Chapter 5

Solutions to Supplementary *Check for Understanding* Problems

Moles and Molar Mass

1. Indicate the appropriate quantity for each of the following.

a) A mole of N atoms contains _________________ atoms.
b) A mole of N_2 molecules contains _________________ molecules.
c) A mole of N_2 molecules contains _________________ atoms.
d) A mole of N atoms has a mass of _________________ grams.
e) A mole of N_2 molecules has a mass of _________________ grams.

Answers:
a) 6.022×10^{23} atoms

b) 6.022×10^{23} molecules

c) 1.2044×10^{24} atoms

d) 14.01 g

e) 28.02 g

Solutions

a) A mole of anything always contains 6.022×10^{23} items of that material.

b) A mole of anything always contains 6.022×10^{23} items of that material.

c) Since a mole of N_2 molecules contains 6.022×10^{23} molecules of N_2 and there are 2 atoms of N per molecule of N_2, the total number of N atoms is given by:

$$\frac{6.022 \times 10^{23}\ \cancel{\text{N}_2\text{ molecules}}}{\text{mol N}_2} \times \frac{2\text{ atoms N}}{1\ \cancel{\text{molecule N}_2}} = \frac{1.2044\ 10^{24}\text{ atoms N}}{\text{mol N}_2}$$

d) The mass in grams of a mole of atoms of any element (its molar mass) is numerically equally to the weighted average atomic mass in atomic mass units. Since the atomic weight of nitrogen is 14.01, its weighted atomic mass is 14.01 u and a mole of nitrogen atoms weighs 14.01 g.

e) The mass in grams of a mole of N_2 molecules (its molar mass) is obtained by summing the molar masses of the atoms in the chemical formula. Since the molar mass of N is 14.01 g/mol (see part d), the molar mass of N_2 is:

$$\frac{14.01\,\text{g}}{\cancel{\text{mol N}}} \times \frac{2\,\cancel{\text{mol N}}}{1\,\text{mol N}_2} = \frac{28.01\,\text{g}}{\text{mol N}_2}$$

2. a) What is the mass of 725 sodium atoms in atomic mass units?

Answer: 1.67×10^3 u

Solution

The numerical value of the atomic weight of sodium (22.99) refers to the mass of a single sodium atom in atomic mass units. Applying this yields:

$$725\,\cancel{\text{Na atoms}} \times \frac{22.99\,\text{u}}{\cancel{\text{Na atom}}} = 1.67 \times 10^3\,\text{u}$$

b) What is the mass of 725 sodium atoms in grams?

Answer: 2.77×10^{-20} g

Solution

What we know: number of Na atoms

Desired answer: g Na

The solution map for this calculation is:

atom Na → mol Na → g Na

The conversion factor needed in the first step is the Avogadro constant expressed in the form $\frac{1\,\text{mol Na}}{6.022 \times 10^{23}\,\text{Na atoms}}$.

The conversion factor needed in the second step is the molar mass of sodium. The numerical value for the molar mass is obtained from the atomic weight of sodium (22.99) and is expressed in the form $\frac{22.99\,\text{g Na}}{1\,\text{mol Na}}$.

Putting these together yields:

$$725\ \cancel{\text{atoms Na}} \times \frac{1\ \cancel{\text{mol Na}}}{6.022 \times 10^{23}\ \cancel{\text{atoms Na}}} \times \frac{22.99\ \text{g Na}}{1\ \cancel{\text{mol Na}}} = 2.77 \times 10^{-20}\ \text{g Na}$$

3. How many atoms of an element are present in a sample of that element if the sample has a mass in grams numerically equal to the atomic weight of the element?

Answer: 6.022×10^{23} atoms

Solution

A mass in grams numerically equal to the atomic weight of an element is its molar mass and thus contains one mole of atoms of that element. Therefore, this sample will contain 6.022×10^{23} atoms.

4. Blackboard chalk is mostly calcium sulfate. How would you determine how many moles of calcium sulfate it takes to write your name in chalk on a blackboard?

Solution

The moles of a pure substance can be determined from the mass of the substance and its molar mass. The mass of the chalk used can be obtained by weighing the chalk before and after writing your name and calculating the difference between the masses. The molar mass of calcium sulfate can be obtained from its chemical formula ($CaSO_4$). The calculations will be:

g $CaSO_4$ used = initial mass of chalk - final mass of chalk

$$\cancel{\text{g CaSO}_4}\ \text{used} \times \frac{1\ \text{mol CaSO}_4}{136.15\ \cancel{\text{g CaSO}_4}} = \text{mol CaSO}_4\ \text{used}$$

5. What mass of zinc metal contains the same number of atoms as 16.1 grams of silver?

Answer: 9.76 g

Solution

What we know: g Ag; number of Ag atoms equals number of Zn atoms

Desired answer: g Zn

The solution map for this calculation is:

g Ag ➔ mol Ag ➔ mol Zn ➔ g Zn

The conversion factor needed in the first step is the molar mass of Ag in the form $\frac{1\,\text{mol Ag}}{107.9\,\text{g Ag}}$.

The problem indicates that the number of atoms of Zn is the same as the number of atoms of Ag. Therefore, mol Zn = mol Ag. This relationship can be applied as the conversion factor in the second step in the form $\frac{1\,\text{mol Zn}}{1\,\text{mol Ag}}$.

The conversion factor needed in the last step is the molar mass of Zn expressed in the form $\frac{65.39\,\text{g Zn}}{\text{mol Zn}}$.

Putting these together yields:

$$16.1\,\cancel{\text{g Ag}} \times \frac{1\,\cancel{\text{mol Ag}}}{107.9\,\cancel{\text{g Ag}}} \times \frac{1\,\cancel{\text{mol Zn}}}{1\,\cancel{\text{mol Ag}}} \times \frac{65.39\,\text{g Zn}}{1\,\cancel{\text{mol Zn}}} = 9.76\,\text{g Zn}$$

6. One atom of an element is found to weigh 2.107×10^{-22} g. What is the atomic weight of this element?

Answer: 126.9

Solution

What we know: g/atom

Desired answer: atomic weight of element

We know that the atomic weight of this element is numerically equal to the molar mass of this element. The solution map for calculating the molar mass is:

$$\frac{\text{g}}{\text{atom}} \rightarrow \frac{\text{g}}{\text{mol}}$$

The conversion factor needed is the Avogadro constant expressed in the form $\frac{6.022 \times 10^{23}\,\text{atoms}}{\text{mol}}$.

Applying this yields:

$$\frac{2.107 \times 10^{-22}\ \text{g}}{\cancel{\text{atom}}} \times \frac{6.022 \times 10^{23}\ \cancel{\text{atoms}}}{\text{mol}} = \frac{126.9\ \text{g}}{\text{mol}}$$

Since the molar mass of this element is 126.9 g/mol, then the atomic weight of the element is 126.9.

7. Which has the larger mass, 1.0 mmol of calcium or 1.5 mmol of sulfur? Justify your choice.

Answer: 1.5 mmol sulfur

<u>Solution</u>

The most direct way to make this determination is to calculate the mass of each and compare. The general solution map for this calculation is:

mmol element ➔ mol element ➔ g element

The conversion factor needed in the first step is that between *mmol* and *mol* in the form $\frac{10^{-3}\ \text{mol}}{1\ \text{mmol}}$.

The conversion factor needed in the second step is the molar mass of the element expressed in the form $\frac{\text{g}}{\text{mol}}$.

Applying these yields:

$$1.0\ \cancel{\text{mmol Ca}} \times \frac{10^{-3}\ \cancel{\text{mol Ca}}}{1\ \cancel{\text{mmol Ca}}} \times \frac{40.08\ \text{g Ca}}{\cancel{\text{mol Ca}}} = 0.040\ \text{g Ca}$$

$$1.5\ \cancel{\text{mmol S}} \times \frac{10^{-3}\ \cancel{\text{mol S}}}{1\ \cancel{\text{mmol S}}} \times \frac{32.07\ \text{g S}}{\cancel{\text{mol S}}} = 0.048\ \text{g S}$$ ← larger mass

8. Which has the larger number of atoms, 0.045 μg of nickel or 0.032 μg of potassium? Justify your choice.

Answer: 0.032 μg potassium

Solution

The most direct way to make this determination is to calculate the moles of each and compare. The general solution map for this calculation is:

µg element ➔ g element ➔ mol element

The conversion factor needed in the first step is that between *µg* and *g* in the form $\frac{10^{-6}\ \text{g}}{1\ \mu\text{g}}$.

The conversion factor needed in the second step is the molar mass of the element expressed in the form $\frac{\text{mol}}{\text{g}}$.

Applying these yields:

$$0.045\ \cancel{\mu\text{g Ni}} \times \frac{10^{-6}\ \cancel{\text{g Ni}}}{1\ \cancel{\mu\text{g Ni}}} \times \frac{1\ \text{mol Ni}}{\cancel{58.69\ \text{g Ni}}} = 7.7 \times 10^{-10}\ \text{mol Ni}$$

$$0.032\ \cancel{\mu\text{g K}} \times \frac{10^{-6}\ \cancel{\text{g K}}}{1\ \cancel{\mu\text{g K}}} \times \frac{1\ \text{mol K}}{\cancel{39.10\ \text{g K}}} = 8.2 \times 10^{-10}\ \text{mol K} \quad \leftarrow \text{larger mol}$$

9. Calculate the molar mass for each of the following compounds.

   a) potassium hydrogen phosphate
   b) $Pb(C_2H_3O_2)_2$

Answers: a) 174.18 g/mol

b) 325.29 g/mol

Solutions

a) What we know: potassium hydrogen phosphate; atomic weights of elements

Desired answer: g potassium hydrogen phosphate/mol

The chemical formula for potassium hydrogen phosphate is K_2HPO_4. The formula indicates that in 1 mole of potassium hydrogen phosphate there are 2 moles of potassium, 1 mole of hydrogen, 1 mole of phosphorus and 4 moles of oxygen. From the periodic table we see that 1 mole of potassium atoms weighs 39.10 g, 1 mole of hydrogen atoms weighs 1.008 g, 1 mole of phosphorus atoms weighs 30.97 g and 1 mole of oxygen atoms weighs 16.00 g. Thus, one mole of K_2HPO_4 will weigh:

2 K $\quad 2\ \cancel{\text{mol K}} \times \frac{39.10\ \text{g}}{\cancel{\text{mol K}}} = \quad 78.20\ \text{g}$

1 H $\quad 1\ \cancel{\text{mol H}} \times \frac{1.008\ \text{g}}{\cancel{\text{mol H}}} = \quad 1.008\ \text{g}$

1 P $\quad 1\ \cancel{\text{mol P}} \times \frac{30.97\ \text{g}}{\cancel{\text{mol P}}} = \quad 30.97\ \text{g}$

4 O $\quad 4\ \cancel{\text{mol O}} \times \frac{16.00\ \text{g}}{1\ \cancel{\text{mol O}}} = \quad 64.00\ \text{g}$

174.18 g

The molar mass of potassium hydrogen phosphate is 174.18 g/mol.

b) What we know: $Pb(C_2H_3O_2)_2$; atomic weights of elements

Desired answer: g $Pb(C_2H_3O_2)_2$/mol

The chemical formula indicates that in 1 mole of lead(II) acetate there are 1 mole of lead, 4 moles of carbon, 6 moles of hydrogen and 4 moles of oxygen. From the periodic table we see that 1 mole of lead atoms weighs 207.2 g, 1 mole of carbon atoms weighs 12.01 g, 1 mole of hydrogen atoms weighs 1.008 g and 1 mole of oxygen atoms weighs 16.00 g. Thus, one mole of K_2HPO_4 will weigh:

1 Pb $\quad 1\,\cancel{\text{mol Pb}} \times \dfrac{207.2\text{ g}}{1\,\cancel{\text{mol Pb}}} = 207.2\text{ g}$

4 C $\quad 4\,\cancel{\text{mol C}} \times \dfrac{12.01\text{ g}}{1\,\cancel{\text{mol C}}} = 48.04\text{ g}$

6 H $\quad 6\,\cancel{\text{mol H}} \times \dfrac{1.008\text{ g}}{1\,\cancel{\text{mol H}}} = 6.048\text{ g}$

4 O $\quad 4\,\cancel{\text{mol O}} \times \dfrac{16.00\text{ g}}{1\,\cancel{\text{mol O}}} = 64.00\text{ g}$

325.29 g

The molar mass of lead(II) acetate is 325.29 g/mol.

10. Calculate the number of moles of compound in each of the following samples.

a) 2.239 g C_2H_5OH
b) 63.1 ng sulfur trioxide
c) 1.48×10^2 kg potassium permanganate

Answers: a) 0.04860 mol

b) 7.88×10^{-10} mol

c) 9.36×10^2 mol

<u>Solutions</u>

a) What we know: g C_2H_5OH

Desired answer: mol C_2H_5OH

The solution map for this calculation is

g C_2H_5OH → mol C_2H_5OH

The conversion factor needed is the molar mass of C_2H_5OH. From the periodic table we can get the molar masses of carbon, hydrogen and oxygen and add them as follows. 2(12.01 g) C + 6(1.008 g) H + 16.00 g O = 46.07 g. Thus the molar mass of C_2H_5OH is 46.07 g/mol. This is used in the form $\frac{1\,mol\,C_2H_5OH}{46.07\,g\,C_2H_5OH}$ to convert units properly. Applying this yields:

$$2.239\,\cancel{g\,C_2H_5OH} \times \frac{1\,mol\,C_2H_5OH}{46.07\,\cancel{g\,C_2H_5OH}} = 0.04860\,mol\,C_2H_5OH$$

b) What we know: ng sulfur trioxide

Desired answer: mol sulfur trioxide

The solution map for this calculation is

ng sulfur trioxide ➔ g sulfur trioxide ➔ mol sulfur trioxide

The formula for sulfur trioxide is SO_3. The conversion factor needed in the first step is that between *ng* and *g* in the form $\frac{10^{-9}\,g}{1\,ng}$. The conversion factor needed in the second step is the molar mass of SO_3. From the periodic table we can get the molar masses of sulfur and oxygen and add them as follows. 32.07 g S + 3(16.00 g) O = 80.07 g. Thus the molar mass of SO_3 is 80.07 g/mol. This is used in the form $\frac{1\,mol\,SO_3}{80.07\,g\,SO_3}$ to convert units properly.

Putting these together yields:

$$63.1\,\cancel{ng\,SO_3} \times \frac{10^{-9}\,\cancel{g\,SO_3}}{1\,\cancel{ng\,SO_3}} \times \frac{1\,mol\,SO_3}{80.07\,\cancel{g\,SO_3}} = 7.88 \times 10^{-10}\,mol\,SO_3$$

c) What we know: kg potassium permanganate

Desired answer: mol potassium permanganate

The solution map for this calculation is

kg potassium permanganate ➔ g potassium permanganate ➔ mol potassium permanganate

The formula for potassium permanganate is $KMnO_4$. The conversion factor needed in the first step is that between *kg* and *g* in the form $\frac{10^3\,g}{1\,kg}$. The conversion factor needed in the second step is the molar mass of $KMnO_4$. From the periodic table we can get the molar masses of potassium, manganese and oxygen and add them as follows. 39.10 g K + 54.94 g Mn + 4(16.00 g) O = 158.04 g. Thus the molar mass of $KMnO_4$ is 158.04 g/mol. This is used in the form $\frac{1\,mol\,KMnO_4}{158.04\,g\,KMnO_4}$ to convert units properly.

Putting these together yields:

$$1.48 \times 10^2\ \cancel{kg\,KMnO_4} \times \frac{10^3\ \cancel{g\,KMnO_4}}{1\ \cancel{kg\,KMnO_4}} \times \frac{1\,mol\,KMnO_4}{158.04\ \cancel{g\,KMnO_4}} = 9.36 \times 10^2\ mol\,KMnO_4$$

11. How many CO molecules are present in 18.4 metric tons of carbon monoxide? One metric tons equals 1000 kg.

Answer: 3.96×10^{29} molecules

Solution

What we know: metric tons CO

Desired answer: number of molecules CO

The solution map for this calculation is

metric tons CO ➔ kg CO ➔ g CO ➔ mol CO ➔ molecules CO

The conversion factor needed in the first step is that between *metric tons* and *kg* in the form $\frac{1000\,kg}{1\,metric\,ton}$. The conversion factor needed in the second step is that between *kg* and *g* in the form $\frac{10^3\,g}{1\,kg}$. The conversion factor needed in the third step is the molar mass of CO. From the periodic table we can get the molar masses of carbon and oxygen and add them as follows. 12.01 g C + 16.00 g O = 28.01 g. Thus the molar mass of CO is 28.01 g/mol. This is used in the form $\frac{1\,mol\,CO}{28.01\,g\,CO}$ to convert units properly.

The conversion factor needed in the last step is the Avogadro constant in the form $\frac{6.022 \times 10^{23} \text{ molecules CO}}{\text{mol CO}}$.

Putting these together yields:

$$18.4 \cancel{\text{metric tons CO}} \times \frac{1000 \cancel{\text{kg CO}}}{1 \cancel{\text{metric ton CO}}} \times \frac{10^3 \cancel{\text{g CO}}}{1 \cancel{\text{kg CO}}} \times \frac{1 \cancel{\text{mol CO}}}{28.01 \cancel{\text{g CO}}} \times \frac{6.022 \times 10^{23} \text{ molecules CO}}{1 \cancel{\text{mol CO}}} = 3.96\ 10^{29} \text{ molecules CO}$$

12. Calculate the mass in grams of each of the following samples.

a) 9.44 mol copper(II) sulfate
b) 7.11 mmol Li_2CO_3

Answers: a) 1.51×10^3 g

b) 0.427 g

<u>Solutions</u>

a) What we know: mol copper(II) sulfate

Desired answer: g copper(II) sulfate

The solution map for this calculation is

mol copper(II) sulfate ➔ g copper(II) sulfate

The formula for copper(II) sulfate is $CuSO_4$. The conversion factor needed is the molar mass of $CuSO_4$. From the periodic table we can get the molar masses of copper, sulfur and oxygen and add them as follows. 63.55 g Cu + 32.07 g S + 4(16.00 g) O = 159.62 g. Thus the molar mass of $CuSO_4$ is 159.62 g/mol. This is used in the form $\frac{159.62 \text{ g } CuSO_4}{1 \text{ mol } CuSO_4}$ to convert units properly.

Applying this yields:

$$9.44 \cancel{\text{mol } CuSO_4} \times \frac{159.62 \text{ g } CuSO_4}{1 \cancel{\text{mol } CuSO_4}} = 1.51 \times 10^3 \text{ g } CuSO_4$$

b) What we know: mmol Li_2CO_3

Desired answer: g Li_2CO_3

The solution map for this calculation is

mmol Li_2CO_3 ➔ mol Li_2CO_3 ➔ g Li_2CO_3

The conversion factor needed in the first step is that between *mmol* and *mol* in the form $\frac{10^{-3}\text{ mol}}{1\text{ mmol}}$. The conversion factor needed in the second step is the molar mass of Li_2CO_3. From the periodic table we can get the molar masses of lithium, carbon and oxygen and add them as follows. 2(6.941) g Li + 12.01 g C + 3(16.00 g) O = 60.01 g. Thus the molar mass of Li_2CO_3 is 60.01 g/mol. This is used in the form $\frac{60.01\text{ g Li}_2\text{CO}_3}{1\text{ mol Li}_2\text{CO}_3}$ to convert units properly.

Putting these together yields:

$$7.11\,\cancel{\text{mmol Li}_2\text{CO}_3} \times \frac{10^{-3}\,\cancel{\text{mol Li}_2\text{CO}_3}}{1\,\cancel{\text{mmol Li}_2\text{CO}_3}} \times \frac{60.01\text{ g Li}_2\text{CO}_3}{1\,\cancel{\text{mol Li}_2\text{CO}_3}} = 0.427\text{ g Li}_2\text{CO}_3$$

13. Calculate the moles of sulfur atoms in each of the following samples.

a) 4.63 g sodium thiosulfate
b) 5.81 µg Na_2S

Answers: a) 0.0586 mol

b) 7.44×10^{-8} mol

<u>Solutions</u>

a) What we know: g sodium thiosulfate

Desired answer: mol sulfur atoms

The solution map for this calculation is:

g sodium thiosulfate ➔ mol sodium thiosulfate ➔ mol S

The formula for sodium thiosulfate is $Na_2S_2O_3$. The conversion factor needed in the first step is the molar mass of $Na_2S_2O_3$. From the periodic table we can get the molar masses of sodium, sulfur and oxygen and add them as follows. 2(22.99 g) Na + 2(32.07 g) S + 3(16.00 g) O = 158.12 g. Thus the molar mass of $Na_2S_2O_3$ is 158.12 g/mol. This is needed in the form $\frac{1\,\text{mol}\,Na_2S_2O_3}{158.12\,\text{g}\,Na_2S_2O_3}$.

The formula of $Na_2S_2O_3$ indicates that in 1 mole of $Na_2S_2O_3$ there are 2 moles of sulfur atoms. This relationship can be applied as the conversion factor in the second step in the form $\frac{2\,\text{mol}\,S}{1\,\text{mol}\,Na_2S_2O_3}$.

Putting these together yields:

$$4.63\,\cancel{\text{g}\,Na_2S_2O_3} \times \frac{1\,\cancel{\text{mol}\,Na_2S_2O_3}}{158.12\,\cancel{\text{g}\,Na_2S_2O_3}} \times \frac{2\,\text{mol}\,S}{1\,\cancel{\text{mol}\,Na_2S_2O_3}} = 0.0586\,\text{mol}\,S$$

b) What we know: µg Na_2S

Desired answer: mol S atoms

The solution map for this calculation is:

µg Na_2S ➔ g Na_2S ➔ mol Na_2S ➔ mol S

The conversion factor needed in the first step is that between *µg* and *g* in the form $\frac{10^{-6}\,\text{g}}{1\,\mu\text{g}}$. The conversion factor needed in the second step is the molar mass of Na_2S. From the periodic table we can get the molar masses of sodium and sulfur and add them as follows. 2(22.99 g) Na + 32.07 g S = 78.05 g. Thus the molar mass of Na_2S is 78.05 g/mol. This is needed in the form $\frac{1\,\text{mol}\,Na_2S}{78.05\,\text{g}\,Na_2S}$.

The formula of Na_2S indicates that in 1 mole of Na_2S there is 1 mole of sulfur atoms. This relationship can be applied as the conversion factor in the last step in the form $\frac{1\,\text{mol}\,S}{1\,\text{mol}\,Na_2S}$.

Putting these together yields:

$$5.81\,\cancel{\mu\text{g}\,Na_2S} \times \frac{10^{-6}\,\cancel{\text{g}\,Na_2S}}{1\,\cancel{\mu\text{g}\,Na_2S}} \times \frac{1\,\cancel{\text{mol}\,Na_2S}}{78.05\,\cancel{\text{g}\,Na_2S}} \times \frac{1\,\text{mol}\,S}{1\,\cancel{\text{mol}\,Na_2S}} = 7.44 \times 10^{-8}\,\text{mol}\,S$$

14. Calculate the number of carbon atoms in a 3.92-g sample of $C_6H_4Cl_2$.

Answer: 9.64×10^{22} atoms

Solution

What we know: g $C_6H_4Cl_2$

Desired answer: number of C atoms

The solution map for this calculation is:

$$\text{g } C_6H_4Cl_2 \rightarrow \text{mol } C_6H_4Cl_2 \rightarrow \text{mol C} \rightarrow \text{atoms C}$$

The conversion factor needed in the first step is the molar mass of $C_6H_4Cl_2$. From the periodic table we can get the molar masses of carbon, hydrogen and chlorine and add them as follows. 6(12.01 g) C + 4(1.008 g) H + 2(35.45 g) Cl = 146.99 g. Thus the molar mass of $C_6H_4Cl_2$ is 146.99 g/mol. This is needed in the form $\frac{1\,\text{mol}\,C_6H_4Cl_2}{146.99\,\text{g}\,C_6H_4Cl_2}$.

The formula of $C_6H_4Cl_2$ indicates that in 1 mole of $C_6H_4Cl_2$ there are 6 moles of carbon atoms. This relationship can be applied as the conversion factor in the second step in the form $\frac{6\,\text{mol}\,C}{1\,\text{mol}\,C_6H_4Cl_2}$.

The conversion factor needed in the last step is the Avogadro constant in the form $\frac{6.022 \times 10^{23}\,\text{atoms}\,C}{1\,\text{mol}\,C}$.

Putting these together yields:

$$3.92\,\cancel{\text{g}\,C_6H_4Cl_2} \times \frac{1\,\text{mol}\,\cancel{C_6H_4Cl_2}}{146.99\,\cancel{\text{g}\,C_6H_4Cl_2}} \times \frac{6\,\cancel{\text{mol}\,C}}{1\,\cancel{\text{mol}\,C_6H_4Cl_2}} \times \frac{6.022 \times 10^{23}\,\text{atoms}\,C}{1\,\cancel{\text{mol}\,C}} = 9.64 \times 10^{22}\,\text{atoms}\,C$$

15. How many moles of oxygen atoms are present in 4.40 mmol calcium phosphate?

Answer: 0.0352 mol

Solution

What we know: mmol calcium phosphate

Desired answer: mol O

The solution map for this calculation is:

mmol calcium phosphate ➔ mol calcium phosphate ➔ mol O

The formula for calcium phosphate is $Ca_3(PO_4)_2$. The conversion factor needed in the first step is that between *mmol* and *mol* in the form $\frac{10^{-3}\text{ mol}}{1\text{ mmol}}$.

The formula of $Ca_3(PO_4)_2$ indicates that in 1 mole of $Ca_3(PO_4)_2$ there are 8 moles of oxygen atoms. This relationship can be applied as the conversion factor in the second step in the form $\frac{8\text{ mol O}}{1\text{ mol }Ca_3(PO_4)_2}$.

Putting these together yields:

$$4.40\ \cancel{\text{mmol }Ca_3(PO_4)_2} \times \frac{10^{-3}\ \cancel{\text{mol }Ca_3(PO_4)_2}}{1\ \cancel{\text{mmol }Ca_3(PO_4)_2}} \times \frac{8\text{ mol O}}{1\ \cancel{\text{mol }Ca_3(PO_4)_2}} = 0.0352\text{ mol O}$$

Mass Percent

1. Calculate the mass percent of each element in the following compounds.

   a) barium chloride
   b) sodium sulfate

Answers: a) 65.95 % Ba and 34.05 % Cl

b) 32.37 % Na, 22.58 % S and 45.05 % O

Solutions

a) What we know: barium chloride; atomic weights of elements

Desired answer: mass % of each element present

The formula for barium chloride is $BaCl_2$. Since the mass percent values apply to any sample of $BaCl_2$, it is convenient to consider one mole of this compound. From the periodic table we can get the molar masses of barium and chlorine and add them as follows. 137.3 g Ba + 2(35.45 g) Cl = 208.2 g. Thus the molar mass of $BaCl_2$ is 208.2 g/mol, so 208.2 g represents the total mass of material. The formula indicates that 1 mole of $BaCl_2$ contains 1 mole of barium. Therefore the mass of barium (=component mass) present is 137.3 g.

These values are used to obtain the mass percent for barium.

$$\text{mass \% Ba} = \frac{137.3\,\text{g Ba}}{208.2\,\text{g BaCl}_2} \text{ x } 100 = 65.95\,\%$$

Since chlorine is the only other element present in the compound, the mass % Cl = 100.00 % - 65.95 % = 34.05 %.

b) What we know: sodium sulfate; atomic weights of elements

Desired answer: mass % of each element present

The formula for sodium sulfate is Na_2SO_4. Since the mass percent values apply to any sample of Na_2SO_4, it is convenient to consider one mole of this compound. From the periodic table we can get the molar masses of sodium, sulfur and oxygen and add them as follows. 2(22.99 g) Na + 32.07 g S + 4(16.00 g) O = 142.05 g. Thus the molar mass of Na_2SO_4 is 142.05 g/mol, so 142.05 g represents the total mass of material. The formula indicates that 1 mole of Na_2SO_4 contains 2 moles of sodium. Therefore the mass of sodium (=component mass) present is 2(22.99 g) = 45.98 g.

These values are used to obtain the mass percent for sodium.

$$\text{mass \% Na} = \frac{45.98\,\text{g Na}}{142.05\,\text{g Na}_2\text{SO}_4} \text{ x } 100 = 32.37\,\%$$

The formula indicates that 1 mole of Na_2SO_4 contains 1 mole of sulfur. Therefore the mass of sulfur (=component mass) present is 32.07 g.

These values are used to obtain the mass percent for sulfur.

$$\text{mass \% S} = \frac{32.07\,\text{g S}}{142.05\,\text{g Na}_2\text{SO}_4} \text{ x } 100 = 22.58\,\%$$

Since oxygen is the only other element present in the compound, the mass % O = 100.00 % - 32.37 % - 22.58% = 45.05 %.

2. Which of the following compounds contains the largest mass percent of nitrogen? Justify your choice.

a) NH_4NO_3 b) HNO_3 c) N_2O_4 d) $Al(NO_3)_3$

Answer: NH_4NO_3

Solution

The most direct way to answer this is to determine the mass percent N in each compound. This is calculated from the mass of one mole of a compound and the mass of nitrogen present in one mole of that compound. The required data are shown below.

Compound	Mass of one mole of compound	Mass of N in one mole of compound	Mass % N
NH_4NO_3	80.05 g	28.02 g	35.00 %
HNO_3	63.02 g	14.01 g	22.23 %
N_2O_4	92.02	28.02 g	30.45 %
$Al(NO_3)_3$	213.01 g	42.03 g	19.72 %

Thus, NH_4NO_3 has the largest mass % N.

3. In a particular molecular compound the mass percent sulfur is 50.0% and the mass percent oxygen is 50.0%. What is the ratio of oxygen atoms to sulfur atoms in a molecule of this compound?

Answer: 2 O atoms/S atom

Solution

What we know: mass % of each element present; atomic weights of elements

Desired answer: number of O atoms/S atom in compound

The ratio of O atoms to S atoms is the same as the ratio of mol O to mol S. In order to obtain the moles of each element, choose an arbitrary mass of the compound (100 g simplifies the math) and use the following solution map:

g compound ➔ g element ➔ mol element

The conversion factor for the first step is the mass percent of the element used in the form $\frac{\text{g element}}{100\,\text{g compound}}$. The conversion factor for the second step is the molar mass of the element in the form $\frac{\text{mol element}}{\text{g element}}$.

Applying these to each element yields:

$$100\,\cancel{\text{g compound}} \times \frac{50.0\,\cancel{\text{g O}}}{100\,\cancel{\text{g compound}}} \times \frac{1\,\text{mol O}}{16.00\,\cancel{\text{g O}}} = 3.12\,\text{mol O}$$

$$100\,\cancel{\text{g compound}} \times \frac{50.0\,\cancel{\text{g S}}}{100\,\cancel{\text{g compound}}} \times \frac{1\,\text{mol S}}{32.07\,\cancel{\text{g S}}} = 1.56\,\text{mol S}$$

The mole ratio of O to S is 3.12 mol O/1.56 mol S = 2 which is the same as the atom ratio.

4. If a type of stainless steel contains 18% chromium by mass, how many moles of chromium are present in a bar of this material weighing 1.5 kg?

Answer: 5.2 mol

Solution

What we know: mass of sample; mass % Cr

Desired answer: mol Cr

The solution map for this calculation is:

kg sample ➔ g sample ➔ g Cr ➔ mol Cr

The conversion factor needed in the first step is that between *kg* and *g* in the form $\frac{10^3\,\text{g}}{1\,\text{kg}}$.

The conversion factor for the second step is the mass percent of Cr used in the form $\frac{18\,\text{g Cr}}{100\,\text{g sample}}$.

The conversion factor for the last step is the molar mass of Cr in the form $\frac{1\,\text{mol Cr}}{52.00\,\text{g Cr}}$.

Putting these together yields:

$$1.5\,\cancel{\text{kg sample}} \times \frac{10^3\,\cancel{\text{g sample}}}{1\,\cancel{\text{kg sample}}} \times \frac{18\,\cancel{\text{g Cr}}}{100\,\cancel{\text{g sample}}} \times \frac{1\,\text{mol Cr}}{52.00\,\cancel{\text{g Cr}}} = 5.2\,\text{mol Cr}$$

Stoichiometric Calculations (mole-to-mole)

1. Balance the following equation and state the meaning of the equation in terms of individual units of reactants and products and in terms of moles of reactants and products.

$$Al(s) + MnO_2(s) \rightarrow Mn(s) + Al_2O_3(s)$$

Answers: $4Al(s) + 3MnO_2(s) \rightarrow 3Mn(s) + 2Al_2O_3(s)$

4 atoms 3 formula units 3 atoms 2 formula units

4 mol 3 mol 3 mol 2 mol

Solution

To balance this equation, place a coefficient of 3 in front of MnO_2 and a coefficient of 2 in front of Al_2O_3 to balance oxygen atoms. This requires a coefficient of 3 in front of Mn to balance manganese atoms. Finally, a coefficient of 4 in front of Al will balance aluminum atoms. The balanced equation is:

$$4Al(s) + 3MnO_2(s) \rightarrow 3Mn(s) + 2Al_2O_3(s)$$

At the most fundamental level this equation suggests that 4 atoms of aluminum react with 3 formula units of MnO_2 to produce 3 atoms of Mn and 2 formula units of Al_2O_3. Multiplying each reactant amount and each product amount by the Avogadro constant yields 4 mol Al reacting with 3 mol MnO_2 to form 3 mol Mn and 2 mol Al_2O_3.

2. How many moles of CO_2 are needed to react completely with 0.675 mol LiOH?

$$LiOH(aq) + CO_2(g) \rightarrow Li_2CO_3(aq) + H_2O(l) \quad \text{(unbalanced)}$$

Answer: 0.338 mol

Solution

What we know: mol LiOH; equation relating CO_2 and LiOH

Desired answer: mol CO_2

The solution map for this calculation is:

mol LiOH ➔ mol CO_2

The conversion factor needed is the mole ratio for these two substances from the balanced equation. To balance this equation place a coefficient of 2 in front of LiOH. The balanced equation is:

$$2LiOH(aq) + CO_2(g) \rightarrow Li_2CO_3(aq) + H_2O(l)$$

Applying the mole ratio yields:

$$0.675\,\cancel{\text{mol LiOH}}\ \frac{1\,\text{mol}\,CO_2}{2\,\cancel{\text{mol LiOH}}} = 0.338\,\text{mol}\,CO_2$$

3. Given the reaction

$$4FeS(s) + 7O_2(g) \rightarrow 2Fe_2O_3(s) + 4SO_2(g)$$

how many moles of O_2 are needed to:

a) produce 0.693 mol Fe_2O_3?

b) react completely with 9.14 mol FeS?

c) form 1.51 mol SO_2?

Answers: a) 2.43 mol

b) 16.0 mol

c) 2.64 mol

Solutions

a) What we know: mol Fe_2O_3; balanced equation relating O_2 and Fe_2O_3

Desired answer: mol O_2

The solution map for this problem is:

mol Fe_2O_3 → mol O_2

The conversion factor needed is the mole ratio for these two substances from the balanced equation in the form $\frac{7\,\text{mol}\,O_2}{2\,\text{mol}\,Fe_2O_3}$.

Applying this yields:

$$0.693\,\cancel{\text{mol}\,Fe_2O_3} \times \frac{7\,\text{mol}\,O_2}{2\,\cancel{\text{mol}\,Fe_2O_3}} = 2.43\,\text{mol}\,O_2$$

b) What we know: mol FeS; balanced equation relating O_2 and FeS

Desired answer: mol O_2

The solution map for this problem is:

mol FeS → mol O_2

The conversion factor needed is the mole ratio for these two substances from the balanced equation in the form $\frac{7\,\text{mol}\,O_2}{4\,\text{mol}\,FeS}$.

Applying this yields:

$$9.14\,\cancel{\text{mol}\,FeS} \times \frac{7\,\text{mol}\,O_2}{4\,\cancel{\text{mol}\,FeS}} = 16.0\,\text{mol}\,O_2$$

c) What we know: mol SO_2; balanced equation relating O_2 and SO_2

Desired answer: mol O_2

The solution map for this problem is:

mol SO_2 → mol O_2

The conversion factor needed is the mole ratio for these two substances from the balanced equation in the form $\frac{7\,\text{mol}\,O_2}{4\,\text{mol}\,SO_2}$.

Applying this yields:

$$1.51\,\cancel{\text{mol}\,SO_2} \times \frac{7\,\text{mol}\,O_2}{4\,\cancel{\text{mol}\,SO_2}} = 2.64\,\text{mol}\,O_2$$

Stoichiometric Calculations (mole-to-mass & mass-to-mole)

1. How many moles of each product can be formed from the decomposition of 1.00 g of the rocket fuel hydrazine (N_2H_4)?

$$3N_2H_4(l) \rightarrow 4NH_3(g) + N_2(g)$$

Answers: 0.0416 mol NH_3 and 0.0104 mol N_2

Solution

What we know: g N_2H_4; balanced equation relating NH_3, N_2 and N_2H_4

Desired answer: mol NH_3 and mol N_2

The solution maps for these calculations are:

g N_2H_4 ➔ mol N_2H_4 ➔ mol NH_3

g N_2H_4 ➔ mol N_2H_4 ➔ mol N_2

In each case, the conversion factor needed in the first step is the molar mass of N_2H_4. From the periodic table we can get the molar masses of nitrogen and hydrogen and add them as follows. 2(14.01 g) N + 4(1.008 g) H = 32.05 g. Thus, the molar mass of N_2H_4 is 32.05 g/mol. It is used in the form $\frac{1\,\text{mol}\,N_2H_4}{32.05\,\text{g}\,N_2H_4}$ in order to cancel units properly.

The conversion factor needed in the second step is the mole ratio for the particular product and N_2H_4: $\frac{4\,\text{mol}\,NH_3}{3\,\text{mol}\,N_2H_4}$ and $\frac{1\,\text{mol}\,N_2}{3\,\text{mol}\,N_2H_4}$

Putting these together yields:

$$1.00\,\cancel{g\,N_2H_4} \times \frac{1\,\cancel{mol\,N_2H_4}}{32.05\,\cancel{g\,N_2H_4}} \times \frac{4\,mol\,NH_3}{3\,\cancel{mol\,N_2H_4}} = 0.0416\,mol\,NH_3$$

$$1.00\,\cancel{g\,N_2H_4} \times \frac{1\,\cancel{mol\,N_2H_4}}{32.05\,\cancel{g\,N_2H_4}} \times \frac{1\,mol\,N_2}{3\,\cancel{mol\,N_2H_4}} = 0.0104\,mol\,N_2$$

Note that the number of moles of NH_3 is four(4) times that of N_2 as suggested by the balanced equation.

2. How many moles of oxygen gas are needed for the complete combustion of 19.6 g acetylene, $C_2H_2(g)$?

Answer: 1.88 mol

<u>Solution</u>

What we know: g C_2H_2

Desired answer: mol O_2

First, a balanced equation is needed for the reaction between C_2H_2 and O_2. Since this is a combustion reaction, the products are $CO_2(g)$ and $H_2O(l)$.

$$C_2H_2(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

Notice that the hydrogen atoms are balanced so first, balance the carbon atoms. Since there are 2 C atoms on the reactant side and only 1 C atom on the product side, a coefficient of 2 in front of CO_2 will balance carbon atoms.

$$C_2H_2(g) + O_2(g) \rightarrow 2CO_2(g) + H_2O(l) \quad \text{(H and C atoms balanced)}$$

Finally, balance the oxygen atoms. Since there are 2 O atoms on the reactant side and 5 O atoms on the product side, multiplying O_2 by 2½, or 5/2, will balance oxygen atoms.

$$C_2H_2(g) + 5/2O_2(g) \rightarrow 2CO_2(g) + H_2O(l) \quad \text{(all atoms balanced)}$$

The fractional coefficient in front of O_2 can be converted to the smallest whole number by multiplying by 2. This requires that all other coefficients be multiplied by 2 in order to retain the atom balance. The resulting balanced equation is:

$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$ (all atoms balanced using whole-number coefficients)

The solution map for this problem is:

g C_2H_2 → mol C_2H_2 → mol O_2

The conversion factor needed in the first step is the molar mass of C_2H_2. From the periodic table we can get the molar masses of carbon and hydrogen and add them as follows. 2(12.01 g) C + 2(1.008 g) H = 26.04 g. Thus, the molar mass of C_2H_2 is 26.04 g/mol. It is used in the form $\frac{1\,\text{mol}\,C_2H_2}{26.04\,\text{g}\,C_2H_2}$ in order to cancel units properly.

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form $\frac{5\,\text{mol}\,O_2}{2\,\text{mol}\,C_2H_2}$.

Putting these together yields:

$$19.6\,\cancel{\text{g}\,C_2H_2} \times \frac{1\,\cancel{\text{mol}\,C_2H_2}}{26.04\,\cancel{\text{g}\,C_2H_2}} \times \frac{5\,\text{mol}\,O_2}{2\,\cancel{\text{mol}\,C_2H_2}} = 1.88\,\text{mol}\,O_2$$

3. How many kilograms of Li_2O are needed to react completely 4.17×10^3 mol H_2O?

$$Li_2O(s) + H_2O(g) \rightarrow 2LiOH(s)$$

Answer: 125 kg

Solution

What we know: mol H_2O; balanced equation relating Li_2O and H_2O

Desired answer: kg Li_2O

The solution map for this problem is:

mol H_2O → mol Li_2O → g Li_2O → kg Li_2O

The conversion factor needed in the first step is the mole ratio for these two substances from the balanced equation in the form $\frac{1\,\text{mol}\,Li_2O}{1\,\text{mol}\,H_2O}$.

The conversion factor needed in the second step is the molar mass of Li_2O. From the periodic table we can get the molar masses of lithium and oxygen and add them as follows. 2(6.941 g) Li + 16.00 g O = 29.88 g. Thus, the molar mass of Li_2O is 29.88 g/mol. It is used in the form $\frac{29.88\,\text{g}\,Li_2O}{1\,\text{mol}\,Li_2O}$ in order to cancel units properly.

The conversion factor needed in the last step is that between *g* and *kg* in the form $\frac{1\,\text{kg}}{10^3\,\text{g}}$.

Putting these together yields:

$$4.17 \times 10^3\ \cancel{\text{mol}\,H_2O} \times \frac{1\ \cancel{\text{mol}\,Li_2O}}{1\ \cancel{\text{mol}\,H_2O}} \times \frac{29.88\ \cancel{\text{g}\,Li_2O}}{1\ \cancel{\text{mol}\,Li_2O}} \times \frac{1\,\text{kg}\,Li_2O}{10^3\ \cancel{\text{g}\,Li_2O}} = 125\,\text{kg}\,Li_2O$$

4. Carbon dioxide is produced in the reaction

$$H_3PO_4(aq) + MgCO_3(s) \rightarrow Mg_3(PO_4)_2(s) + CO_2(g) + H_2O(l) \quad \text{(unbalanced)}$$

How many grams of $MgCO_3$ are needed to produce 14.8 moles of CO_2?

Answer: 1.25×10^3 g

Solution

What we know: mol CO_2; equation relating $MgCO_3$ and CO_2

Desired answer: g $MgCO_3$

First, a balanced equation is needed. To balance this equation, place a coefficient of 2 in front of $MgCO_3$ to balance magnesium atoms. Then place a coefficient of 2 in front of H_3PO_4 to balance PO_4 groups. Next, a coefficient of 3 in front of CO_2 will balance carbon atoms. Finally, a coefficient of 3 in front of H_2O will balance hydrogen and oxygen atoms. The balanced equation is:

$$2H_3PO_4(aq) + 3MgCO_3(s) \rightarrow Mg_3(PO_4)_2(s) + 3CO_2(g) + 3H_2O(l)$$

The solution map for this problem is:

mol CO_2 ➔ mol $MgCO_3$ ➔ g $MgCO_3$

The conversion factor needed in the first step is the mole ratio for these two substances from the balanced equation in the form $\frac{3\,\text{mol MgCO}_3}{3\,\text{mol CO}_2}$.

The conversion factor needed in the second step is the molar mass of $MgCO_3$. From the periodic table we can get the molar masses of magnesium, carbon and oxygen and add them as follows. 24.31 g Mg + 12.01 g C + 3(16.00 g) O = 84.32 g. Thus, the molar mass of $MgCO_3$ is 84.32 g/mol. It is used in the form $\frac{84.32\,\text{g MgCO}_3}{1\,\text{mol MgCO}_3}$ in order to cancel units properly.

Putting these together yields:

$$14.8\,\cancel{\text{mol CO}_2} \times \frac{3\,\cancel{\text{mol MgCO}_3}}{3\,\cancel{\text{mol CO}_2}} \times \frac{84.32\,\text{g MgCO}_3}{1\,\cancel{\text{mol MgCO}_3}} = 1.25 \times 10^3\,\text{g MgCO}_3$$

Stoichiometric Calculations (mass-to-mass)

1. How many grams of sulfur can react with 1.79 g of copper according to the following equation?

$$\text{Cu(s)} + \text{S(s)} \rightarrow \text{CuS(s)}$$

Answer: 0.903 g

Solution

What we know: g Cu; balanced equation relating S and Cu

Desired answer: g S

The solution map for this problem is:

g Cu → mol Cu → mol S → g S

The conversion factor needed in the first step is the molar mass of Cu. From the periodic table we see that the molar mass of copper is 63.55 g/mol. It is used in the form $\frac{1\,\text{mol Cu}}{63.55\,\text{g Cu}}$ in order to cancel units properly.

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form $\frac{1\,\text{mol S}}{1\,\text{mol Cu}}$.

The conversion factor needed in the last step is the molar mass of S. From the periodic table we see that the molar mass of sulfur is 32.07 g/mol. It is used in the form $\frac{32.07\,\text{g S}}{1\,\text{mol S}}$ in order to cancel units properly.

Putting these together yields:

$$1.79\,\cancel{\text{g Cu}} \times \frac{1\,\cancel{\text{mol Cu}}}{63.55\,\cancel{\text{g Cu}}} \times \frac{1\,\cancel{\text{mol S}}}{1\,\cancel{\text{mol Cu}}} \times \frac{32.07\,\text{g S}}{1\,\cancel{\text{mol S}}} = 0.903\,\text{g S}$$

2. How many grams of chlorine gas are required to react completely with 0.455 g iron to form iron(III) chloride?

Answer: 0.867 g

<u>Solution</u>

What we know: g Fe

Desired answer: g Cl_2

First, a balanced equation is needed for the reaction between Fe and Cl_2. The unbalanced equation for this combination reaction is:

$$Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$$

To balance this equation, place a coefficient of 2 in front of $FeCl_3$ and a coefficient of 3 in front of Cl_2 to balance chlorine atoms. This requires a coefficient of 2 in front of Fe to balance iron atoms. The balanced equation is:

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

The solution map for this problem is:

$$\text{g Fe} \rightarrow \text{mol Fe} \rightarrow \text{mol } Cl_2 \rightarrow \text{g } Cl_2$$

The conversion factor needed in the first step is the molar mass of Fe. From the periodic table we see that the molar mass of iron is 55.84 g/mol. It is used in the form $\frac{1\,\text{mol Fe}}{55.84\,\text{g Fe}}$ in order to cancel units properly.

$$\frac{70.90\,\text{g}\,Cl_2}{1\,\text{mol}\,Cl_2}$$

The conversion factor needed in the second step is the mole ratio for these two substances from the balanced equation in the form $\frac{3\,\text{mol}\,Cl_2}{2\,\text{mol}\,Fe}$.

The conversion factor needed in the last step is the molar mass of Cl_2. From the periodic table we see that the molar mass of chlorine is 35.45 g/mol so the molar mass of Cl_2 is 2(35.45 g/mol) = 70.90 g/mol. It is used in the form in order to cancel units properly.
Putting these together yields:

$$0.455\,\cancel{\text{g Fe}} \times \frac{1\,\cancel{\text{mol Fe}}}{55.84\,\cancel{\text{g Fe}}} \times \frac{3\,\cancel{\text{mol}\,Cl_2}}{2\,\cancel{\text{mol Fe}}} \times \frac{70.90\,\text{g}\,Cl_2}{1\,\cancel{\text{mol}\,Cl_2}} = 0.867\,\text{g}\,Cl_2$$

3. How many grams of each product can be formed from the decomposition of 14.0 g of sodium chlorate?

$$2NaClO_3(s) \rightarrow 2NaCl(s) + 3O_2(g)$$

Answers: 7.69 g NaCl and 6.31 g O_2

Solution

What we know: g $NaClO_3$; balanced equation relating NaCl, O_2 and $NaClO_3$

Desired answer: g NaCl and g O_2

The solution maps for these calculations are:

g $NaClO_3$ ➔ mol $NaClO_3$ ➔ mol NaCl ➔ g NaCl

g $NaClO_3$ ➔ mol $NaClO_3$ ➔ mol O_2 ➔ g O_2

In each case, the conversion factor needed in the first step is the molar mass of $NaClO_3$. From the periodic table we can get the molar masses of sodium, chlorine and oxygen and add them as follows. 22.99 g Na + 35.45 g Cl + 3(16.00 g) O = 106.44 g. Thus, the molar mass of $NaClO_3$ is 106.44 g/mol. It is used in the form $\frac{1\,\text{mol}\,NaClO_3}{106.44\,\text{g}\,NaClO_3}$ in order to cancel units properly.

The conversion factor needed in the second step is the mole ratio for the particular product and N_2H_4: $\frac{2\,\text{mol NaCl}}{2\,\text{mol}\,NaClO_3}$ and $\frac{3\,\text{mol}\,O_2}{2\,\text{mol}\,NaClO_3}$

The conversion factor needed in the last step is the molar mass of the product: $\frac{58.44\,g\,NaCl}{1\,mol\,NaCl}$

and $\frac{32.00\,g\,O_2}{1\,mol\,O_2}$

Putting these together yields:

$$14.0\,\cancel{g\,NaClO_3} \times \frac{1\,\cancel{mol\,NaClO_3}}{106.44\,\cancel{g\,NaClO_3}} \times \frac{2\,\cancel{mol\,NaCl}}{2\,\cancel{mol\,NaClO_3}} \times \frac{58.44\,g\,NaCl}{1\,\cancel{mol\,NaCl}} = 7.69\,g\,NaCl$$

$$14.0\,\cancel{g\,NaClO_3} \times \frac{1\,\cancel{mol\,NaClO_3}}{106.44\,\cancel{g\,NaClO_3}} \times \frac{3\,\cancel{mol\,O_2}}{2\,\cancel{mol\,NaClO_3}} \times \frac{32.00\,g\,O_2}{1\,\cancel{mol\,O_2}} = 6.31\,g\,O_2$$

Note that the total mass of the products equals the mass of the reactant.

4. How many grams of potassium are needed to produce 16.5 kg K_2O?

$$KNO_3(s) + K(s) \rightarrow K_2O(s) + N_2(g) \quad \text{(unbalanced)}$$

Answer: 0.867 g

Solution

What we know: kg K_2O; equation relating K and K_2O

Desired answer: g K

First, a balanced equation is needed. To balance this equation, place a coefficient of 2 in front of KNO_3 to balance nitrogen atoms. Then place a coefficient of 6 in front of K_2O to balance oxygen atoms. Finally, a coefficient of 10 in front of K will balance potassium atoms. The balanced equation is:

$$2KNO_3(s) + 10K(s) \rightarrow 6K_2O(s) + N_2(g)$$

The solution map for this problem is:

$$\text{kg } K_2O \rightarrow \text{g } K_2O \rightarrow \text{mol } K_2O \rightarrow \text{mol K} \rightarrow \text{g K}$$

The conversion factor needed in the first step is that between *kg* and *g* in the form $\frac{10^3\,g}{1\,kg}$.

The conversion factor needed in the second step is the molar mass of K_2O in the form $\frac{1\,\text{mol}\,K_2O}{94.20\,\text{g}\,K_2O}$.

The conversion factor needed in the third step is the mole ratio for these two substances from the balanced equation in the form $\frac{10\,\text{mol}\,K}{6\,\text{mol}\,K_2O}$.

The conversion factor needed in the last step is the molar mass of K. From the periodic table we see that the molar mass of potassium is 39.10 g/mol. It is used in the form $\frac{39.10\,\text{g}\,K}{1\,\text{mol}\,K}$ in order to cancel units properly.

Putting these together yields:

$$16.5\,\cancel{\text{kg}\,K_2O} \times \frac{10^3\,\cancel{\text{g}\,K_2O}}{1\,\cancel{\text{kg}\,K_2O}} \times \frac{1\,\cancel{\text{mol}\,K_2O}}{94.20\,\cancel{\text{g}\,K_2O}} \times \frac{10\,\cancel{\text{mol}\,K}}{6\,\cancel{\text{mol}\,K_2O}} \times \frac{39.10\,\text{g}\,K}{1\,\cancel{\text{mol}\,K}} = 1.14 \times 10^4\,\text{g}\,K$$

Theoretical Yield and Limiting Reactant

1. Which is the limiting reactant when 0.68 g magnesium reacts with 17 mmol nitrogen gas to form Mg_3N_2?

Answer: Mg

Solution

What we know: g Mg; mmol N_2

Desired answer: which reactant, Mg or N_2, is the limiting reactant

The balanced equation for this combination reaction is:

$$3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$$

First, determine the limiting reactant by calculating how many moles of Mg_3N_2 can form from each starting amount of reactant. The solution maps for these calculations are:

g Mg → mol Mg → mol Mg_3N_2

mmol N_2 → mol N_2 → mol Mg_3N_2

For the first calculation the conversion factor needed in the first step is the molar mass of Mg in the form $\frac{1\,\text{mol Mg}}{24.31\,\text{g Mg}}$.

The conversion factor needed for the second step is the Mg_3N_2/Mg mole ratio from the balanced equation.

Applying these yields:

$$0.68\,\cancel{\text{g Mg}}\ \text{x}\ \frac{1\,\cancel{\text{mol Mg}}}{24.31\,\cancel{\text{g Mg}}}\ \text{x}\ \frac{1\,\text{mol Mg}_3\text{N}_2}{3\,\cancel{\text{mol Mg}}} = 0.0093\,\text{mol Mg}_3\text{N}_2$$

For the second calculation the conversion factor needed in the first step is that between *mmol* and *mol* in the form $\frac{10^{-3}\text{ mol}}{1\,\text{mmol}}$.

The conversion factor needed in the second step is the Mg_3N_2/N_2 mole ratio from the balanced equation.

Applying this yields:

$$17\,\cancel{\text{mmol N}_2}\ \text{x}\ \frac{10^{-3}\ \cancel{\text{mol N}_2}}{1\,\cancel{\text{mmol N}_2}}\ \text{x}\ \frac{1\,\text{mol Mg}_3\text{N}_2}{1\,\cancel{\text{mol N}_2}} = 0.017\,\text{mol Mg}_3\text{N}_2$$

Since the starting amount of Mg produces the smaller amount of Mg_3N_2, magnesium is the limiting reactant.

2. How many moles of AsF_5 can be produced when 14 moles of arsenic react with 29 mol fluorine gas?

Answer: 12 mol

<u>Solution</u>

What we know: mol As; mol F_2

Desired answer: mol AsF_5

The balanced equation for this combination reaction is:

$$2As(s) + 5F_2(g) \rightarrow 2AsF_5(g)$$

First, determine the limiting reactant by calculating how many moles of AsF_5 can form from each starting amount of reactant. The solution maps for these calculations are:

mol As $\rightarrow$ mol AsF_5

mol F_2 $\rightarrow$ mol AsF_5

For the first calculation the conversion factor needed is the AsF_5/As mole ratio from the balanced equation.

Applying this yields:

$$14\ \cancel{\text{mol As}} \times \frac{2\ \text{mol AsF}_5}{2\ \cancel{\text{mol As}}} = 14\ \text{mol AsF}_5$$

For the second calculation the conversion factor needed is the AsF_5/F_2 mole ratio from the balanced equation.

Applying this yields:

$$29\ \cancel{\text{mol F}_2} \times \frac{2\ \text{mol AsF}_5}{5\ \cancel{\text{mol F}_2}} = 12\ \text{mol AsF}_5$$

Since the starting amount of F_2 produces the smaller amount of AsF_5, fluorine is the limiting reactant and the maximum moles of AsF_5 is 12 moles.

3. When 26.5 g CO and 3.7 g H_2 are allowed to react as shown below,

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

a) which is the limiting reactant?

b) what is the theoretical yield in grams of CH_3OH?

c) how much of the reactant in excess remains?

Answers: a) H_2

b) 29.32 g

c) 0.8 g

Solutions

a) What we know: g CO; g H_2; balanced equation relating CO, H_2 and CH_3OH

Desired answer: which reactant, CO or H_2, is the limiting reactant

Determine the limiting reactant by calculating how many moles of CH_3OH can form from each starting amount of reactant. The solution maps for these calculations are:

g CO ➔ mol CO ➔ mol CH_3OH

g H_2 ➔ mol H_2 ➔ mol CH_3OH

For the first calculation the conversion factors needed are the molar mass of CO and the CH_3OH/CO mole ratio.

Putting these together yields:

$$26.5\,\cancel{\text{g CO}} \times \frac{1\,\cancel{\text{mol CO}}}{28.01\,\cancel{\text{g CO}}} \times \frac{1\,\text{mol CH}_3\text{OH}}{1\,\cancel{\text{mol CO}}} = 0.946\,\text{mol CH}_3\text{OH}$$

For the second calculation the conversion factors needed are the molar mass of H_2 and the CH_3OH/H_2 mole ratio.

Putting these together yields:

$$3.7\,\cancel{\text{g H}_2} \times \frac{1\,\cancel{\text{mol H}_2}}{2.016\,\cancel{\text{g H}_2}} \times \frac{1\,\text{mol CH}_3\text{OH}}{2\,\cancel{\text{mol H}_2}} = 0.918\,\text{mol CH}_3\text{OH}$$

Since the starting amount of H_2 produces the smaller amount of CH_3OH, H_2 is the limiting reactant.

b) What we know: mol CH_3OH produced from limiting reactant

Desired answer: theoretical yield of CH_3OH in grams

Since the theoretical yield is the amount of product produced from the limiting reactant, the solution map for this calculation is:

mol CH_3OH ➔ g CH_3OH

The conversion factor needed for this calculation is the molar mass of CH_3OH.

Applying this yields:

$$0.918\,\cancel{\text{mol}\,CH_3OH} \times \frac{32.04\,\text{g}\,CH_3OH}{1\,\cancel{\text{mol}\,CH_3OH}} = 29.32\,\text{g}\,CH_3OH$$

c) What we know: mol CH_3OH produced from limiting reactant

Desired answer: g CO remaining

Use the theoretical yield of CH_3OH to calculate how much CO was used and subtract this from the starting quantity. The solution maps for these calculations are:

mol CH_3OH ➔ mol CO ➔ g CO used

g CO remaining = g CO initially - g CO used

For the first calculation the conversion factors needed are the CO/CH_3OH mole ratio and the molar mass of CO.

Putting these together yields:

$$0.918\,\cancel{\text{mol}\,CH_3OH} \times \frac{1\,\cancel{\text{mol}\,CO}}{1\,\cancel{\text{mol}\,CH_3OH}} \times \frac{28.01\,\text{g}\,CO}{1\,\cancel{\text{mol}\,CO}} = 25.7\,\text{g}\,CO\,\text{used}$$

Therefore, the mass of CO that remains equals 26.5 g - 25.7 g = 0.8 g.

4. How many kilograms of the excess reactant remain when a mixture of 2.50 kg of SiO_2 and 2.50 kg of carbon react?

$$SiO_2(s) + 3C(s) \rightarrow SiC(s) + 2CO(g)$$

Answer: 1.00 kg

Solution

What we know: kg SiO_2; kg C; balanced equation relating SiO_2 and C

Desired answer: kg of excess reactant remaining

First, determine the limiting reactant by calculating how many moles of SiC can form from each starting amount of reactant. The solution maps for these calculations are:

$$\text{kg } SiO_2 \rightarrow \text{g } SiO_2 \rightarrow \text{mol } SiO_2 \rightarrow \text{mol SiC}$$

$$\text{kg C} \rightarrow \text{g C} \rightarrow \text{mol C} \rightarrow \text{mol SiC}$$

For the first calculation the conversion factors needed are that between *kg* and *g*, the molar mass of SiO_2 and the SiC/SiO_2 mole ratio from the balanced equation.

Putting these together yields:

$$2.50\ \cancel{\text{kg } SiO_2} \times \frac{10^3\ \cancel{\text{g } SiO_2}}{1\ \cancel{\text{kg } SiO_2}} \times \frac{1\ \cancel{\text{mol } SiO_2}}{60.09\ \cancel{\text{g } SiO_2}} \times \frac{1\ \text{mol SiC}}{1\ \cancel{\text{mol } SiO_2}} = 41.6\ \text{mol SiC}$$

For the second calculation the conversion factors needed are that between *kg* and *g*, the molar mass of C and the SiC/C mole ratio from the balanced equation.

Putting these together yields:

$$2.50\ \cancel{\text{kg C}} \times \frac{10^3\ \cancel{\text{g C}}}{1\ \cancel{\text{kg C}}} \times \frac{1\ \cancel{\text{mol C}}}{12.01\ \cancel{\text{g C}}} \times \frac{1\ \text{mol SiC}}{3\ \cancel{\text{mol C}}} = 208\ \text{mol SiC}$$

Since the starting amount of SiO_2 produces the smaller amount of SiC, SiO_2 is the limiting reactant and carbon is the reactant in excess.

Now use the theoretical yield of SiC (41.6 mol) to calculate how much C was used and subtract this from the starting quantity. The solution maps for these calculations are:

$$\text{mol SiC} \rightarrow \text{mol C used} \rightarrow \text{g CO used} \rightarrow \text{kg CO used}$$

kg C remaining = kg C initially - kg C used

For the first calculation the conversion factors needed are the C/SiC mole ratio, the molar mass of C and the relationship between *g* and *kg* .

Putting these together yields:

$$41.6\ \cancel{\text{mol SiC}} \times \frac{3\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol SiC}}} \times \frac{12.01\ \cancel{\text{g C}}}{1\ \cancel{\text{mol C}}} \times \frac{1\ \text{kg C}}{10^3\ \cancel{\text{g C}}} = 1.50\ \text{kg C used}$$

Therefore, the mass of C that remains equals 2.50 kg - 1.50 kg = 1.00 kg.