

Mass Relationship



Periodic Table

1 IA 1 H 1.01	2 IIA 4 Be 9.01											13 IIIA 5 B 10.81	14 IVA 6 C 12.01	15 VA 7 N 14.01	16 VIA 8 O 16.00	17 VIIA 9 F 19.00	18 VIIIA 2 He 4.00
3 Li 6.94	12 Mg 24.31	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIII B	10	11 IB	12 IIB	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
11 Na 22.99	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.6	53 I 126.9	54 Xe 131.29
55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac^ (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (264)	108 Hs (265)	109 Mt (268)	110 Ds (271)	111 Rg (272)							

* 58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
^ 90 Th 232.0	91 Pa (231)	92 U 238.0	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Molar Calculations



The Mole and Molar Mass

- **The mole (mol)** is the amount of a substance that contains the same number of entities as there are atoms in exactly **12 g** of **carbon-12**.
 - 1 mole of glucose molecule ($\text{C}_6\text{H}_{12}\text{O}_6$) contains 6 moles of C atoms, 12 moles of H atoms, and 6 moles of O atoms.
- One mole (1 mol) contains **6.022×10^{23}** entities. This constant value is known as **Avogadro's Number (N)**.
- **The molar mass** of a substance is the mass of one mole of the substance in grams.
 - Atomic mass of Na = 22.99 amu. Molar mass of Na = 22.99 g



Sample Problem 3.1

Silver (Ag, $Z = 47$) is used in jewelry and tableware but no longer in U.S. coins. How many grams of Ag are in 0.0342 mol of Ag?

SOLUTION

$$0.0342 \text{ mol } \cancel{\text{Ag}} \times \frac{107.9 \text{ g Ag}}{1 \cancel{\text{ mol Ag}}} = 3.69 \text{ g Ag}$$



Sample Problem 3.3

Iron (Fe, $Z = 26$) is the main component of steel and is therefore the most important metal in society. How many Fe atoms are in 95.8 g of Fe?

SOLUTION

$$95.8 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \times \frac{6.022 \times 10^{23} \text{ Fe atoms}}{1 \text{ mol Fe}} = 1.03 \times 10^{24} \text{ Fe atoms}$$



Sample Problem 3.5

Ammonium carbonate, or $(\text{NH}_4)_2\text{CO}_3$, a white solid that decomposes on warming, is a component of baking powder.

- How many *formula units* are in 41.6 g of ammonium carbonate?
- How many O atoms are in this sample?

SOLUTION

a) Molar mass $(\text{NH}_4)_2\text{CO}_3 = 2(14 + 4 \times 1) + 12 + 3 \times 16 \text{ g}$
 $= 96 \text{ g.}$

$$41.6 \text{ g } (\text{NH}_4)_2\text{CO}_3 \times \frac{1 \text{ mol } (\text{NH}_4)_2\text{CO}_3}{96 \text{ g } (\text{NH}_4)_2\text{CO}_3} \times \frac{6.022 \times 10^{23} \text{ For. Unit.}}{1 \text{ mol } (\text{NH}_4)_2\text{CO}_3}$$
$$= 2.61 \times 10^{23} \text{ for. units}$$

b) $\frac{3 \text{ Oxy atoms}}{1 \text{ For. Unit } (\text{NH}_4)_2\text{CO}_3} \times (2.61 \times 10^{23}) (\text{NH}_4)_2\text{CO}_3 \text{ units}$
$$= 7.83 \times 10^{23} \text{ O atoms}$$

Mass Percent



Sample Problem 3.6

Ammonium nitrate, or NH_4NO_3 , is a common fertilizer. What is the mass percent of each element in ammonium nitrate?

SOLUTION

Molar mass $\text{NH}_4\text{NO}_3 = (14 + 4 \times 1 + 14 + 3 \times 16) \text{ g}$
 $= 80 \text{ g}$

Mass of N atoms in $\text{NH}_4\text{NO}_3 = 2 \times 14 \text{ g} = 28 \text{ g}$.

Mass of N atoms = $4 \times 14 = 56$ g
Mass of H atoms = $4 \times 1 = 4$ g

Mass of O atoms " " = $3 \times 16g = 48g$.

$$\% \text{ mass N} = \frac{28}{80} \times 100\% = 35\%$$

$$\% \text{ mass H} = \frac{4g}{80g} \times 100\% = 5\%$$

$$\% \text{ mass O} = \frac{48 \text{ g}}{80 \text{ g}} \times 100\% = 60\%$$

Molecular Formula and Empirical Formula



Molecular and Empirical Formulas

- The **molecular formula** shows the ***actual*** number of atoms of each element in a molecule.
 - The **empirical formula** is the simplest formula for a molecule — it shows the ***lowest whole number of moles*** and gives the ***relative*** number of atoms of each element present.
-
- The **molecular formula** for glucose is $\text{C}_6\text{H}_{12}\text{O}_6$ while its **empirical formula** is CH_2O .
 - The **molecular formula** for hydrogen peroxide is H_2O_2 while its **empirical formula** is HO .

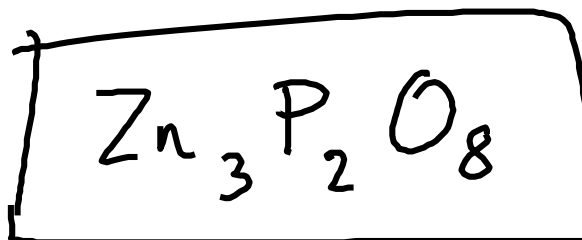


Sample Problem 3.8

A sample of an unknown compound contains 0.21 mol of zinc, 0.14 mol of phosphorus, and 0.56 mol of oxygen. What is its empirical formula?

SOLUTION

$$\begin{aligned}\text{Zn} : \text{P} : \text{O} &= 0.21 \text{ mol} : 0.14 \text{ mol} : 0.56 \text{ mol} \\ &= \frac{0.21 \text{ mol}}{0.14 \text{ mol}} : \frac{0.14 \text{ mol}}{0.14 \text{ mol}} : \frac{0.56 \text{ mol}}{0.14 \text{ mol}} \\ &= (1.5 : 1 : 4) 2 \\ &= 3 : 2 : 8\end{aligned}$$

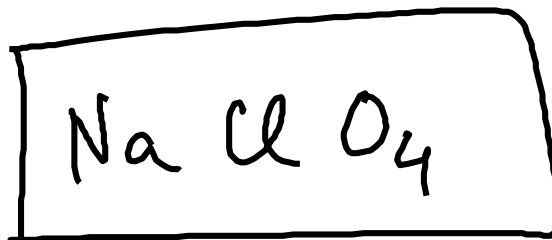


Sample Problem 3.9

Analysis of a sample of an ionic compound yields 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What is the empirical formula?

SOLUTION

$$\begin{aligned}\text{Na} : \text{Cl} : \text{O} &= 2.82\text{g} : 4.35\text{g} : 7.83\text{g} \\ &= \cancel{2.82\text{g}} \times \frac{1\text{mol Na}}{\cancel{23\text{g Na}}} : \cancel{4.35\text{g}} \times \frac{1\text{mol Cl}}{\cancel{35.5\text{g Cl}}} : \cancel{7.83\text{g}} \times \frac{1\text{mol O}}{\cancel{16\text{g O}}} \\ &= \frac{0.12\text{ mol}}{0.12} : \frac{0.12\text{ mol}}{0.12} : \frac{0.49\text{ mol}}{0.12} \\ &= 1 : 1 : 4\end{aligned}$$



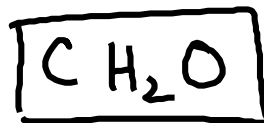
Sample Problem 3.10

Elemental analysis of lactic acid (Molar Mass = 90.08 g/mol) shows that this compound contains 40.0 mass % C, 6.71 mass % H, and 53.3 mass % O. Determine the empirical formula and the molecular formula for lactic acid.

SOLUTION

a

$$\begin{aligned} \text{C} : \text{H} : \text{O} &= 40.0\% : 6.71\% : 53.3\% \\ \text{If we consider the total mass of lactic acid to be } 100\text{g,} \\ \text{then } \text{C} : \text{H} : \text{O} &= 40.0\text{g} : 6.71\text{g} : 53.3\text{g} \\ &= 40.0\text{g} \times \frac{1\text{mol C}}{12\text{g C}} : 6.71\text{g} \times \frac{1\text{mol H}}{1\text{g H}} : 53.3\text{g} \times \frac{1\text{mol O}}{16\text{g O}} \\ &= \frac{3.33\text{mol}}{3.33} : \frac{6.71\text{mol}}{3.33} : \frac{3.33\text{mol}}{3.33} \\ &= 1 : 2 : 1 \end{aligned}$$



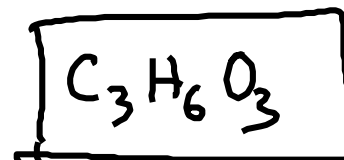
b

$n(\text{CH}_2\text{O}) = \text{molecular formula}$

$$n(\text{CH}_2\text{O}) = 90.08\text{g/mol}$$

$$n(12 + 2 \times 1 + 16)\text{g/mol} = 90.08\text{g/mol}$$

$$\Rightarrow n = 3$$



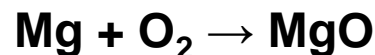
Balancing Chemical Equations



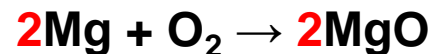
Balancing a Chemical Equation

Translate the statement

Magnesium and oxygen gas react to give magnesium oxide:



Balance the atoms using *coefficients*; *formulas cannot be changed*



Adjust coefficients if necessary

Check that all atoms balance

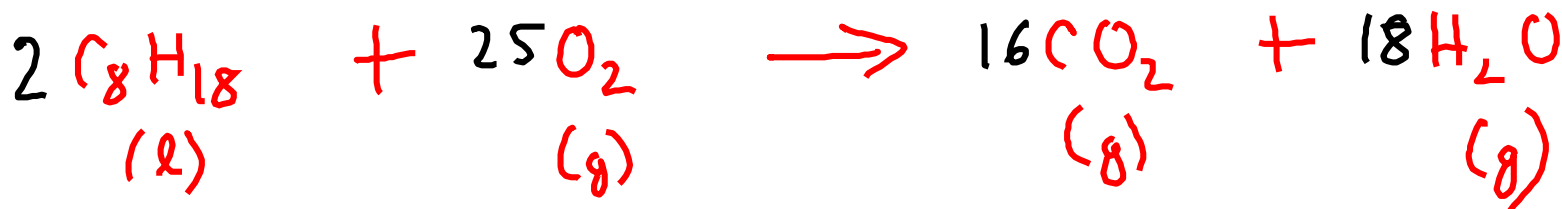
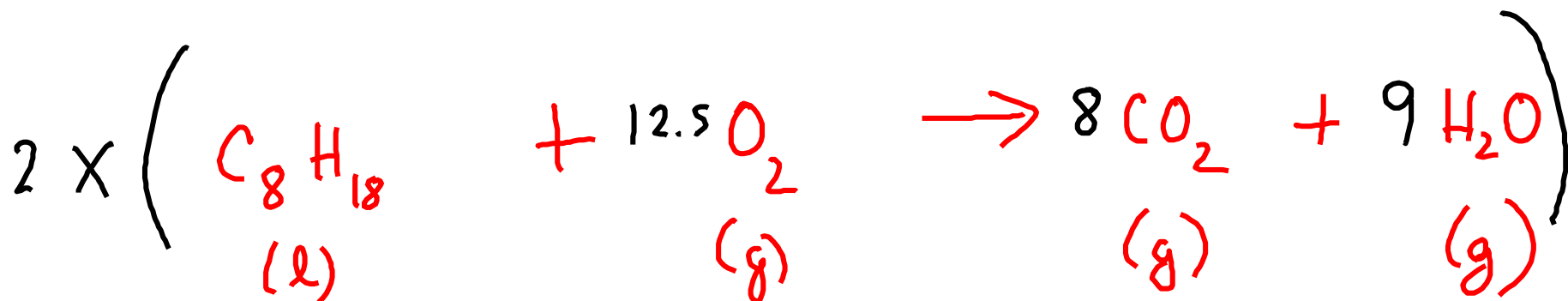
Specify states of matter



Sample Problem 3.12

Octane (C_8H_{18}) mixes with oxygen from the air and burns to form carbon dioxide and water vapor. Write a balanced equation for this reaction.

SOLUTION



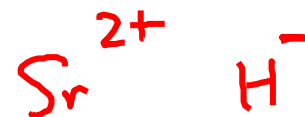
Related Problems

Problem 3.11 (d)

Calculate the molar mass of $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$.

Problem 3.15 (c)

Calculate the number of H^- ions in 5.82g of SrH_2 .



Problem 3.21 (a)

Calculate the mass % of I in strontium periodate.



Related Problems

Problem 3.41 (d)

What is the Molecular Formula when the given Empirical Formula is $\text{C}_7\text{H}_4\text{O}_2$ (Molar Mass = 240.20 g/mol).

Problem 3.44

An oxide of nitrogen contains 30.45 mass % N.

- (a) What is the empirical formula of the oxide?
- (b) If the molar mass is 90 ± 5 g/mol, what is the molecular formula?

Problem 3.46

A sample of 0.600 mol of a metal M reacts completely with excess fluorine to form 46.8 g of MF_2 .

- (a) How many moles of F are in the sample of MF_2 that forms?
- (b) How many grams of M are in this sample of MF_2 ?
- (c) What element is represented by the symbol M?



Related Problems

Problem 3.56 (b)

Balance: $\text{___P}_4\text{O}_{10} (\text{s}) + \text{___H}_2\text{O} (\text{l}) \rightarrow \text{___H}_3\text{PO}_4 (\text{l})$

Problem 3.58 (c)

Balance: $\text{___H}_3\text{PO}_4 (\text{aq}) + \text{___NaOH} (\text{aq}) \rightarrow \text{___Na}_2\text{HPO}_4 (\text{aq}) + \text{___H}_2\text{O} (\text{l})$

Problem 3.60 (b)

Balance: Liquid hexane burns in oxygen gas to form carbon dioxide gas and water vapor.

Problem 3.61 (a)

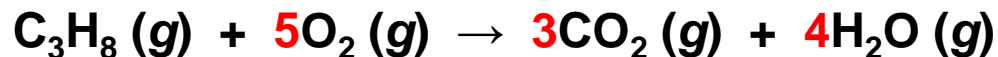
Balance: When lead(II) nitrate solution is added to potassium iodide solution, solid lead(II) iodide forms and potassium nitrate solution remains.



Reaction Stoichiometry



Stoichiometric Calculations



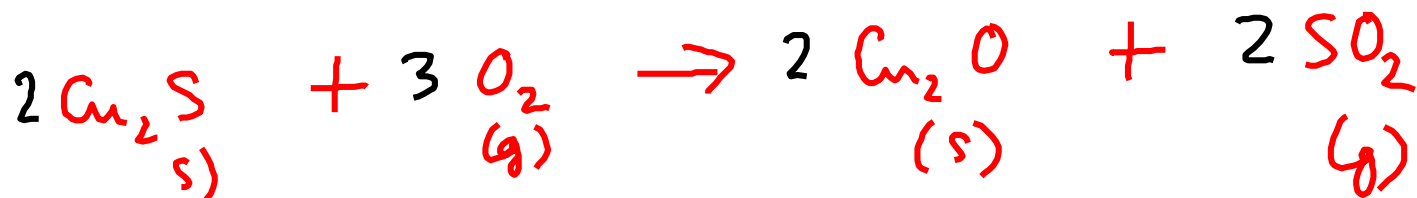
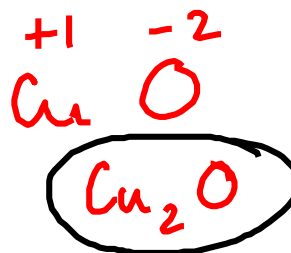
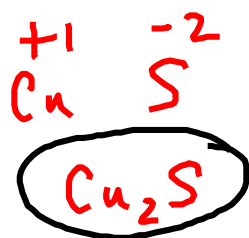
- The coefficients in a balanced chemical equation
 - represent the relative number of reactant and product particles
 - and the relative number of moles of each.
- Since moles are related to mass
 - the equation can be used to calculate masses of reactants and/or products for a given reaction.
- The **mole ratios** from the balanced equation are used as **conversion factors**.



Sample Problem 3.14

Copper is obtained from copper(I) sulfide by roasting it in the presence of oxygen gas to form powdered copper(I) oxide and gaseous sulfur dioxide. How many moles of oxygen are required to roast 10.0 mol of copper(I) sulfide?

SOLUTION

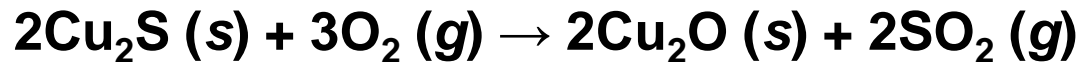


$$10.0 \text{ mol } \cancel{\text{Cu}_2\text{S}} \times \frac{3 \text{ mol } \text{O}_2}{2 \text{ mol } \cancel{\text{Cu}_2\text{S}}} = 15.0 \text{ mol } \text{O}_2$$



Sample Problem 3.15

During the process of roasting copper(I) sulfide, how many grams of sulfur dioxide form when 10.0 mol of copper(I) sulfide reacts?



SOLUTION

$$10.0 \text{ mol } \cancel{\text{Cu}_2\text{S}} \times \frac{2 \text{ mol } \cancel{\text{SO}_2}}{2 \text{ mol } \cancel{\text{Cu}_2\text{S}}} \times \frac{64 \text{ g } \text{SO}_2}{1 \text{ mol } \cancel{\text{SO}_2}}$$

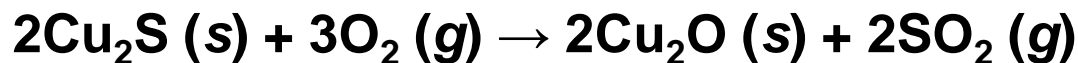
$$= \boxed{640 \text{ g SO}_2}$$

$$\begin{aligned} \text{M.M. SO}_2 &= (32 + 2 \times 16) \text{ g} \\ &= 64 \text{ g} \end{aligned}$$



Sample Problem 3.16

During the roasting of copper(I) sulfide, how many kilograms of oxygen are required to form 2.86 kg of copper(I) oxide?



SOLUTION

$$2.86 \text{ kg } \cancel{\text{Cu}_2\text{O}} \times \frac{1000 \text{ g } \cancel{\text{Cu}_2\text{O}}}{1 \text{ kg } \cancel{\text{Cu}_2\text{O}}} \times \frac{1 \text{ mol } \cancel{\text{Cu}_2\text{O}}}{142 \text{ g } \cancel{\text{Cu}_2\text{O}}} \times \frac{3 \text{ mol } \cancel{\text{O}_2}}{2 \text{ mol } \cancel{\text{Cu}_2\text{O}}}$$

$$\times \frac{32 \text{ g } \cancel{\text{O}_2}}{1 \text{ mol } \cancel{\text{O}_2}} \times \frac{1 \text{ kg } \cancel{\text{O}_2}}{1000 \text{ g } \cancel{\text{O}_2}}$$

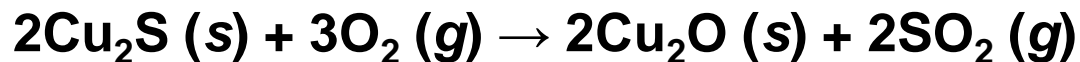
$$= \boxed{0.96 \text{ kg O}_2}$$

$$\begin{aligned} \text{M.M. Cu}_2\text{O} &= 142 \text{ g} \\ \text{M.M. O}_2 &= 32 \text{ g} \end{aligned}$$



Limiting Reactant

- So far we have assumed that reactants are present in the correct amounts to react completely.
- In reality, one reactant may *limit* the amount of product that can form.
- The *limiting* reactant gets completely used up in the reaction.
- The reactant that is **not limiting** is in **excess** – some of this reactant will be left over.



Sample Problem 3.19

In the balanced equation:



0.750 mol of Cl_2 reacts with 3.00 mol of F_2 . Find the limiting reactant.

SOLUTION

$$0.750 \text{ mol } \cancel{\text{Cl}_2} \times \frac{2 \text{ mol ClF}_3}{1 \text{ mol } \cancel{\text{Cl}_2}} = 1.50 \text{ mol ClF}_3$$

$$3.00 \text{ mol } \cancel{\text{F}_2} \times \frac{2 \text{ mol ClF}_3}{3 \text{ mol } \cancel{\text{F}_2}} = 2.00 \text{ mol ClF}_3$$

Cl_2 is
limiting
reactant

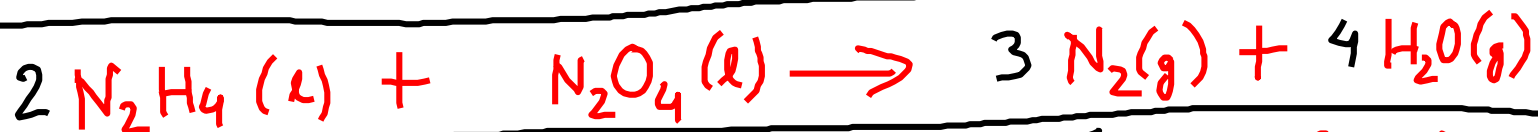


Sample Problem 3.20

Hydrazine (N_2H_4) and dinitrogen tetroxide (N_2O_4) ignite on contact to form nitrogen gas and water vapor.

- (a) How many grams of nitrogen gas form when $1.00 \times 10^2 \text{ g}$ of N_2H_4 and $2.00 \times 10^2 \text{ g}$ of N_2O_4 are mixed?

SOLUTION



i

$$1.00 \times 10^2 \text{ g } \cancel{\text{N}_2\text{H}_4} \times \frac{1 \text{ mol } \cancel{\text{N}_2\text{H}_4}}{32 \text{ g } \cancel{\text{N}_2\text{H}_4}} \times \frac{3 \text{ mol } \cancel{\text{N}_2}}{2 \text{ mol } \cancel{\text{N}_2\text{H}_4}} \times \frac{28 \text{ g } \text{N}_2}{1 \text{ mol } \cancel{\text{N}_2}} = 1.31 \times 10^2 \text{ g } \text{N}_2$$

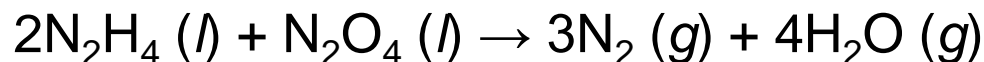
ii

$$2.00 \times 10^2 \text{ g } \cancel{\text{N}_2\text{O}_4} \times \frac{1 \text{ mol } \cancel{\text{N}_2\text{O}_4}}{92 \text{ g } \cancel{\text{N}_2\text{O}_4}} \times \frac{3 \text{ mol } \cancel{\text{N}_2}}{1 \text{ mol } \cancel{\text{N}_2\text{O}_4}} \times \frac{28 \text{ g } \text{N}_2}{1 \text{ mol } \cancel{\text{N}_2}} = 1.82 \times 10^2 \text{ g } \text{N}_2$$

$\text{N}_2\text{H}_4 \rightarrow \text{lim. reactant}$
 $1.31 \times 10^2 \text{ g of } \text{N}_2$

Sample Problem 3.20 (continued)

(b) How many grams of the excess reactant remain unreacted when the reaction is over?



SOLUTION

$$1.00 \times 10^2 \text{ g } \cancel{\text{N}_2\text{H}_4} \times \frac{1 \cancel{\text{ mol N}_2\text{H}_4}}{32 \text{ g } \cancel{\text{N}_2\text{H}_4}} \times \frac{1 \cancel{\text{ mol N}_2\text{O}_4}}{2 \cancel{\text{ mol N}_2\text{H}_4}} \times \frac{92 \text{ g N}_2\text{O}_4}{1 \cancel{\text{ mol N}_2\text{O}_4}} \\ = 1.43 \times 10^2 \text{ g N}_2\text{O}_4$$

Excess N_2O_4 left behind after the reaction is complete = $(2.00 \times 10^2 - 1.43 \times 10^2) \text{ g}$
 $= 57 \text{ g.}$

Percentage Yield

- The **theoretical yield** is the amount of product *calculated* using the mole ratios from the balanced equation.
- The **actual yield** is the amount of product actually obtained in an experiment or any reaction process.
- The actual yield is usually less than the theoretical yield.

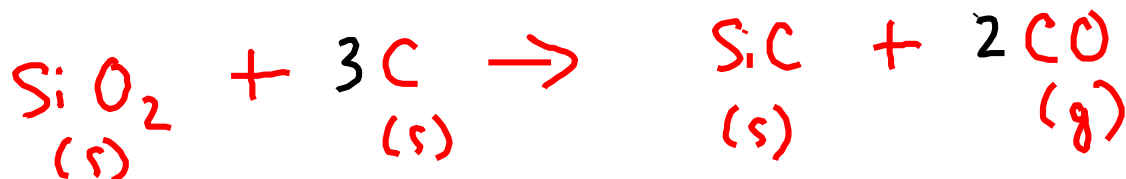
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$



Sample Problem 3.21

Silicon carbide (SiC) is made by reacting sand (silicon dioxide, SiO₂) with powdered carbon at high temperature. Carbon monoxide is also formed. What is the percent yield if 51.4 kg of SiC is recovered from processing 100.0 kg of sand?

SOLUTION



$$\begin{aligned} & 100.0 \text{ kg SiO}_2 \times \frac{1000 \text{ g SiO}_2}{1 \text{ kg SiO}_2} \times \frac{1 \text{ mol SiO}_2}{60 \text{ g SiO}_2} \times \frac{1 \text{ mol SiC}}{1 \text{ mol SiO}_2} \\ & \quad \times \frac{40 \text{ g SiC}}{1 \text{ mol SiC}} \times \frac{1 \text{ kg SiC}}{1000 \text{ g SiC}} \\ & = \boxed{66.6 \text{ kg SiC}} \leftarrow \text{Theoretical yield} \end{aligned}$$

$$\% \text{ yield} = \frac{51.4 \text{ kg SiC}}{66.6 \text{ kg SiC}} \times 100 \% = \boxed{77.1 \%}$$

Related Problems

Problem 3.70

Potassium nitrate decomposes on heating to produce potassium oxide and gaseous nitrogen and oxygen. Write a balanced equation for the process.

To produce 56.6 kg of oxygen, how many **(a)** moles and **(b)** grams of potassium nitrate must be heated?

Problem 3.71

Chromium (III) oxide reacts with hydrogen sulfide (H_2S) gas to form chromium (III) sulfide and water. Write a balanced equation for the process.

To produce 421 g of chromium (III) sulfide, how many **(a)** moles and **(b)** grams of chromium (III) oxide are required?

Problem 3.75

Elemental sulfur occurs as octatomic molecules, S_8 . What mass (in grams) of fluorine gas is needed to react completely with 17.8 g of sulfur to form sulfur hexafluoride?



Related Problems

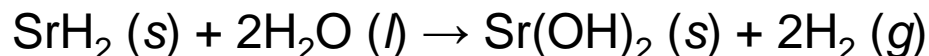
Problem 3.76

Solid iodine trichloride is prepared in two steps: first, a reaction between solid iodine and gaseous chlorine to form solid iodine monochloride; then treatment with more chlorine.

- (a) Write a balanced equation for each step.
- (b) Write a balanced equation for the overall reaction.
- (c) How many grams of iodine are needed to prepare 2.45 kg of final product?

Problem 3.79

Metal hydrides react with water to form hydrogen gas and the metal hydroxide:



You wish to calculate the mass in grams of H_2 gas that can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O .

- (a) How many moles of H_2 can be produced from the given mass of SrH_2 ?
- (b) How many moles of H_2 can be produced from the given mass of H_2O ?
- (c) Which is the limiting reactant?
- (d) How many grams of H_2 can be produced?



Related Problems

Problem 3.85

Calcium nitrate and ammonium fluoride react to form calcium fluoride, dinitrogen monoxide, and water vapor. How many grams of each substance are present after 16.8g of calcium nitrate and 17.50g of ammonium fluoride react completely?

Problem 3.86

Two successive reactions, $A \rightarrow B$ and $B \rightarrow C$, have yields of 73% and 68%, respectively. What is the overall percent yield for conversion of $A \rightarrow C$?

Problem 3.96

Sodium borohydride (NaBH_4) is used industrially in many organic syntheses. One way to prepare it is by reacting sodium hydride with gaseous diborane (B_2H_6). Assuming an 88.5% yield, how many grams of NaBH_4 can be prepared by reacting 7.98g of sodium hydride and 8.16g of diborane?



