Solutions 4a (Chapter 4 problems) **Chem151 [Kua]**

4.10 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 the equation, start by determining the number of atoms on each side of the chemical equation.

(a) $P_4 + \mathbf{12}\ Na \rightarrow \mathbf{4}\ Na_3P$

(b) $Na_3P + \mathbf{3}\ H_2O \rightarrow PH_3 + \mathbf{3}\ NaOH$

(c) $\mathbf{4}\ PH_3 + \mathbf{8}\ O_2 \rightarrow P_4O_{10} + \mathbf{6}\ H_2O$

(d) $P_4O_{10} + \mathbf{6}\ H_2O \rightarrow \mathbf{4}\ H_3PO_4$

4.16 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 center on the mole:

$$375\ \text{kg}\ C_6H_{12}\left(\frac{10^3\ \text{g}}{1\ \text{kg}}\right)\left(\frac{1\ \text{mol}}{84.16\ \text{g}}\right)\left(\frac{2\ \text{mol}\ C_6H_{10}O_4}{2\ \text{mol}\ C_6H_{12}}\right)\left(\frac{146.14\ \text{g}}{1\ \text{mol}}\right)\left(\frac{1\ \text{kg}}{10^3\ \text{g}}\right)$$

$$= 651\ \text{kg}\ C_6H_{10}O_4$$

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

$$\text{Theoretical yield (if 100\% efficient)} = 3.50\ \text{kg}\left(\frac{100\%}{76.5\%}\right) = 4.575\ \text{kg};$$

$$\text{Amt required: } 4.575\ \text{kg AA}\left(\frac{1\ \text{mol}}{146.14\ \text{g}}\right)\left(\frac{2\ \text{mol}\ C_6H_{12}}{2\ \text{mol AA}}\right)\left(\frac{84.16\ \text{g}}{1\ \text{mol}}\right) = 2.63\ \text{kg}\ C_6H_{12}$$

4.78 The problem asks about amount of a product that forms from a given amount of reactant. First balance the chemical equation. The reaction is:

$C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$

8C + 18H + 2O → 1C + 2H + 3O

Give CO_2 a coefficient of 8, and H_2O a coefficient of 9 to balance C and H:

$C_8H_{18} + O_2 \rightarrow \mathbf{8}\ CO_2 + \mathbf{9}\ H_2O$

8C + 18H + 2O → 8C + 18H + 25O

There are 25 O on the product side, so give O_2 a coefficient of 25/2 to balance O and multiply through by 2 to eliminate the fraction:

$2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$

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

(a) $n_{CO_2} = 3.5\text{ g}\left(\frac{1\text{ mol}}{114.2\text{ g}}\right)\left(\frac{16\text{ mol } CO_2}{2\text{ mol } C_8H_{18}}\right) = 0.25\text{ mol } CO_2$

(b) $\# = n\,N_A = 0.25\text{ mol}\left(\frac{6.022 \times 10^{23}\text{ molecules}}{1\text{ mol}}\right) = 1.5 \times 10^{23}\text{ molecules}$

(c) $m = n\,MM = 0.25\text{ mol}\left(\frac{44.01\text{ g}}{1\text{ mol}}\right) = 11\text{ g}$

4.28 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 10^3 mol amounts. The balanced equation is given in the problem.

Calculate the initial amounts:

$1.50 \times 10^3\text{ kg propene}\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol}}{42.08\text{ g}}\right) = 35.6 \times 10^3\text{ mol;}$

$6.80 \times 10^2\text{ kg } NH_3\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol}}{17.04\text{ g}}\right) = 39.9 \times 10^3\text{ mol;}$

$1.92 \times 10^3\text{ kg } O_2\left(\frac{10^3\text{ g}}{1\text{ kg}}\right)\left(\frac{1\text{ mol}}{32.00\text{ g}}\right) = 60.0 \times 10^3\text{ mol;}$

Next construct an amounts table:

Reaction	$2\ C_3H_6$ +	$2\ NH_3$ +	$3\ O_2$ →	$2\ C_3H_3N$	+ $6\ H_2O$
Amt (10^3 mol)	35.6	39.9	60.0	0	0
kmol/coeff	17.8 (LR)	19.95	20.0		
Change (10^3 mol)	- 35.6	-35.6	$-(\frac{3}{2})(35.6)$	+35.6	+3(35.6)
Final (10^3 mol)	0.0	4.3	6.6	35.6	106.8

The mass of acrylonitrile that could be produced is:

$$35.6 \times 10^3\text{ mol}\left(\frac{53.06\text{ g}}{1\text{ mol}}\right)\left(\frac{10^{-3}\text{ kg}}{1\text{ g}}\right) = 1.89 \times 10^3\text{ kg.}$$

4.30 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 10^3 kilograms (10^3 kg = 10^6 g), so it will be convenient to work with 10^6 mol amounts. See the answers to problem 4.6 for the balanced equations.

(a) Begin by calculating the initial amounts:

$$7.50 \text{ x } 10^3 \text{ kg } NH_3 \left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol}}{17.03 \text{ g}}\right) = 0.440 \text{ x } 10^6 \text{ mol;}$$

$$7.50 \text{ x } 10^3 \text{ kg } O_2 \left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right)\left(\frac{1 \text{ mol}}{32.00 \text{ g}}\right) = 0.234 \text{ x } 10^6 \text{ mol;}$$

Next construct the table of amounts using the balanced equation from 4.6:
which is $4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$

Reaction	$4\ NH_3$ +	$5\ O_2$ →	4 NO +	$6\ H_2O$
Amt (10^6 mol)	0.440	0.234	0	0
10^6 mol/coeff	0.110	0.0468 (LR)		
Change (10^6 mol)	$-(\frac{4}{5})(0.234)$	-0.234	$+(\frac{4}{5})(0.234)$	$+(\frac{6}{5})(0.234)$
Final (10^6 mol)	0.253	0	0.187	0.281

Now using the results from the amounts table, determine the mass of products produced:

Mass NO: $0.187 \text{ x } 10^6 \text{ mol}\left(\frac{30.01 \text{ g}}{1 \text{ mol}}\right)\left(\frac{1 \text{ ton}}{10^6 \text{ g}}\right) = 5.61$ metric ton or $5.61 \text{ x } 10^3$ kg

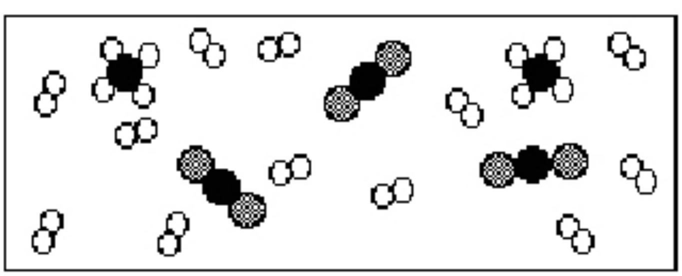
Mass H_2O: $0.281 \text{ x } 10^6 \text{ mol}\left(\frac{18.02 \text{ g}}{1 \text{ mol}}\right)\left(\frac{1 \text{ ton}}{10^6 \text{ g}}\right) = 5.06$ metric ton or $5.06 \text{ x } 10^3$ kg

4.76 Molecular pictures and limiting reactant calculations require a balanced chemical equation for the reaction under consideration. For this process, the balanced equation is

$$CH_4 + 2\ H_2O \rightarrow CO_2 + 4\ H_2.$$

(a) There are 6 H_2O molecules and 5 CH_4 molecules in the picture. Divide number of molecules by coefficient to see which reactant is limiting: $\frac{5}{1} = 5$ for CH_4, $\frac{6}{2} = 3$ for H_2O. The smaller number identifies the limiting reactant, H_2O.

(b) The new molecular picture must show that all the H_2O has been consumed and CO_2 and H_2 have been produced, but the number of atoms of each element must be the same as before: 5 atoms of C, 32 of H, and 6 of O. Six H_2O molecules react with three CH_4 molecules to form three molecules of CO_2 and 12 of H_2:



(c) Construct an amounts table for the stoichiometric calculations:

Reaction	CH_4 +	2 H_2O →	CO_2	4 H_2
Amt (mol)	5	6	0	0
Change (mol)	-6/2	-6	+6/2	+(4/2)(6)
Final (mol)	2	0	3	12

Use the results from the amounts table to calculate the desired masses:

$$2.0 \text{ mol } CH_4\left(\frac{16.04\text{ g}}{1\text{ mol}}\right) = 32.08 \text{ g } CH_4 \text{ remaining}$$

$$3 \text{ mol } CO_2\left(\frac{44.01\text{ g}}{1\text{ mol}}\right) = 132.0 \text{ g } CO_2 \text{ produced}$$

$$12 \text{ mol } H_2\left(\frac{2.016\text{ g}}{1\text{ mol}}\right) = 24.19 \text{ g } H_2 \text{ produced}$$

4.98 Because amounts of each reactant are provided, this is a limiting reactant problem, for which it is convenient to use an amounts table. We are not asked about the second product, CO, so its amounts do not need to be tabulated.

Calculate the initial amounts:

$$SiO_2\text{:}\quad 75.0 \text{ g}\left(\frac{1\text{ mol}}{60.09\text{ g}}\right) = 1.25 \text{ mol}$$

$$C\text{:}\quad 75.0 \text{ g}\left(\frac{1\text{ mol}}{12.01\text{ g}}\right) = 6.24 \text{ mol}$$

$$Cl_2\text{:}\quad 75.0 \text{ g}\left(\frac{1\text{ mol}}{70.90\text{ g}}\right) = 1.06 \text{ mol}$$

Reaction	SiO_2 +	2 C +	2 Cl_2 →	$SiCl_4$
MM (g/mol)	60.09	12.01	70.90	169.89
Amt (mol)	1.25	6.24	1.06	0
mol/coeff	1.25	3.12	0.530 (LR)	
Change (mol)	-(1/2)(1.06)	- 1.06	-1.06	+(1/2)(1.06)
Final (mol)	0.72	5.18	0	0.530

The theoretical yield is $0.530 \text{ mol}\left(\frac{169.89\text{ g}}{1\text{ mol}}\right) = 90.0 \text{ g}$

The actual production is $90.0 \text{ g}\left(\frac{95.7\%}{100\%}\right) = 86.1 \text{ g}$.

To determine the amount of each reactant that remains unreacted, redo the amounts table using the correct change:

$$86.1\text{ g}\left(\frac{1\text{ mol}}{169.9\text{ g}}\right) = 0.507\text{ mol}$$

Reaction	SiO_2 +	2 C +	2 Cl_2 →	$SiCl_4$
MM (g/mol)	60.09	12.01	70.90	169.89
Amt (mol)	1.25	6.24	1.058	0
Change (mol)	-0.507	-2(0.507)	-2(0.507)	0.507
Final (mol)	0.74	5.23	0.046	0.507

Calculate the final amounts:

SiO_2: $0.74\text{ mol}\left(\frac{60.09\text{ g}}{1\text{ mol}}\right) = 44.5\text{ g}$

C: $5.23\text{ mol}\left(\frac{12.01\text{ g}}{1\text{ mol}}\right) = 62.8\text{ g}$

Cl_2: $0.046\text{ mol}\left(\frac{70.90\text{ g}}{1\text{ mol}}\right) = 3.26\text{ g}$

4.40 A net ionic equation shows which ions combine to give new products. Spectator ions do not appear in the net equation, and charges must be balanced. Start by determining what ions are present in solution. Then use the Solubility Guidelines from your text to determine what precipitates can form.

(a) The ions present are: Ag^+, NO_3^-, Li^+, and Cl^-. So the possible combinations are: $AgNO_3$, AgCl, LiCl, and $LiNO_3$. Guidelines 1 and 2 predict that salts of lithium and nitrate are soluble. Since AgCl is not soluble by the guidelines, it must be an insoluble salt:

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$; spectator ions: Li^+, NO_3^-

(b) The ions present are: Mg^{2+}, SO_4^{2-}, Na^+, and PO_4^{3-}. Guidelines 1 and 3 state that all salts of sodium and most of sulfate are soluble. $Mg_3(PO_4)_2$ is not covered by guidelines 1 or 2, and is not in the exceptions of 4 and 5, thus by guideline 3 it is insoluble:

$3\,Mg^{2+}(aq) + 2\,PO_4^{3-}(aq) \rightarrow Mg_3(PO_4)_2(s)$; spectator ions: Na^+, SO_4^{2-}

(c) The ions present are: Ba^{2+}, SO_4^{2-}, Na^+, and OH^-. Guidelines 1 and 3 state that all salts of sodium and most of sulfate are soluble. $Ba(OH)_2$ is soluble by guideline 5. $BaSO_4$ is an exception in guideline 4 and is thus insoluble:

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$; spectator ions: Na^+, OH^-

(d) The ions present are: Al^{3+}, Cl^-, K^+, and OH^-. Guidelines 1 and 2 state that all salts of potassium and most of chloride are soluble. $Al(OH)_3$ is not soluble by the guidelines and is thus insoluble by guideline 3:

$Al^{3+}(aq) + 3\ OH^-(aq) \rightarrow Al(OH)_3(s)$; spectator ions: K^+, Cl^-.

4.42 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 product mass. Start by determining the balanced net ionic reaction using the solubility guidelines.

The ions present are Pb^{2+}, NO_3^-, NH_4^+, and Cl^-. Guideline 1 and 2 state that all salts containing ammonium and nitrate are soluble. $PbCl_2$ is an exception of guideline 4 and thus insoluble:

$2\ Pb^{2+} + Cl^- \rightarrow PbCl_2$.

Calculations of initial amounts:

For Pb^{2+}, 75.0 ml = 0.0750 L and 0.750 M = 0.750 mol Pb^{2+}/L

$$0.0750\ L\left(\frac{0.750 mol}{1\ L}\right) = 0.5625 \times 10^{-1} mol\ Pb^{2+}$$

For Cl^-, 125 mL = 0.125 L and 0.855 M = 0.855 mol Cl^-/L

$$0.125\ L\left(\frac{0.855 mol}{1\ L}\right) = 1.069 \times 10^{-1} mol\ Cl^-$$

Next set up the amounts table:

Reaction	Pb^{2+} +	$2\ Cl^-$ →	$PbCl_2(s)$
Start (10^{-1} mol)	0.5625	1.069	0
Mol/coeff	0.5625	0.5345 (LR)	
Change (10^{-1} mol)	-1.069/2	-1.069	+1.069/2
Final (10^{-1} mol)	0.0280	0	0.5344

The mass of solid that forms is: $5.344 \times 10^{-2}\ mol\left(\frac{278.1\ g}{1\ mol}\right) = 14.9\ g.$

The ions remaining in solution are the excess Pb^{2+} and the spectator ions, NO_3^- and NH_4^+.

4.60 The first two parts of this problem ask about an ionic reaction. First identify the species present in the solution, then use the identified products to find the net ionic reaction and the spectator ions. The species present are H_2O, NH_4^+, SO_4^{2-}, Ca^{2+}, and NO_3^-. All ammonium and nitrate salts are soluble, so these ions do not precipitate.

(a) The solid product is $CaSO_4$, and the net ionic reaction is $Ca^{2+} + SO_4^{2-} \rightarrow CaSO_4(s)$;

(b) The spectator ions are NH_4^+ and NO_3^-

(c) The second two parts of this problem involve stoichiometric calculations. 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 product formed and the final concentrations of all species.

Calculations of initial amounts:

Ca^{2+}: $n = 0.1000\text{ L}\left(\frac{1.50\text{ mol}}{1\text{ L}}\right) = 0.150\text{ mol}$

SO_4^{2-}: $n = 0.0750\text{ L}\left(\frac{3.00\text{ mol}}{1\text{ L}}\right) = 0.225\text{ mol}$

The balanced equation shows that the starting materials react in a 1:1 mole ratio, so we can identify the limiting reactant by inspection; the limiting reactant is Ca^{2+}. Here is the complete table of amounts:

Reaction	Ca^{2+} +	SO_4^{2-} $\rightarrow$	$CaSO_4$
Start (mol)	0.150	0.225	0
Change (mol)	-0.150	-0.150	+0.150
Final (mol)	0	0.075	0.150

Obtain the mass of precipitate from the final amount of $CaSO_4$ from the table:

MM = 40.08 g/mol + 32.07 g/mol + 4(16.00 g/mol) = 136.15 g/mol

The mass of precipitate is $0.150\text{ mol}\left(\frac{136.15\text{ g}}{1\text{ mol}}\right) = 20.4\text{ g}$;

(d) To find ion concentrations in the final solution, first determine how many moles of each ion remain in solution, then divide by the final volume of the solution, which is:

V_{final} = 100.0 mL + 75.0 mL = 175.0 mL or 0.1750 L

The amounts of the two spectator ions are unaffected by the reaction:

$$[NH_4^+] = \left(\frac{0.225\text{ mol }(NH_4)_2SO_4}{0.1750\text{ L}}\right)\left(\frac{2\text{ mol }NH_4^+}{1\text{ mol }(NH_4)_2SO_4}\right) = 2.57\text{ M};$$

$$[NO_3^-] = \left(\frac{0.150\text{ mol }Ca(NO_3)_2}{0.1750\text{ L}}\right)\left(\frac{2\text{ mol }NO_3^-}{1\text{ mol }Ca(NO_3)_2}\right) = 1.71\text{ M};$$

The final moles of each reactant can be obtained from the amounts table in part c:

$[Ca^{2+}]$ =0 M; $[SO_4^{2-}] = \frac{0.075\text{ mol}}{0.175\text{ L}} = 0.43\text{ M}.$

4.46 This problem describes an acid-base reaction. We are asked to determine the mass of base required to completely neutralize the acid. The starting materials are HCl, a strong acid, and $Mg(OH)_2$, a weak base. In addition to water, the major species are:

$$H_3O^+, Cl^-, Mg(OH)_2$$

The presence of hydronium ions and a weak base together as major species will result in an acid-base reaction:

The balanced reaction is:

$$Mg(OH)_{2(s)} + 2\ H_3O^+(aq) \rightarrow Mg^{2+}(aq) + 4\ H_2O(l).$$

The problem gives information about the amount of acid present. This is a mol-mass conversion problem:

Calculation of initial amount of acid:

For H_3O^+: $0.125\ L\left(\frac{0.115\ mol}{1\ L}\right) = 1.44 \times 10^{-2}$ mol;

Now determine the mass of $Mg(OH)_2$ required:

$MM\ [Mg(OH)_2] = 24.31\ g/mol + 2(16.00\ g/mol) + 2(1.008\ g/mol) = 58.33\ g/mol$

$$m_{Mg(OH)_2} = 1.44 \times 10^{-2} mol\ H_3O^+\left(\frac{1\ mol\ Mg(OH)_2}{2\ mol\ H_3O^+}\right)\left(\frac{58.33\ g}{1\ mol}\right) = 0.420\ g\ Mg(OH)_2$$

4.50 This problem describes an acid-base titration. We are asked to determine the concentration of acid used to neutralize the base. The starting materials are HCl, a strong acid, and NaOH, a strong base. In addition to water, the major species are:

$$H_3O^+, Cl^-, Na^+, OH^-$$

The presence of hydronium ions and hydroxide ions together as major species will result in an acid-base reaction:

The balanced reaction is:

$$OH^-(aq) + H_3O^+(aq) \rightarrow 2\ H_2O(l).$$

Calculation of the amount of base added:

$$OH^-: 32.45\ mL\left(\frac{10^{-3}\ L}{1\ mL}\right)\left(\frac{0.0965\ mol}{1\ L}\right) = 3.13 \times 10^{-3}\ mol$$

Since this is a completed titration, the moles of base added to the solution will equal the moles of acid in the solution.

$$n(OH^-) = 3.13 \times 10^{-3}\ mol = n(H_3O^+)$$

The concentration of the acid solution is obtained by dividing the number of moles of H_3O^+ by the volume of the HCl solution.

$$V_{soln} = 0.01000\ L; \qquad [H_3O^+] = \frac{3.13 \times 10^{-3} mol}{0.01000\ L} = 0.313\ M$$

4.88 (a) The reaction is:

$$Mg + H_3O^+ \rightarrow Mg^{2+} + H_2 + H_2O.$$

$$1Mg + 3H + 1O \rightarrow 1Mg + 4H + 1O$$

Magnesium and oxygen are already balanced. Give H_3O^+ a coefficient of 2 to balance the charges and H_2O a coefficient of 2 to balance O. Then H is also balanced (6 on each side):

$$Mg + 2\,H_3O^+ \rightarrow Mg^{2+} + H_2 + 2\,H_2O.$$

(b and c)
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 amounts.

Calculations of initial amounts:

Mg: $1.215 \text{ g Mg}\left(\frac{1\text{ mol}}{24.305\text{ g}}\right) = 4.999 \times 10^{-2}\text{ mol Mg}$

$H_3O^+ = 0.0500\text{ L }(4.0\text{ M}) = 20 \times 10^{-2}\text{ mol } H_3O^+$

Use the balanced equation and the initial amounts to construct the amounts table:

Reaction	Mg +	$2\,H_3O^+ \rightarrow$	Mg^{2+} +	H_2
Amt (10^{-2} mol)	4.999	20.0	0	0
10^{-2} mol/coeff	4.999 (LR)	10.0		
Change (10^{-2} mol)	-4.999	-9.998	+4.999	+4.999
Final (10^{-2} mol)	0	10	4.999	4.999

(b) $m = n\,MM = 4.999 \times 10^{-2}\text{ mol}\left(\frac{2.0158\text{ g}}{1\text{ mol}}\right) = 0.1008\text{ g } H_2$ formed;

(c) The spectator ion concentration remains unchanged: $[Cl^-] = 4.0$ M;

$$[Mg^{2+}] = \frac{4.999 \times 10^{-2}\text{ mol}}{50.0 \times 10^{-3}\text{ L}} = 1.00\text{ M};$$

$$[H_3O^+] = \frac{1 \times 10^{-1}\text{ mol}}{50.0 \times 10^{-3}\text{ L}} = 2.0\text{ M}.$$

4.66 When determining reaction products, first identify the type of substances present to determine what kind of reaction can occur.

(a) $Ca(OH)_2$ is a strong base, HCl is a strong acid:

$OH^-(aq) + H_3O^+(aq) \rightarrow 2\,H_2O(l)$ (acid-base reaction);

(b) Ni displaces hydrogen gas from strong acids such as HCl:

$Ni(s) + 2\,H_3O^+(aq) \rightarrow Ni^{2+}(aq) + H_2(g) + 2\,H_2O(l)$ (redox reaction, but Ni also acts as a base, so its also an acid-base reaction)

(c) AgOH is an insoluble salt so its a precipitation reaction. However Ag^+ also acts as a weak acid (so this could also be an acid-base reaction).

$Ag^+(aq) + OH^-(aq) \rightarrow AgOH(s)$ (precipitation reaction);

(d) A strong acid and a strong base react to form water:

$H_3O^+(aq) + OH^-(aq) \rightarrow 2\ H_2O(l)$ (acid-base reaction);

(e) Ca is a reactive metal, which reacts with water to generate H_2 gas:

$Ca(s) + 2\ H_2O(l) \rightarrow Ca^{2+}(aq) + H_2(g) + 2\ OH^-(aq)$ (redox reaction, but Ca also acts as a base so this could be an acid-base reaction);

(f) NH_3 is a weak base, which reacts with strong acid:

$NH_3(aq) + H_3O^+(aq) \rightarrow H_2O(l) + NH_4^+(aq)$ (acid-base reaction).

4.102 The statement of the problem indicates what reactions take place:

$H_2S + O_2 \rightarrow SO_2 + H_2O$. H and S are balanced; give O a coefficient of 3/2 to balance O, then multiply through by 2 to clear fractions:

$$2\ H_2S + 3\ O_2 \rightarrow 2\ SO_2 + 2\ H_2O$$

$SO_2 + H_2S \rightarrow S + H_2O$. H is already balanced; give H_2S and H_2O coefficients of 2 to balance O and S a coefficient of 3 to balance S:

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

The second part of the problem asks for the amount of H_2S required to form a given amount of S. From the first reaction, each mole of SO_2 requires 1 mol of H_2S, so the amount of H_2S required is the sum of the moles of H_2S and SO_2 required in the second reaction. From the second reaction, 3 mol S requires 2 + 1 = 3 mol ($H_2S + SO_2$), so $n_{H_2S} = n_S$:

$$m_{H_2S} = 1.25\ \text{kg}\left(\frac{1\ \text{mol}}{32.07\ \text{g}}\right)\left(\frac{1\ \text{mol}\ H_2S}{1\ \text{mol}\ S}\right)\left(\frac{34.09\ \text{g}}{1\ \text{mol}}\right) = 1.33\ \text{kg}.$$