Chemical Conversions and Problems

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Temperature Conversions

To convert between temperature in Kelvins (TK) and degrees Celsius (T°C):

$$TK = T°C + 273.15 \quad \text{or} \quad T°C = TK - 273.15$$

To convert between temperature in degrees Celsius (T°C) and degrees Fahrenheit (T°F):

$$T°C = \frac{T°F - 32}{1.80} \quad \text{or} \quad T°F = 1.80(T°C) + 32$$

Example 1: What is the boiling point of water, 100°C, in degrees Fahrenheit?

T°F = 1.80(T°C) + 32 & T°C = 100
T°F = 1.80(100) + 32
T°F =180 + 32
T°F = 212

Basic Unit Conversions

To do a basic conversion from one unit to another:

1) Start with the original number you are given. In example 2, the original number given is 34 minutes.

2) Multiply/divide that original number by a known relationship between that original unit and the unit you want to end up with. In example 2, the known relationship is 60 seconds equals 1 minute.

How you determine whether to put the 60 s on the top and the 1 min on the bottom, or put the 1 min on the top and the 60 s on the bottom, depends on which unit you are starting with and which unit you need to end up with. You want to be able to cancel out the original unit leaving the ending unit on top; put the unit you began with, min, on bottom and the unit you want to end up with, s, on top. In Example 2, you want to be able to cancel out the min, so you put min on bottom, and you want to end up with s, so put s on top. If you were to do a problem converting from seconds to minutes, however, you would put min on top and s on the bottom as in the Example 3

Example 2: How many seconds are in 34 minutes?

$$34 \ \text{min} \times \frac{60 \ \text{s}}{1 \ \text{min}} = 2040 \ \text{s}$$

Example 3: How many minutes are in 25 seconds?

$$25 \ \text{s} \times \frac{1 \ \text{min}}{60 \ \text{s}} = 0.42 \ \text{min}$$

Often in conversion problems, you will have to use more than one *conversion factor* as in Example 4. A conversion factor is a known relationship. Just take these multi-step conversions one step at a time, canceling out all units except the one you want to end up with.

Example 4: How many centimeters are in 10 miles?

Known relationships: 1 mi = 5280 ft, 1 ft = 12 in, and 1 in = 2.54 cm

$$10 \ \text{mi} \times \frac{5280 \ \text{ft}}{1 \ \text{mi}} \times \frac{12 \ \text{in}}{1 \ \text{ft}} \times \frac{2.54 \ \text{cm}}{1 \ \text{in}} = 1.61 \times 10^6 \ \text{cm}$$

(If you need to review how to put numbers in scientific notation, see our handout *Exponents, Radicals, and Scientific Notation*.)

Another type of common conversion problem deals with conversions between some unit and a prefix of that unit such as a conversion from meters to millimeters. The following table provides a list of some widely used prefixes. For example, 1 gigameter (Gm) = 1,000,000,000 meters (m) or 10^9 m, and 1 microinch (μm) = 0.000001 in or 10^{-6} in.

Prefix		Value	
Name	Symbol	Numerical	Exponential
giga	G	1,000,000,000	10^9
mega	M	1,000,000	10^6
kilo	k	1,000	10^3
hecto	h	100	10^2
deka	da	10	10^1
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}

(If you need help converting between different metric prefixes, see our handout on *Metric Conversion*.)

Example 5: How many picograms are there in 4 hectograms?

Known relationships: 1 pg = 0.000000000001 g, and 1 hg = 100 g

$$4 \, \cancel{hg} \times \frac{100 \, \cancel{g}}{1 \, \cancel{hg}} \times \frac{1 \, pg}{0.000000000001 \, \cancel{g}} = 4 \times 10^{14} \, pg$$

Known relationships: 1 pg = 10^{12} g, and 1 hg = 10^2 g

$$4 \, hg \times \frac{10^2 \, \cancel{g}}{1 \, \cancel{hg}} \times \frac{1 \, pg}{10^{12} \, \cancel{g}} = 4 \times 10^{14} \, pg$$

In the above example, it is possible to do the problem in either of the two ways. Both are correct, but the second way is easier since you don't have to deal with very large and very small numbers. The method you choose to use depends on how comfortable you are with using terms in scientific notation.

Chemical Quantity Conversions

Conversions between different units of chemical quantities are a very important part of chemistry. The conversions you will encounter the most will be those between grams, moles, numbers of molecules, and numbers of atoms. The mole is a central unit of quantity in chemistry, and it represents the number of atoms in 12 grams of the carbon-12 isotope, which is 6.022 x 10^{23} atoms. This number, which equals 1 mole, is called

Avogadro's number. You can have a mole of anything, not just atoms. You can have a mole of molecules or a mole of test tubes or a mole of people. A mole, like a dozen, is a standard amount of any object. A conversion between moles of an object and the number of that object can be set up just like the previous unit conversion examples.

Example 6: How many water molecules are there in 6.0 moles of water?

Known relationship: $1 \text{ mol} = 6.022 \times 10^{23}$ molecules (mlcls)

$$6.0 \text{ mol } H_2O \times \frac{6.022 \times 10^{23} \text{ mlcls}}{1 \text{ mol}} = \textbf{3.6} \times \textbf{10}^{24} \textbf{ mlcls } \textbf{H}_2\textbf{O}$$

Example 7: How many moles of Na are there if there are 5,900,000 atoms of Na?

Known relationship: $1 \text{ mol} = 6.022 \times 10^{23}$ atoms

$$5{,}900{,}000 \text{ atoms Na} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} = \textbf{9.8} \times \textbf{10}^{-18} \textbf{ moles Na}$$

Conversions between moles and grams of a chemical can be done by using molar mass, which is the addition of the atomic masses of the atoms making up that molecule. Atomic mass is the mass in grams of 1 mole, 6.022×10^{23} atoms, of an element. These atomic masses are listed on most periodic tables.

Example 8: What is the molar mass of H_2SO_4?

H: 2(1 g/mol) = 2 g/mol
S: 1(32 g/mol) = 32 g/mol
O: 4(16 g/mol) = 64 g/mol
                      98 g/mol

In Example 8, H_2SO_4 was made up of 4 O atoms, so the atomic mass of O was multiplied by 4. Likewise, the atomic masses of H and S were multiplied by 2 and 1 since there were 2 atoms of H and 1 atom of S. Once you have solved for the molar mass of a molecule, you can use it to convert between moles and grams in a conversion problem set up just like basic unit conversions. This is possible because the molar mass of any compound equals 1 mole of that compound.

Example 9: How many grams are in 50. moles of H_2SO_4?

Known relationship: 1 mol = the molar mass of H_2SO_4 = 98 g

$$50. \text{ mol } H_2SO_4 \times \frac{98 \text{ g}}{1 \text{ mol}} = \textbf{4900 g } \textbf{H}_2\textbf{SO}_4$$

Once you know how to convert between moles and grams and between moles and numbers of atoms or molecules, you can perform conversions between grams and number of atoms or molecules by putting together the two known relationships between moles and grams and between moles and number of atoms or molecules.

Example 10: How many kilograms of Ne gas are there if there are 104 atoms of Ne?

Known relationships: 1 mol = 6.022 x 10^{23} atoms, 1 mol Ne = molar mass Ne = 20g

1 kg = 10^3 g (or 1000 g)

$$104 \text{ atoms Ne} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ atoms}} \times \frac{20 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 3.45 \times 10^{-24} \text{ kg Ne}$$

In some problems, you may have to convert between moles or grams of some molecules and atoms of an element in that molecule. For example, you might be given moles or grams of C_2H_6 and be asked to find the number of H atoms in that sample. In problems like this, you can use the fact that there are 6 atoms of H in 1 mlcl of C_2H_6 as a conversion factor.

Example 11: How many atoms of H are there in 124 g of NH_3?

Known relationships: 1 mol = molar mass of NH_3 = 17 g, 1 mol = 6.022 x 10^{23} mlcls, 1 mlcl NH_3 = 3 atoms H

$$124 \text{ g } NH_3 \times \frac{1 \text{ mol}}{17 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ mlcls}}{1 \text{ mol}} \times \frac{3 \text{ atoms H}}{1 \text{ mlcl } NH_3} = 1.31 \times 10^{25} \text{ atoms H}$$

Density

The density of a substance is the ratio of that substance's mass to its volume. The higher the density of a substance, the less space it takes to fill that space with some amount of the substance. For example, imagine you have 100 pounds of lead and 100 pounds of cotton balls. It would take a space with a much larger volume to accommodate 100 pounds of cotton balls than lead. This is because lead is much denser than cotton balls. To determine the density of a substance, divide the amount of that substance by the volume it takes to accommodate that amount. If you know the density, you can solve for the volume a certain mass will occupy or you can solve for the mass a certain volume will accommodate.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \quad \text{or} \quad \text{Volume} = \frac{\text{Mass}}{\text{Density}} \quad \text{or} \quad \text{Mass} = \text{Volume x Density}$$

Example 12: What is the density of 908 g of cadmium metal that occupies 105 mL?

$$Density = \frac{Mass}{Volume} = \frac{908 \text{ g}}{105 \text{ mL}} = \textbf{8.65 g/mL}$$

The important thing is to first make sure you are dealing with the correct units before solving for density. Mass should be in grams, and volume should either be in mL or cm^3 since 1 mL = 1 cm^3. This means density should always be in g/mL or g/cm^3. You may have to do some basic unit converting first.

Example 13: What mass of Ne gas with a density of 0.00090 g/mL occupies a volume of 70 ounces?

Known relationships: 32 oz = 1 qt (quart), 1.0567 qt = 1 L, 1 mL = 10^{-3} L (or 0.001 L)

$$70 \, \cancel{oz} \times \frac{1 \, \cancel{qt}}{32 \, \cancel{oz}} \times \frac{1 \, \cancel{L}}{1.0567 \, \cancel{qt}} \times \frac{1 \text{ mL}}{10^{-3} \, \cancel{L}} = 2069.54 \text{ mL}$$

$$Mass = Volume \times Density = 2069.54 \, \cancel{mL} \times 0.0009 \text{ g/}\cancel{mL} = \textbf{2 g Ne}$$

Sometimes density can be incorporated into other types of problems, for example, chemical quantity conversions. In Example 14, you must find the mass from the density and volume before the number of moles can be found.

Example 14: How many moles are there in a sample of titanium metal, density 4.54 g/mL, that can occupy a volume of 25.0 mL?

Known relationships: 1 mol = molar mass of Ti = 48 g

$$Mass = Volume \times Density = 4.54 \text{ g/}\cancel{mL} \times 25.0 \, \cancel{mL} = 113.5 \text{ g}$$

$$113.5 \, \cancel{g} \text{ Ti} \times \frac{1 \text{ mol}}{48 \, \cancel{g}} = \textbf{2.36 mol Ti}$$

Percent Composition & Empirical & Molecular Formula Determination

The molecular formula of a compound is the chemical formula that specifies the number of atoms of each element in one molecule of that compound and specifies the number of moles of each element in one mole of that compound. The molecular formula for sugar glucose is $C_6H_{12}O_6$. This means that for every 1 molecule of glucose, there are 6

atoms of C, 12 atoms of H, and 6 atoms of O. This also means that for every 1 mole of glucose, there are 6 moles of C, 12 moles of H, and 6 moles of O. The empirical formula of a compound is the smallest whole number ratio of the molecular formula and gives the relative number of atoms of each element in the compound. The empirical formula for glucose is CH_2O. This means that the ratio of C, H, and O in glucose is 1 : 2 : 1. The percent composition of a compound gives the percentage of each element in that compound. The percent composition of K_2CO_3, for example, is 56.52% K, 8.70% C, and 34.78% O. Knowing percent composition can help you determine empirical formula, and knowing empirical formula can help you determine molecular formula, so we will go over how to determine each of these.

To determine the percent composition of a compound:
1) Convert moles of each element into grams.
2) Add the mass in grams of each element to get a total mass.
3) Divide each element's mass in grams by the total mass in grams and multiply by 100%.
4) Check your answer by making sure that the sum of the percentages equals about 100%. The sum may not equal exactly 100% depending on how much you rounded your answers.

Example 15: What is the percent composition of K_2CO_3?

$$2 \text{ mol K} \times \frac{39 \text{ g}}{1 \text{ mol}} = 78 \text{ g K} \qquad 1 \text{ mol C} \times \frac{12 \text{ g}}{1 \text{ mol}} = 12 \text{ g C}$$

$$3 \text{ mol O} \times \frac{16 \text{ g}}{1 \text{ mol}} = 48 \text{ g O} \qquad \text{Total Mass} = 48 \text{ g} + 12 \text{ g} + 78 \text{ g} = 138 \text{ g}$$

$$\% \text{ K} = \frac{78 \text{ g}}{138 \text{ g}} \times 100\% = 56.52\% \quad \% \text{ C} = \frac{12 \text{ g}}{138 \text{ g}} \times 100\% = 8.70\%$$

$$\% \text{ O} = \frac{48 \text{ g}}{138 \text{ g}} \times 100\% = 34.78\% \quad \text{Check: } 8.70\% + 34.78\% + 56.52\% = 100\%$$

Percent Composition: 56.52% K, 8.70% C, and 34.78% O

You can find the empirical formula of a compound if you know the percent composition of that compound. To do so:

1) Designate the total amount of sample as 100 g so that the percentage of each element will equal the amount in grams of each element.
2) Convert grams of each element into moles.

3) Divide each number of moles by the smallest number of moles. In example 16, the smallest number of moles is Cr, so moles of K, moles of Cr, and moles of O will each be divided by moles of Cr.
4) Put the numbers obtained from dividing by the smallest number into a ratio.
5) Multiply that ratio by the smallest whole number that will give a ratio of close to whole numbers.

Example 16: What is the empirical formula of a compound containing 26.57% K, 35.36% Cr, and 38.07% O?

$$26.57 \text{ g K} \times \frac{1 \text{ mol}}{39 \text{ g}} = 0.6813 \text{ mol K} \qquad 35.36 \text{ g Cr} \times \frac{1 \text{ mol}}{52 \text{ g}} = 0.6800 \text{ mol Cr}$$

$$38.07 \text{ g O} \times \frac{1 \text{ mol}}{16 \text{ g}} = 2.379 \text{ mol O} \qquad K: \frac{0.6813 \text{ mol}}{0.6800 \text{ mol}} = 1.002$$

$$Cr: \frac{0.6800 \text{ mol}}{0.6800 \text{ mol}} = 1.000 \qquad O: \frac{2.379 \text{ mol}}{0.6800 \text{ mol}} = 3.499$$

$$K : Cr : O = (1.002 : 1.000 : 3.499) \times 2 = 2.004 : 2.000 : 6.998 = 2 : 2 : 7$$

Empirical formula: $K_2Cr_2O_7$

The molecular formula of a compound can be determined if you know the empirical formula and the molar mass of the compound. To do so:

1) Convert moles of each element into grams.
2) Add the grams of each element together to get a total empirical formula weight.
3) Divide the molar mass by the empirical formula weight.
4) Multiply the subscripts of each element by the number obtained from step 3.

Example 17: What is the molecular formula of ethylene glycol if its empirical formula is CH_3O and its molar mass is 62.1 g?

$$1 \text{ mol C} \times \frac{12 \text{ g}}{1 \text{ mol}} = 12 \text{ g C} \quad 3 \text{ mol H} \times \frac{1 \text{ g}}{1 \text{ mol}} = 3 \text{ g H} \quad 1 \text{ mol O} \times \frac{16 \text{ g}}{1 \text{ mol}} = 16 \text{ g O}$$

Empirical formula weight: $16 \text{ g} + 12 \text{ g} + 3 \text{ g} = 31 \text{ g}$

$$\frac{62.1 \text{ g}}{31 \text{ g}} = 2 \qquad C : H : O = (1 : 3 : 1) \times 2 = 2 : 6 : 2$$

Molecular formula: $C_2H_4O_2$

Stoichiometric Calculations

   Stoichiometric calculations are calculations that involve the relationship between the quantity of products and of reactants of a chemical equation. With a balanced chemical equation, you can convert between moles of reactants and products using the coefficients. To convert from the amount of one participant in a reaction to another participant in a reaction:

1) Make sure the reaction given is balanced (for balancing review, see our *Balancing Chemical Equations* handout).
2) Convert the amount of the compound given to moles if it is not already.
3) Convert to the desired compound by doing a mole to mole ratio using the coefficients of the two compounds you're converting between. Treat the mole to mole ratio like a conversion factor (known relationship) from basic unit conversion problems. In example 14, 4 mol NH_3 = 3 mol O_2, 4 mol NH_3 = 2 mol N_2, 4 mol NH_3 = 6 mol H_2O, 3 mol O_2 = 2 mol N_2, 3 mol O_2 = 6 mol H_2O, and 2 mol N_2 = 6 mol H_2O are all conversion factors that can be used to convert between compounds in the reaction $4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$. Remember to put the compound you want to end up with on the top and the compound you started with on the bottom of your ratio so that the compound you started with can be canceled out just like a unit.
4) If the problem asks for the end compound to be in some unit other than moles, do that conversion.

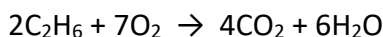Example 18: How many moles of H_2O are produced if 40 moles of NH_3 are used?

$$4NH_3 + 3O_2 \longrightarrow 2N_2 + 6H_2O$$
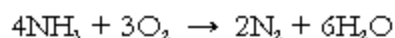
Known relationship: 4 mol NH_3 = 6 mol H_2O

Mole to mole ratio: $\dfrac{6 \text{ mol } H_2O}{4 \text{ mol } NH_3}$

$40 \text{ mol } NH_3 \times \dfrac{6 \text{ mol } H_2O}{4 \text{ mol } NH_3} = \textbf{60 mol } H_2O$

Example 19: If 16 grams of O_2 are used in the following reaction, how many grams of C_2H_6 are used?

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

Known relationships: 2 mol C_2H_6 = 7 mol O_2, 1 mol = molar mass of C_2H_6 = 30 g, and 1 mol = molar mass of O_2 = 32 g

Mole to mole ratio: $\dfrac{2\ \text{mol }C_2H_4}{7\ \text{mol }O_2}$

$16\ \text{g }\cancel{O_2} \times \dfrac{1\ \cancel{\text{mol}}}{32\ \text{g}} \times \dfrac{2\ \cancel{\text{mol }C_2H_4}}{7\ \cancel{\text{mol }O_2}} \times \dfrac{30\ \text{g}}{1\ \cancel{\text{mol}}} = \textbf{4.3 g } C_2H_4$

Limiting Reagents

When chemical reactions are performed for industrial purposes, chemical reactants are often used in specific amounts so that there won't be any leftover reactants. For instance, in the following reaction, if you react 5.0 g of Cs with 3.0 g of Br_2, both reactants will be completely used up and there will be no leftovers of either Cs or Br_2.

$$2Cs + Br_2 \rightarrow 2CsBr$$
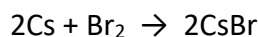
However, if you mix 5.0 g of Cs with 2.5 g of Br_2, there will be leftover Cs since there wasn't enough Br_2 to completely consume the Cs. In this situation, Br_2 is called the limiting reactant or limiting reagent. Note that Br_2 won't always be the limiting reagent for this reaction. Whether Cs or Br_2 or neither is the limiting reagent depends on the amount of each reactant you begin with. The following steps can help you take a reaction and determine which, if any, reactant is limiting, determine how much of the non-limiting reactant will be leftover, and determine the amount of each product formed.

1) Make sure the given reaction is balanced.
2) Convert the amount of each reactant to moles if they are not in moles already.
3) Start with the given amount of one of the reactants in moles, reactant A, and convert to moles of the other reactant, reactant B, using a mole to mole ratio of the reactants' coefficients as your conversion factor (see pg. 8 on stoichiometric calculations for review on mole to mole ratios). This will give you the amount of B needed to completely consume the A.
4) Compare the actual amount of B you have with the amount you would need to consume all the A to determine which reactant, if any, is limiting. If the actual amount of B is less than is needed to consume all the A, B will be limiting. If the actual amount of B is more than is needed to consume all the A, A will be limiting. If the actual amount of B equals the amount needed to consume all the A, neither reactant will be limiting.
5) If reactant A is limiting, you can determine the leftover amount of B by subtracting the actual amount of B used by the amount of B needed to completely consume A. If reactant B is limiting, you will first need to determine how much of A is needed to consume B (see step 3). Next, subtract the actual amount of A used by the amount that you just calculated (the amount needed to consume B). If neither reactant is limiting, there will be no leftover of either reactant.

6) To calculate the amount of a product that will be formed, use the amount of limiting reagent used in moles to convert to moles of that product using a mole to mole ratio.

7) If the problem asks for the product to be in a unit besides moles, or if the reactant amounts are given in a unit besides moles, convert the products to that unit.

Example 20: A solution containing 2.00 g of $C_5H_{11}OH$ was added to a solution containing 2.00 g of O_2. Which reagent is the limiting reagent? How much of the reagent that is not limiting will be left over?

$$2C_5H_{11}OH + 15O_2 \rightarrow 10CO_2 + 12H_2O$$
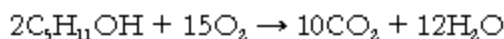
$$2.00 \text{ g } C_5H_{11}OH \times \frac{1 \text{ mol}}{88 \text{ g}} = 2.27 \times 10^{-2} \text{ mol } C_5H_{11}OH$$

$$2.00 \text{ g } O_2 \times \frac{1 \text{ mol}}{16 \text{ g}} = 0.125 \text{ mol } O_2$$

$$0.125 \text{ mol } O_2 \times \frac{2 \text{ mol } C_5H_{11}OH}{15 \text{ mol } O_2} = 1.67 \times 10^{-2} \text{ mol } C_5H_{11}OH$$

Since the amount of $C_5H_{11}OH$ that will react with all the O_2, 1.67×10^{-2} mol is less than the amount of $C_5H_{11}OH$ actually used, 2.27×10^{-2} mol, there will be leftover $C_5H_{11}OH$, and **O_2 will be the limiting reagent**. This could also be determined if you determined how much O_2 would be needed to react with all the $C_5H_{11}OH$. That amount would be more than the available amount of O_2.

$$2.27 \times 10^{-2} \text{ mol } C_5H_{11}OH \times \frac{15 \text{ mol } O_2}{2 \text{ mol } C_5H_{11}OH} = 0.170 \text{ mol } O_2$$

Since O_2 is the limiting reagent, all of it will be consumed, and there will be leftover $C_5H_{11}OH$.

$$2.27 \times 10^{-2} \text{ mol} - 1.67 \times 10^{-2} \text{ mol} = 6.00 \times 10^{-3} \text{ mol } C_5H_{11}OH$$

$$6.00 \times 10^{-3} \text{ mol } C_5H_{11}OH \times \frac{88 \text{ g}}{1 \text{ mol}} = \textbf{0.528 g } C_5H_{11}OH \textbf{ leftover}$$

Example 21: How many grams of $Ca_3(PO_4)_2$ and grams of KCl can be produced by mixing 5.00 g of $CaCl_2$ with 8.00 g of K_3PO_4?

$$3CaCl_2 + 2 K_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6KCl$$
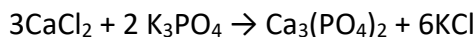
$$5.00 \text{ g } CaCl_2 \times \frac{1 \text{ mol}}{110 \text{ g}} = 0.0455 \text{ mol } CaCl_2$$

$$8.00 \text{ g } K_3PO_4 \times \frac{1 \text{ mol}}{212 \text{ g}} = 0.0377 \text{ mol } K_3PO_4$$

$$0.0377 \text{ mol } K_3PO_4 \times \frac{3 \text{ mol } CaCl_2}{2 \text{ mol } K_3PO_4} = 0.0566 \text{ mol } CaCl_2$$

$CaCl_2$ is the limiting reagent since there is not enough of it for all the K_3PO_4 being used to be consumed. 0.0566 mol $CaCl_2$ would be needed for all the K_3PO_4 to be consumed, but we only have 0.0455 mol $CaCl_2$.

$$0.0455 \text{ mol } CaCl_2 \times \frac{1 \text{ mol } Ca_3(PO_4)_2}{3 \text{ mol } CaCl_2} \times \frac{310 \text{ g}}{1 \text{ mol}} = 4.70 \text{ g } Ca_3(PO_4)_2$$

$$0.0455 \text{ mol } CaCl_2 \times \frac{6 \text{ mol } KCl}{3 \text{ mol } CaCl_2} \times \frac{74 \text{ g}}{1 \text{ mol}} = 6.73 \text{ g } KCl$$

Percent Yield

The theoretical yield of a reaction is the amount of product that was calculated to be produced. Example 21 shows how to calculate the theoretical yield of a product or products. Percent yield is the percentage of the theoretical yield that was actually obtained when the chemical reaction was performed. To calculate the percent yield, use the following formula:

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

Example 22: The theoretical yield of H_2O in the combustion of CH_4 was calculated to be 55 g, but when the reaction was actually performed, the amount of H_2O yielded was 51.7 g. What is the percent yield?

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

$$\text{Percent Yield} = \frac{51.7 \text{ g}}{55 \text{ g}} \times 100\% = 94\%$$

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- *[Periodic Table](#)* – Gives a traditional periodic table with atomic masses and atomic numbers with additional trends and values such as electronegativity and ionization energy in one convenient handout. This handout is available in color and in black/white, but the color table is suggested since it is easier to read.
- *[Balancing Chemical Equations](#)* – Gives step-by-step instructions on how to balance chemical equations and contains many practice problems to help you perfect your balancing.
- *[Naming Compounds](#)* – Gives step-by-step instructions on how to name ionic compounds, binary molecular compounds, and acids with practice problems.
- *[Aqueous Reactions](#)* – Explains the following types of aqueous reactions: precipitation, acid/base (neutralization) , and oxidation/reduction. Also gives helpful examples and practice problems.
- *[Significant Figure Rules](#)* – Explains the rules of determining significant figures and how to use these rules in addition, subtraction, multiplication, division, and converting. Also has practice problems.
- *[Exponents, Radicals, and Scientific Notation](#)* – Under the math handout section, this handout includes how to turn a number into scientific notation and vice versa with practice problems.

* Information for this handout was obtained from the following sources:

- Brown, LeMay, & Bursten. *Chemistry: The Central Science*. 9th Ed. Pearson Education, Inc. 2003.
- Gilbert, Kriss, & Davies. *Chemistry: The Science in Context*. 1st Ed. W. W. Norton & Company. 2004.
- Rosenberg & Epstein. *Schaum's Outlines: College Chemistry*. 8th Ed. McGraw-Hill. 19