HCl →** | **2 C2H4Cl2 +** | **2 H2O** |
| Initial amount (106 mol) | 0.0357 | 0.0313 | 0.0274 | 0 | 0 |
| 106 mol/coeff | 0.0179 | 0.0313 | 0.00685 (LR) |  |  |
| Change (106 mol) |  |  | −0.0274 |  |  |
| Final amount (106 mol) | 0.0220 | 0.0245 | 0 | 0.0137 | 0.0137 |

The masses that could be produced are





1.74 The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the final masses of the products. Starting amounts are in 103 kilograms (103 kg = 106 g), so it will be convenient to work with 106 mol amounts. See the answers to Problem 1.50 for the balanced equations:

(a) Begin by calculating the initial amounts:



Next, construct a table of amounts using the balanced equation from the solution to Problem 1.50:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Reaction | 4 NH3 + | 5 O2 → | 4 NO + | 6 H2O |
| Initial amount (106mol) | 0.440 | 0.234 | 0 |  |
| 106 mol/coeff | 0.110 | 0.0468 (LR) |  |  |
| Change (106 mol) | – (0.234) | –0.234 | +  (0.234) | + (0.234) |
| Final amount (106 mol) | 0.253 | 0 | 0.187 | 0.281 |

Now use the results from the amounts table to determine the mass of products produced:



(b) Calculate the initial amounts:



Next, construct an amounts table:

|  |  |  |  |
| --- | --- | --- | --- |
| Reaction | 2 NO + | O2 → | 2 NO2 |
| Initial amount (106 mol) | 0.250 | 0.234 | 0 |
| 106 mol/coeff | 0.125 (LR) | 0.234 |  |
| Change (106 mol) | –0.250 | –(0.250) | +0.250 |
| Final amount (106 mol) | 0 | 0.109 | 0.250 |

Now use the results from the amounts table to determine the mass of products produced:



(c) Calculate the initial amounts:



Next, construct an amounts table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Reaction | 3 NO2 + | H2O → | 2 HNO3 | NO |
| Initial amount (106 mol) | 0.163 | 0.416 | 0 | 0 |
| 106 mol/coeff | 0.0543 (LR) | 0.416 |  |  |
| Change (106 mol) | –0.163 | –(0.163) | +(0.163) | +(0.163) |
| Final amount (106 mol) | 0 | 0.362 | + 0.1086 | +0.0543 |

Now use the results from the amounts table to determine the mass of products produced:



(d) Calculate the initial amounts:



Next, construct an amounts table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Reaction | 4 NH3 + | 3 O2 → | 2 N2 | 6 H2O |
| Initial amount (106 mol) | 0.440 | 0.234 | 0 | 0 |
| 106 mol/coeff | 0.110 | 0.0781(LR) |  |  |
| Change (106 mol) | –(0.234) | –0.234 | +(0.234) | +(0.234) |
| Final amount (106 mol) | 0.128 | 0 | 0.156 | 0.468 |

Now use the results from the amounts table to determine the mass of products produced:



(e) Calculate the initial amounts:



Next, construct an amounts table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Reaction | 4 NH3 + | 6 NO → | 5 N2 | 6 H2O |
| Initial amount (106 mol) | 0.440 | 0.250 | 0 | 0 |
| 106 mol/coeff | 0.110 | 0.0417(LR) |  |  |
| Change (106 mol) | –(0.250) | –0.250 | +(0.250) | +0.250 |
| Final amount (106 mol) | 0.273 | 0 | 0.208 | 0.250 |

Now use the results from the amounts table to determine the mass of products produced:



1.75 The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the masses of the product formed and the remaining reactant.

Begin by determining the balanced chemical equation:

P4 + O2 → P4O10

P is already balanced. To balance O, give O2 a coefficient of 5:

P4 + **5** O2 → P4O10

Next, calculate the initial amounts:

Construct an amounts table:

|  |  |  |  |
| --- | --- | --- | --- |
| **Reaction** | **P4 +** | **5 O2 →** | **P4O10** |
| Initial amount (mol) | 0.0303 | 0.205 | 0 |
| mol/coeff | 0.0303 (LR) | 0.0410 |  |
| Change (mol) | −0.0303 | −5(0.0303) | +0.0303 |
| Final amount (mol) | 0 | 0.0535 | 0.0303 |

Now obtain the mass of P4O10 produced, using the information from the amounts table and the molar mass: *M* = 4(30.97 g/mol) + 10(16.00 g/mol) = 283.9 g/mol

The mass that could be produced is 

Finally, determine the mass of O2 left over:

There would be of O2 left unreacted.

1.76 The problem gives information about the amounts of both starting materials, so this is a limiting reactant situation. We must calculate the number of moles of each species, construct a table of amounts, and use the results to determine the final masses of all species.

Begin by balancing the chemical equation:

Al + Fe3O4 → Fe + Al2O3

1 Al + 3 Fe + 4 O →2 Al + 1 Fe + 3 O

Balance oxygen by giving Fe3O4 a coefficient of 3 and Al2O3 a coefficient of 4:

Al + **3** Fe3O4 →Fe + **4** Al2O3

1 Al + 9 Fe + 12 O →8 Al + 1 Fe + 12 O

Now balance the remaining elements: give Al a coefficient of 8 and Fe a coefficient of 9:

8 Al + 3 Fe3O4 →9 Fe + 4 Al2O3

Calculate the initial amounts:







Now construct an amounts table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Reaction | 8 A1+ | 3 FeO4 **→** | 9 Fe + | 4 A12O3 |
| Initial amount (mol) | 7.41 | 3.02 | 0 | 0 |
| mol/coeff | 0.926 (LR) | 1.01 |  |  |
| Change (mol) | –7.41 | – (7.41) | + (7.41) | + (7.41) |
| Final amount (mol) | 0 | 0.24 | 8.34 | 3.71 |

Determine the masses of species remaining:

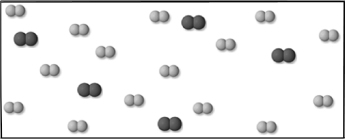




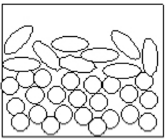




1.77 Both molecular fluorine and molecular chlorine are diatomic (two atoms per molecule) species. To have a total of 20 molecules in a 3:1 ratio there should be 5 molecules of chlorine (dark) and 15 molecules of fluorine (light).



1.78 Both gasoline and water are liquids, and therefore they should be drawn in a tight but not ordered manner. Because gasoline does not dissolve in water and is less dense than water, the gasoline will be on top of the water:



1.79 Kelvin–Celsius conversions involve addition or subtraction rather than multiplication or division, because the two scales have the same unit size but different zero points:

*T*(K) = *T*(ºC) + 273.15 *T* = –11.5 ºC + 273.15 = 261.7 K

1.80 Kelvin–Celsius conversions involve addition or subtraction rather than multiplication or division, because the two scales have the same unit size but different zero points:

|  |  |
| --- | --- |
| *T*(K) = *T*(ºC) + 273.15 | *T*(ºC) = 1235 K – 273.15 = 962 ºC |

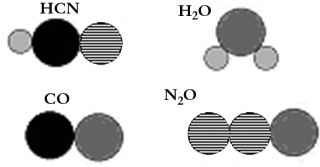
1.81 The number of significant figures in a result of multiplication/division is determined by the number that contains the least number of significant figures:



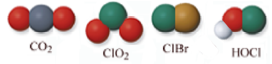
1.82 When adding and subtracting powers of ten, be sure to include the power of ten with each value:



1.83 To draw molecular models, make use of the colour-coded, scaled atoms shown in Figure 1-4 in your textbook. Here, we use shading to indicate different elements:



1.84 To draw molecular models, make use of the colour-coded, scaled atoms shown in Figure 1-4 in your textbook:



1.85 “Same number of atoms” also means “same number of moles,” so work with moles:





1.86 “Same number of atoms” also means “same number of moles,” so work with moles:



1.87 The chemical formula of a substance provides all the information needed to compute its molar characteristics:

(a) *M* = 2(26.98 g/mol) + 3[32.07 g/mol + 4(16.00 g/mol)] = 342.17 g/mol

(b) 

(c) To determine the percent composition, work with 1 mole of substance. Take the ratio of the mass of each element that 1 mole contains to the mass of 1 mole (molar mass):



(d) 

1.88 The chemical formula of a substance provides all the information needed to compute its molar characteristics:



(c) To determine the percent composition, work with 1 mole of substance. Take the ratio of the mass of each element that 1 mole contains to the mass of 1 mole (molar mass):



1.89 Convert from mass to moles to determine molarity:





1.90 Convert from mass to moles to determine molarity:



1.91 The question asks about mole–mass–number quantities. Although there is much interesting biomedical information provided, the only relevant data are the chemical formula of verapamil, C27H38O4N2, and the amount of verapamil per tablet, 120.0 mg:

(a) *M* = 27(12.01 g/mol) + 38(1.008 g/mol) + 4(16.00 g/mol) + 2(14.01 g/mol) = 454.59 g/mol

(b) 

(c) 

1.92 The question asks about mole–mass–number quantities. Although there is much interesting biomedical information provided, the only relevant data are the chemical formula of penicillin, C16H18O4N2S, and the mass of penicillin:



1.93 This problem describes a redox reaction. We are asked to determine the percent yield for the reaction. The starting materials are Xe and F2, and the product is XeF4:

Xe + 2 F2 → XeF4

The problem gives information about the amount of one of the starting materials (the other is in excess), so this is a simple mass–mole–mass problem. We must calculate the theoretical yield, determine the percent yield, and use the actual yield to determine the final amount of Xe left over.

Calculations of theoretical yield:

 = 131.3 g/mol + 4(19.0 g/mol) = 207.3 g/mol

The theoretical yield is determined from the starting amount of Xe:





Now determine the percent yields using the actual yield and the theoretical yield:



The mass of Xe left unreacted is the difference between the starting mass and the amount converted into XeF4:



1.94 This problem describes an acid–base reaction. We are asked to determine the percent yield for the reaction. Begin by analyzing the chemistry. The starting materials are CaF2, a weak base, and H2SO4, a strong acid. In addition to water, the major species are CaF2, H3O+, and HSO4–. The presence of hydronium ions and CaF2 together as major species will result in an acid–base reaction.

The balanced reaction is

CaF2 (*s*) + 2 H3O+ (*aq*) →2 HF (*aq*) + Ca2+ (*aq*) + 2 H2O (*l*)

The problem gives information about the amount of one of the starting materials (the other is in excess), so this is a simple mass–mole–mass problem. We must calculate the theoretical yield, determine the percent yield, and use the actual yield to determine the final amount of CaF2 left over.

Calculate the theoretical yield from the starting amount of CaF2:



(Carry an extra decimal to avoid rounding errors.)

The actual yield of HF is 155 kg 

Now determine the percent yields, using the actual yield and the theoretical yield:

The percent yield is (100%)

The mass of CaF2 left unreacted is the difference between the starting mass and the amount converted into HF:



1.95 Examine the molecular picture to see what chemical reaction occurs. The reactants are atomic X and atomic Y and the product of the reaction is YX2 (there are atoms of X left over). Thus, the reaction is Y + 2 X → YX2.

1.96 Examine the molecular picture to see what chemical reaction occurs. The reactants are atomic U and diatomic V2 and the product of the reaction is UV (there is a molecule of V2 left over). Thus, the reaction is: 2 U + V2 →2 UV

1.97 This problem describes a combustion reaction. Begin by analyzing the chemistry. The problem asks about the amount of a product that forms from a given amount of reactant. First balance the chemical equation. The reaction is

C2H5OH + O2 → CO2 + H2O

2 C + 6 H + 3 O → 1 C + 2 H + 3 O

Give CO2 a coefficient of 2 and H2O a coefficient of 3 to balance C and H:

C2H5OH + O2 → **2** CO2 + **3** H2O

2 C + 6 H + 3 O → 2 C + 6 H + 7 O

There are 7 O on the product side, so give O2 a coefficient of 3 to balance O:

C2H5OH + 3 O2 → 2 CO2 + 3 H2O

Use the balanced equation to do the appropriate mass–mole–number calculations:

(a) 

(b) 

(c) 

1.98 The problem asks about amount of a product that forms from a given amount of reactant.

First balance the chemical equation. The reaction is

C8H18 + O2 → CO2 + H2O

8 C + 18 H + 2 O → 1 C + 2 H + 3 O

Give CO2 a coefficient of 8, and H2O a coefficient of 9 to balance C and H:

C8H18 + O2 → **8** CO2 + **9** H2O

8 C + 18 H + 2 O → 8 C + 18 H + 25 O

There are 25 O on the product side, so give O2 a coefficient of 25/2 to balance O and multiply through by 2 to eliminate the fraction:

2 C8H18 + 25 O2 → 16 CO2 + 18 H2O

Use the balanced equation to do the appropriate mass–mole–number calculations:



1.99 There is much information about propylene oxide provided in this problem, but all that is needed for the calculations are the balanced chemical equation, the amount of starting material, and molar masses. The stoichiometry of this reaction is 1:1 in all reagents:

*M*propylene oxide = 3(12.01 g/mol) + 6(1.01 g/mol) + 1(16.00 g/mol) = 58.1 g/mol

*M*propene = 3(12.01 g/mol) + 6(1.01 g/mol) = 42.1 g/mol

*M*hydroperoxide = 4(12.01 g/mol) + 10(1.01 g/mol) + 2(16.00 g/mol) = 90.1 g/mol

(a) 

(b) 

1.100 There is much information about vinyl chloride provided in this problem, but all that is needed for the calculations are the balanced chemical equation, the amount of starting material, and molar masses. The stoichiometry of this reaction is 1:1 with respect to ethylene, HCl, and vinyl chloride.



1.101 Each atom contributes a distance equal to its diameter, so the number of atoms is the total distance divided by the diameter of an atom. A unit conversion is required between centimetres and picometres:



1.102 This is a unit-conversion problem. Multiply by dimensional ratios to convert units:



1.103 To determine how many atoms of an isotope are in a sample of an element, we must first determine how many atoms of that element are in the sample. Then, we can use the percentage composition to calculate how many atoms of one particular isotope there are. Use the molar mass of chromium to determine the number of moles, and then multiply by Avogadro’s number to convert to number of atoms:



If 9.5% of the Cr is chromium-53:

# = (1.425 × 1021 atoms)(0.095) = 1.35 × 1020 atoms of 53Cr

1.104 This problem involves a sequence of mole–mass–number–volume conversions:

(a) = 3(22.99 g/mol Na) + 1(30.97 g/mol P) + 4(16.00 g/mol O) = 163.94 g/mol



1.105 Moles and mass in grams are related through the equation, . When masses are not given in grams, unit conversions must be made (1 lb = 453.6 g):

= 3(1.008 g/mol) + 1(30.97 g/mol) + 4(16.00g/mol) = 97.99 g/mol



The chemical formula shows that there is 1 mole of P in every mole of H3PO4. The problem states that 15% of the annual production of H3PO4 comes from elemental P:





1.106 This is an empirical formula problem involving combustion analysis. Do a mass analysis of the combustion data to obtain the mass percent composition of the compound. Determine the mass percent oxygen by difference. Once the mass percentages are known, use elemental molar masses to determine relative amounts of the elements.

% N = 3.8% (stated in the problem)



Assume that the remainder of the contents of the compound is oxygen:

% O = 100% – (3.8% + 68.73% + 6.016%) = 21.5%

Now calculate the number of moles of each element in 100. g of the compound:



Divide each by the smallest among them, 0.27 mol N, to determine the relative amounts of each element in the compound:



These values match those for heroin, C21H22O5N.

1.107 The problem asks about yields and provides information about the amounts of both starting materials, so it is both a limiting reagent and a yield problem. Use a table of amounts to determine the theoretical yield.

Calculations of initial amounts:



|  |  |  |  |
| --- | --- | --- | --- |
| **Reaction** | **N2 +** | **3 H2 →** | **2 NH 3** |
| Initial amount (103 mol) | 3.00 | 11.9 | 0 |
| mol/coeff | 3.00 (LR) | 3.97 |  |
| Change (103 mol) | −3.00 | −9.00 | +6.00 |
| Final amount (103 mol) | 0 | 2.9 | 6.00 |

Use the table to obtain the theoretical yield of NH3:



The theoretical yield is 102 kg, and the actual yield is 68 kg:



Multiply the theoretical change by the percent yield to get the actual change:

Actual change = (67%)(6.00 × 103 mol) = 4.00 × 103 mol

Repeat the table of amounts using the actual change rather than the theoretical change:

|  |  |  |  |
| --- | --- | --- | --- |
| **Reaction** | **N2 +** | **3 H2 →** | **2 NH3** |
| Amount (103 mol) | 3.00 | 11.9 | 0 |
| Change (103 mol) | −2.00 | −6.00 | +4.00 |
| Final (103 mol) | 1.00 | 5.9 | 4.00 |

Do mole–mass conversions to determine the masses of leftover reactants:



1.108 The quantity asked for is a volume ratio, which can be related to a mass ratio using densities and to a mole ratio using molar masses. First balance the chemical reaction:

N2H4 (*l*) + H2O2 (*l*) → N2 (*g*) + H2O (*g*)

2 N + 6 H + 2 O → 2 N + 2 H + O

N is already balanced; give H2O a coefficient of 2 to balance O:

N2H4 (*l*) + H2O2 (*l*) → N2 (*g*) + 2 H2O (*g*)

2 N + 6 H + 2 O → 2 N + 4 H + 2 O

Now there are 6 H on the left and 4 on the right; multiply the coefficients of both H2O2 and H2O by 2 to give 8 on each side:

N2H4 (*l*) + 2 H2O2 (*l*) → N2 (*g*) + 4 H2O (*g*)

The two reactants should be present in a mole ratio of 1 N2H4 : 2 H2O2.

Use *m* = *nM* and , giving :



1.109 This problem asks about a mixture of two salts, CaSO4 and Ca(H2PO4)2:

(a) The 2:1 mole ratio indicates the coefficients for the balanced equation. There is no water present during this reaction, so we need not ask about species present:

2 H2SO4 + Ca3(PO4)2 → 2 CaSO4 + Ca(H2PO4)2

(b) Because the mixture has fixed stoichiometric proportions of 2:1, it has a combined molar mass of :

*M*combination = 2(136.1 g/mol) + 1(234.1 g/mol) = 506.3 g/mol









(c) Each mole of Ca3(PO4)2 yields 2 moles of phosphate ion, so the amount of phosphate ion available is = 2(98.76 mol) = 198 mol

1.110 The problem asks about the amount of reactant needed to react completely with a given amount of another reactant. Use mole amounts and stoichiometric ratios:



1.111 This is a density problem. The data provided are sufficient to determine the number density of red blood cells. The number of red blood cells (#rbc) in an adult human is then the volume of blood multiplied by the number density:



1.112 The key to solving this problem is to recognize that the average density of the sphere plus its air contents must be less than the density of water, 1.00 g/cm3:



The mass of each component can be related to its density, the total volume of the sphere, and the fraction of the total volume occupied by that component:

*m*component=component *f*component *V*total

There are only two components, air and the metal, so

*f*metal = 1 – *f*air from which *m*total=  *f*air *V*total + metal(1 – *f*air) *V*total

Substitute this expression for the total mass into the density equation:



Cancel the volume from this equality, substitute the appropriate densities (all in g/cm3), and solve for the fraction of the total volume that must be air.

1.00 g/cm3 = (0.001189 g/cm3) *f*air + metal −metal *f*air

metal− 1.00 g/cm3 = metal *f*air  − (0.001189 g/cm3) *f*air



The density of air is much smaller than the density of any metal, so the second term in the denominator can be neglected, giving the final result:



For A1, = 2.70 g/cm3 and *f*air = 0.63; for Cu,  = 8.96 g/cm3 and *f*air = 0.89

For any volume of air, *V*air, the total volume of the sphere must be 

For an air volume of 1.50 L the total sphere volumes are as follows:

For A1, *V*total = 1.50 L/0.63 = 2.38 L; for Cu, *V*total = 1.50 L/0.89 = 1.685 L

We can calculate the radius of the air sphere, whose volume is 1500 cm3, using



We can calculate the radius of each total sphere in the same fashion:



Thus, the A1 shell must have a thickness of (8.28 − 7.10) = 1.18 cm, and the Cu shell must have a radius of (7.38 – 7.10) = 0.28 cm.

1.113 Make use of tabulated data in the *CRC Handbook of Chemistry and Physics* to identify the *Z­-*values of elements falling in each category, and use the periodic table to determine the symbol for each:

Elements with only one stable isotope, with their *Z-*values, are the following: 4, Be; 9, F; 11, Na; 13, Al; 15, P; 21, Sc; 23, V; 25, Mn; 27, Co; 33, As; 39, Y; 41, Nb; 45, Rh; 53, I; 55, Cs; 59, Pr; 65, Tb; 67, Ho; 69, Tm; 79, Au; and 83, Bi. Elements with four or more stable isotopes, with their *Z-*values, are the following: 16, S; 20, Ca; 22, Ti; 24, Cr; 26, Fe; 28, Ni; 30, Zn; 32, Ge; 34, Se; 36, Kr; 38, Sr; 40, Zr; 42, Mo; 44, Ru; 46, Pd; 48, Cd; 50, Sn; 52, Te; 54, Xe; 56, Ba; 58, Ce; 60, Nd; 62, Sm; 64, Gd; 66, Dy; 68, Er; 70, Yb; 72, Hf; 74, W; 76, Os; 78, Pt; and 82, Pb. Apart from Be (*Z* = 4), all the elements with just one stable isotope have an odd number of protons (odd *Z-*value). In sharp contrast, all the elements having four or more stable isotopes have an even number of protons (even *Z-*value). In addition, five light elements (*Z* < 16) have just one stable isotope, whereas none of the light elements has four or more.

1.114 This problem contains a series of mole–mass–number problems.

*M*aluminum sulphate = 2(26.98 g/mol) + 3(32.07 g/mol) + 30(16.00 g/mol) + 36(1.01 g/mol) = 666.53 g/mol



1.115 (a) Molar mass of HFC-134a (CF3CH2F)

= 2(12.011 g/mol) + 4(18.998 g/mol) + 2(1.008 g/mol)

= 102.03 g/mol

Molar mass of trichloroethylene (CCl2CHCl)

= 2(12.011 g/mol) + 3(35.453 g/mol) + 1.008 g/mol

= 131.389 g/mol

Moles of HFC-134a produced = (65 × 106 kg)(1000 g/kg)/(102.03 g/mol)

= 6.37 × 108 moles

The reaction yield is 47%; therefore, to produce this amount of HFC-134a, you would require enough raw materials to make 

*n*trichloroethylene = *n*HFC-134a



(b) If the process were 53% efficient (rather than 47%), then the 178 million kg (1.36 × 109 moles) of trichloroethylene would produce:



(c) Excess production = (74 – 65) million kg = 9 million kg



1.116First we need a balanced reaction to work with. Because there are four F atoms on the right, we need four HF molecules on the left:

UO2 + 4 HF → UF4 + H2O

Now the H and O atoms can be balanced by increasing the coefficient of H2O to 2:

UO2 + 4 HF → UF4 + 2 H2O

25.0 tonnes of UF4 per day is converted into moles:



From the balanced reaction, we see that the manufacture of 1 mol UF4 requires 4 mol HF. Accounting for the yield, the total HF required is therefore



Now we can calculate the required volume of HF solution:



1.117To solve this problem requires molecular weights of the starting material and of the product, but not of the two intermediate molecules. The molecular formula of isobutylbenzene is C10H14, which has a molecular weight of 134.2 g/mol, and that of ibuprofen is C13H18O2, which has a molecular weight of 206.3 g/mol.

If the yields of all three reactions were 100%, the theoretical mass of the product would be limited by the mass of the reactant:



But the actual overall yield will be a function of the product of the yields of the three reactions, namely 91% x 90% x 85% = 69.6%, so the actual mass of ibuprofen produced will be 