**Answer Key to Stoichiometry Problem Set**

Most stoichiometry problems follow the strategy:

Quantity A → mol A → mol B→ Quantity B

You will be using portions of this strategy or the entire strategy to complete this problem set.

1. Sum the atomic weights (AW) of each of the elements in the compound according to the number of each element present. The molar mass is defined as the number of grams of substance per mole of substance. It has the same mass in g/mol as one formula unit in amu.

 a. AlCl3

 Molar mass = AW Al + 3(AW Cl) = 26.98 g/mol + 3(35.45 g/mol) = 133.33 g/mol

 b. Na2SO4

 Molar mass = 2(AW Na) + AW S + 4(AW O) = 2(22.99 g/mol) + 32.07 g/mol +

 4(16.00 g/mol) = 142.05 g/mol

 c. Cr2(CO3)3

 Molar Mass = 2(AW Cr) + 2(AW C) + 6(AW O) = 2(52.00 g/mol) + 3(12.01 g/mol) +

 9(16.00 g/mol) = 284.03 g/mol

 d. C7H5N3O6

 Molar Mass = 7(AW C) + 5(AW H) + 3(AW N) + 6(AW O) = 7(12.01 g/mol) +

 5(1.01 g/mol) + 3(14.01 g/mol) + 6(16.00 g/mol) = 227.15 g/mol

 e. CuSO4∙5H2O (a hydrate)

 Molar Mass = AW Cu + AW S + 9(AW O) + 10(AW H) = 63.54 g/mol + 32.07 g/mol +

 10(1.01 g/mol) + 9(16.00 g/mol) = 249.71 g/mol

2. a. Use the molar mass of sodium sulfate as a “conversion factor” to calculate the moles of sodium sulfate. In terms of the stoichiometry strategy, the molar mass allows you to convert a Quantity A to mol A or mol B to Quantity B. There are 142.05 grams of Na2SO4 in 1 mole of Na2SO4. Let the units decide how to set up the problem.
$$115 g Na\_{2}SO\_{4} x\frac{1 mol Na\_{2}SO\_{4}}{142.05 g Na\_{2}SO\_{4}} = 0.810 Na\_{2}SO\_{4}$$

 b. Here we are taking Quantity A to mol A to mol B. The subscript of 2 in the formula for Na2SO4 indicates there are 2 moles of sodium ions per mole of sodium sulfate. Let the units help you set up the problem:

$$115 g Na\_{2}SO\_{4} x\frac{1 mol Na\_{2}SO\_{4}}{142.05 g Na\_{2}SO\_{4}} x \frac{2 mol Na^{+}}{1 mol Na\_{2}SO\_{4}} = 1.62 mol Na^{+}$$

 c. Again, here we are taking Quantity A to mol A to mol B. In mole of sodium sulfate there is one mole of sulfate ion. The units help you set up the problem in a similar fashion as in problem 2a:

$$115 g Na\_{2}SO\_{4} x\frac{1 mol Na\_{2}SO\_{4}}{142.05 g Na\_{2}SO\_{4}} x \frac{1 mol SO\_{4}^{2-}}{1 mol Na\_{2}SO\_{4}} = 0.810 mol SO\_{4}^{2-}$$

 d. In one mole of sodium sulfate there is 1 mole of sulfur as indicated by the subscript of one in the formula. Once again, units help set up the problem for converting Quantity A to mol A to mol B:

$$115 g Na\_{2}SO\_{4} x\frac{1 mol Na\_{2}SO\_{4}}{142.05 g Na\_{2}SO\_{4}} x \frac{1 mol S}{1 mol Na\_{2}SO\_{4}} = 0.810 mol S$$

 e. In mole of sodium sulfate there are four moles of oxygen as indicated by the subscript of four in the formula. Once again, units help set up the problem for converting Quantity A to mol A to mol B:

$$115 g Na\_{2}SO\_{4} x\frac{1 mol Na\_{2}SO\_{4}}{142.05 g Na\_{2}SO\_{4}} x \frac{4 mol O}{1 mol Na\_{2}SO\_{4}} = 3.24 mol O$$

3 a. Here we already know mol A and will convert to mol B to Quantity B of the strategy. We will need the molar mass of potassium carbonate to do the mol B to Quantity B conversion:

 Molar mass = 2(AW K) + AW C + 3(AW O) = 2(39.10 g/mol) + 12.01 g/mol +

 3(16.00 g/mol) = 138.21 g/mol

Let the units help you set up the problem. One mole of potassium carbonate has a mass of 138.21 grams:

$$3.5 mol K\_{2}CO\_{3} x \frac{138.21 g K\_{2}CO\_{3}}{1 mol K\_{2}CO\_{3}} = 4.8 x 10^{2} g K\_{2}CO\_{3}$$

 b. Convert mol A to mol B. In one mole of potassium carbonate there are two potassium ions:

$$3.5 mol K\_{2}CO\_{3} x \frac{2 mol K^{+}}{1 mol K\_{2}CO\_{3}} = 7.0 mol K^{+}$$

 c. Once again, convert mol A to mol B. In one mole of potassium carbonate there is one mole of carbonate:

$$3.5 mol K\_{2}CO\_{3} x \frac{1 mol CO\_{3}^{2-}}{1 mol K\_{2}CO\_{3}} = 3.5 mol CO\_{3}^{2-}$$

 d. Here, the number of particles of potassium ions is required. This tips you off that Avogadro’s number is needed somewhere in the problem. Avogadro’s number is the number of particles in one mole of substance. There will be 6.02 x 1023 potassium ions in one mole of potassium ions. Convert mol A to mol B to Quantity B. In this case, mol potassium carbonate to mol potassium ions to number of potassium ions:

$$3.5 mol K\_{2}CO\_{3} x \frac{2 mol K^{+}}{1 mol K\_{2}CO\_{3}} x \frac{6.02 x 10^{23} K^{+} ions}{1 mol K^{+}} = 4.2 x 10^{24} K^{+} ions^{}$$

 e. This problem is similar to 3d. Convert mol A to mol B to Quantity B. In this case, mol potassium carbonate to mol carbon to number of carbon atoms:

$$3.5 mol K\_{2}CO\_{3} x \frac{1 mol C}{1 mol K\_{2}CO\_{3}} x \frac{6.02 x 10^{23} C atoms}{1 mol C} = 2.1 x 10^{24} C atoms$$

4. When balancing chemical reactions there are a few things to keep in mind. Never change the subscripts in the formulas found in the reaction. Place coefficients in front of the substance, never in the middle of the substance. Polyatomic ions may be balanced as a unit, if they are unchanged on both sides of the reaction. There are a few tricks one can use to quickly balance chemical reactions. If an element is off by itself, balance that element last. You can always put any number needed in the end to balance that element. If there is an odd number of an element on one side the reaction, and an even number on the other, make the odd side even by multiplying by two.

 a. Balance the sulfur first, as there is an odd number on the right hand side and an even number on the left hand side. In this case, don’t multiply by 2, rather balance it with an 8:

S8 + F2 → 8 SF4

Now the sulfur is balance, but there are 2 F on the left hand side and 32 F on the right hand side. Balance the F with a coefficient of 16 on the left hand side:

S8 + 16 F2 → 8 SF4

 b. Balance the Cl last, as it found as an element on the left hand side. The carbon is already balanced, so balance the H. There are 4 H on the left hand side, so place a coefficient of 4 with the HCl to balance the H:

CH4 + Cl2 → CCl4 + 4 HCl

The carbon is still balanced along with the H. There are 2 Cl on the left hand side and 8 Cl on the right hand side Don’t forget to look at all the products – 4 Cl are found in CCl4 and 4 Cl are with the 4 HCl. Balance the Cl by placing a coefficient of 4 with the Cl on the left hand side:

CH4 + 4 Cl2 → CCl4 + 4 HCl

 c. Oxygen is off by itself, so balance it last. There are 5 C on the left hand side and 1 C on the right hand side. Balance the carbon by placing a coefficient of 5 in front of CO2:

C5H12 + O2 → 5 CO2 + H2O

Next balance the H. There are 12 H on the left hand side and 2 H on the right hand side. Place a coefficient of 6 in front of H2O to balance the H:

C5H12 + O2 → 5 CO2 + 6 H2O

Lastly, balance the O. There are 2 O on the left hand side and 16 O on the right hand side. Place a coefficient of 8 in front of O2 to balance the O:

C5H12 + 8 O2 → 5 CO2 + 6 H2O

 d. In this problem, there isn’t any easy place to start. To keep things simpler, notice that the polyatomic ions can be balanced as the ion, rather than as separate elements. Barium is balanced, but there are 2 NO3─ ions on the left hand side and 1 NO3─ ion on the right hand side. Place a coefficient of 2 in front of NaNO3 to balance nitrate:

Ba(NO3)2 + Na2CO3 → BaCO3 + 2 NaNO3

This also balances the Na. There are now 2 Na on both sides of the reaction. Also, CO32- is balanced as well. There is one carbonate in both sides of the reaction.

 e. Once again, there isn’t an easy place to start, but you should recognize that phosphate ions, PO43- can be balanced as the polyatomic ion. Start with balancing the calcium (arbitrary). There is 1 Ca on the left hand side and 3 Ca on the right hand side. Place a coefficient of 3 in front of CaCl2 to balance the calcium:

3 CaCl2 + K3PO4 → Ca3(PO4)2 + KCl

From here, continue on sequentially. There are now 6 Cl on the left hand side and 1 Cl on the right hand side. Balance the Cl by placing a coefficient of 6 in front of KCl to balance the Cl:

3 CaCl2 + K3PO4 → Ca3(PO4)2 + 6 KCl

Now there are 6 K on the right hand side and 3 K on the left hand side. Balance the K by placing a coefficient of 2 in front of K3PO4:

3CaCl2 + 2 K3PO4 → Ca3(PO4)2 + 6 KCl

This also balances the PO43-. There are now 2 phosphate ions on the left hand side and 2 phosphate ions on the right hand side.

5. Balance the reaction first: C3H8 + 5 O2 → 3 CO2 + 4 H2O. We will be back to using the stoichiometry strategy, Quantity A → mol A → mol B → Quantity B.

 a. The molar mass of propane, C3H8, will be needed:

 Molar mass = 3(AW C) + 8(AW H) = 3(12.01 g/mol) + 8(1.01 g/mol) = 44.11 g/mol

Use the molar mass to convert grams of propane to mol of propane and the balance reaction to convert mol of propane to mol of oxygen gas needed. From the strategy, we are using Quantity A → mol A → mol B:
$$12.5 g C\_{3}H\_{8} x \frac{1 mol C\_{3}H\_{8} }{44.11 g C\_{3}H\_{8}} x \frac{5 mol O\_{2}}{1 mol C\_{3}H\_{8}} = 1.42 mol O\_{2}$$

 b. From the balanced reaction, 3 moles of carbon dioxide are formed from 1 mole of propane. Once again, we are converting grams of propane to moles of propane using its molar mass and then converting moles of propane to moles of carbon dioxide. From the strategy, we are using Quantity A → mol A → mol B:
$$12.5 g C\_{3}H\_{8} x \frac{1 mol C\_{3}H\_{8} }{44.11 g C\_{3}H\_{8}} x \frac{3 mol CO\_{2}}{1 mol C\_{3}H\_{8}} = 0.850 mol CO\_{2}$$

 c. Here we will use the entire strategy. Convert the grams of propane to moles of propane, then moles of propane to moles of carbon dioxide and finally moles of carbon dioxide to grams of carbon dioxide. We will need to calculate the molar mass of carbon dioxide. All other information has already been previously found in parts a – c of the problem:

 Molar mass of CO2 = AW C + 2(AW O) = 12.01 g/mol + 2(16.00 g/mol) = 44.01 g/mol

$$12.5 g C\_{3}H\_{8} x \frac{1 mol C\_{3}H\_{8} }{44.11 g C\_{3}H\_{8}} x \frac{3 mol CO\_{2}}{1 mol C\_{3}H\_{8}} x \frac{44.01 g CO\_{2}}{1 mol CO\_{2}} = 37.4 g CO\_{2}$$

 d. The mass of carbon dioxide calculated in part c of this problem is often referred to as the theoretical yield. The theoretical yield represents the maximum amount of product that can be produced in an ideal world with ideal chemists who can collect every single particle of product formed. Needless to say in the real world this doesn’t happen. As a result one often can calculate a percent yield. A percent is always a “part/whole” ratio. The actual yield, 35.2 g CO2 represents the “part” and the theoretical yield, 37.4 g CO2, represents the “whole.” To calculate the percent yield:

$$\% yield = \frac{actual yield}{theoretical yield} x 100 = \frac{35.2 g CO\_{2}}{37.4 g CO\_{2}} x 100 = 94.1 \% yield$$

 e. Once again, the entire strategy is used. Since we are asked for the number of molecules (particles) of CO2, Avogadro’s number will be needed. Convert the mass of C3H8 to moles of C3H8 to moles of CO2 to molecules of CO2:

$$12.5 g C\_{3}H\_{8} x \frac{1 mol C\_{3}H\_{8} }{44.11 g C\_{3}H\_{8}} x \frac{3 mol CO\_{2}}{1 mol C\_{3}H\_{8}} x \frac{6.02 x 10^{23}}{1 mol CO\_{2}} = 5.1 x 10^{23} molecules CO\_{2}$$

6. a. Convert the number of copper atoms to moles copper to grams copper:

$$5.39 x 10^{24} Cu atoms x \frac{1 mol Cu}{6.02 x 10^{23} Cu atoms} x \frac{63.54 g Cu}{1 mol Cu} = 569 g Cu$$

 b. Convert the number of copper atoms to moles copper to moles of nitric acid:

$$5.39 x 10^{24} Cu atoms x \frac{1 mol Cu}{6.02 x 10^{23} Cu atoms} x \frac{4 mol HNO\_{3}}{1 mol Cu} = 35.81 mol HNO\_{3}$$

 c. Convert the number of copper atoms to moles copper to moles copper(II) nitrate to grams of copper(II) nitrate using the balanced reaction:

$$5.39 x 10^{24} Cu atoms x \frac{1 mol Cu}{6.02 x 10^{23} Cu atoms} x \frac{1 mol Cu(NO\_{3})\_{2}}{1 mol Cu} x \frac{187.56 g Cu(NO\_{3})\_{2}}{1 mol Cu(NO\_{3})\_{2}}$$

 = 1.68 x 103 g Cu(NO3)2

 d. Since molecules are particles and directly analogous to moles, one can convert molecules copper directly to molecules nitrogen dioxide using the coefficients in the balanced reaction:

$$5.39 x 10^{24} Cu atoms x \frac{2 NO\_{2} molecules }{1 Cu atom} = 1.08 x 10^{25} NO\_{2} molecules$$

 e. Convert the atoms of copper to moles of copper to moles of water using the balanced reaction:

$$5.39 x 10^{24} Cu atoms x \frac{1 mol Cu}{6.02 x 10^{23} Cu atoms} x \frac{2 mol H\_{2}O}{1 mol Cu} = 17.9 mol H\_{2}O$$

7. A limiting reactant problem is recognizable because quantities of both reactants are given in the problem. One of these reactants will be used up before the other is totally consumed.

 a. To find the limiting reactant, first find the moles of each reactant so that they may be compared. Don’t jump to any conclusions about the limiting reactant. The reactant with a lower mass isn’t always the limiting reactant.

$$6.20 g C\_{2}H\_{6} x \frac{1 mol C\_{2}H\_{6}}{30.08 g C\_{2}H\_{6}} = 0.206 mol C\_{2}H\_{6}$$

$$92.0 g Br\_{2} x \frac{1 mol Br\_{2}}{159.82 g Br\_{2}} = 0.576 mol Br\_{2}$$

Once again, don’t jump to any conclusion about the limiting reactant. Even though it appears that there are fewer moles of ethane, a comparison has not been made to establish which reactant will be consumed first. Pick one of the reactants and calculate how many moles of the other reactant is needed. Let’s arbitrarily choose ethane and calculate the moles of bromine needed using the balanced reaction:

$$0.206 mol C\_{2}H\_{6} x \frac{6 mol Br\_{2} needed}{1 mol C\_{2}H\_{6}} = 1.24 mol Br\_{2} needed$$

Although it appears that there is a greater amount of bromine present, bromine turns out to be the limiting reactant: 0.618 moles of bromine are needed to react all 0.206 mol of ethane, and only 0.576 mole of bromine is available.

 b. Use the moles of limiting reactant and the balanced reaction to calculate the moles of HBr formed:
$$0.576 mol Br\_{2} x \frac{6 mol HBr}{6 mol Br\_{2}} = 0.576 mol HBr $$

 c. Using the moles of limiting reactant and the balanced reaction to calculate the theoretical yield in grams of C2Br6:

$$0.576 mol Br\_{2} x \frac{1 mol C\_{2}Br\_{6}}{6 mol Br\_{2}} x \frac{503.74 g C\_{2}Br\_{6}}{1 mol C\_{2}Br\_{6}} = 48.4 g C\_{2}Br\_{6}$$

 d. A percent is always a “part/whole” ratio. Use the actual yield as the “part” and the theoretical yield as the “whole”:
$$\% yield = \frac{actual yield}{theoretical yield} x 100 = \frac{45.9 g actual}{96.7 g theoretical } x 100 = 94.9 \% yield $$

 e. The reactant in excess was ethane. Find the moles of ethane that actually reacted using the balanced reaction:
$$0.576 mol Br\_{2}x \frac{1 mol C\_{2}H\_{6} used }{6 mol Br\_{2}} = 0.0960 mol C\_{2}H\_{6} used$$

Subtract the mole ethane used from the total moles of ethane available:

0.206 mol C2H6 – 0.0960 mol C2H6 = 0.110 mol C2H6 in excess

Convert the moles of ethane in excess to grams:

$$0.110 mol C\_{2}H\_{6} x \frac{30.08 g C\_{2}H\_{6}}{1 mol C\_{2}H\_{6}} = 3.31g C\_{2}H\_{6} in excess$$

8. a. This is a limiting reactant problem because the quantities of both reactants are given. It is impossible to tell which reactant is used up before the other. A comparison of both reactants must be made in moles to determine the limiting reactant. Convert each of the reactant quantities to moles. The atomic weight of copper is used to convert grams copper to moles copper:
$$14.7 g Cu x \frac{1 mol Cu}{63.54 g Cu} = 0.231 mol Cu$$

Think of molarity as a conversion factor that allows you to convert moles of solute to liters of solution or in this case, volume of solution to moles of solute:

$$450.0 mL soln x \frac{1 x 10^{-3} L}{1 mL soln} x \frac{2.50 mol HNO\_{3}}{1 L soln} = 1.12 mol HNO\_{3}$$

Choose one reactant and calculate the moles of the other reactant needed. In this case, mole of copper is arbitrarily chosen and the mole of nitric acid needed is calculated:

$$0.231 mol Cu x \frac{4 mol HNO\_{3} needed }{1 mol Cu} = 0.924 mol HNO\_{3} needed$$

Compare the moles of nitric acid needed to the moles of nitric acid available: 0.924 mol of HNO3 are needed and 1.12 mol of HNO3 is available. It is clear that HNO3 is in excess, thus Cu is the limiting reactant. Use the mole of limiting reactant to calculate the theoretical yield in grams of copper(II) nitrate formed:

$$0.231 mol Cu x \frac{1 mol Cu(NO\_{3})\_{2}}{1 mol Cu} x \frac{187.56 g Cu(NO\_{3})\_{2}}{1 mol Cu(NO\_{3})\_{2}} = 43.3 g Cu(NO\_{3})\_{2}$$

 b. The total volume of the solution is 450.0 mL since we are assuming the volume copper is not significantly adding to the total volume of the solution. First find the moles of Cu(NO3)2 formed and divide the moles of solute by the volume in liters:

$$0.231 mol Cu x \frac{1 mol Cu(NO\_{3})\_{2}}{1 mol Cu} = 0.231 mol Cu(NO\_{3})\_{2}$$

$$450.0 mL soln x \frac{1 x 10^{-3} L}{1 mL soln} = 0.4500 L solution$$

$$\frac{0.231 mol Cu(NO\_{3})\_{2}}{0.4500 L soln} = 0.531 M Cu(NO\_{3})\_{2}$$

 c. Convert the mole of limiting reactant to mole of nitrogen dioxide to molecules of nitrogen dioxoide:
$$0.231 mol Cu x \frac{2 mol NO\_{2}}{1 mol Cu} \frac{6.02 x 1)^{23} molecules NO\_{2}}{1 mol NO\_{2}} = 2.78 x 1O^{23}molecules NO\_{2}$$

 d. The mole of HNO3 needed and the total mole of HNO3 available has already been found in part a of the problem. To find the mole of HNO3 in excess, subtract the mole of HNO3 needed from the total mole of HNO3 available:

1.12 mol HNO3 – 0.924 mol HNO3 = 0.20 mol HNO3 in excess

 e. A percent is always a “part/whole” ratio. Use the actual yield as the “part” and the theoretical yield as the “whole”:
$$\% yield = \frac{actual yield}{theoretical yield} x 100 = \frac{39.5 g actual}{43.3 g theoretical } x 100 = 91.2 \% yield $$