

The Roadmap

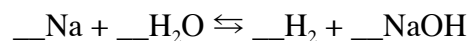
Molecules A ----- Moles A ----- Mass A ----- Volume A

Balanced Chemical Equation

Molecules B ----- Moles B ----- Mass B ----- Volume B

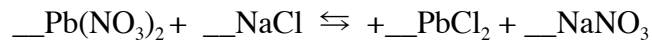
Mass A to Mass B

For the reaction:



10.00 g of Na reacts with excess water. How many grams of NaOH can be produced?

For the reaction:



What mass of PbCl_2 can be formed from the reaction of 50.0 mL of 0.250 M NaCl with excess lead (II) nitrate?

Percent Yields

A 1.00 g sample of Ammonium Chloride (NH_4Cl) is decomposed. 0.25 g of ammonia are produced. What is the percent yield?

Reaction:

Conversions:

Final Answer

Consider 100.0 g of CaC_2 reacted with excess water to produce $\text{Ca}(\text{OH})_2$ and C_2H_2 . What is the percent yield if 28.3 g of C_2H_2 is produced?

Reaction:

Conversions:

Final Answer

Excess Reactants

Consider 100.0 g of CCl_4 mixed with 30.00 g of HF. What mass of CCl_2F_2 can be produced? How much excess reactant is left over?

Reaction:

Conversion for first reactant:

Conversion for second reactant:

Work backwards for the excess:

What mass of excess remains in a reaction mixture that consists of 1.54 g of $\text{Cr}(\text{NO}_3)_3$ dissolved in 120.0 mL of 0.100 M H_2S ?

Reaction:

Conversion for first reactant:

Conversion for second reactant:

Work backwards for the excess:

Now you practice!

A volume of 3.42 mL of 0.500M AgNO_3 is reacted with 2.50 g of copper (II) chloride. How many grams of the excess reactant is left over?

Phosphorus is P_4 and it reacts with oxygen gas to make P_4O_{10} . How much of the excess reactant remains if 3.75 g of P_4 is mixed with 6.55 g of O_2 ?

Calculate the % yield when 5.23 g of ZnCl_2 is mixed with 35.0 mL of 0.325 M AgNO_3 and 1.35 g of AgCl is formed. Be careful writing the equation!

HCN and Water are produced by the reaction of ammonia, oxygen gas and methane(CH_4). If you mix 35.00 grams of methane and 35.00 grams of ammonia with excess oxygen and you get 53.22 g of HCN what is the percent yield?

Formula Problems

Remember empirical versus molecular formulas.

Chemical formulas are based on what type of ratios?

Determine the molecular formulas from the empirical formula of the following compounds based on their molecular weights:

CH	78.13	HCO ₂	90.00
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NH ₂ Cl	103	CH ₂ O	180
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Percent by Mass

Find the percent by mass of Carbon in CH₄.

Find the percent by mass of Hydrogen in PH₃.

How do you determine an empirical formula?

- 1.
- 2.
- 3.
- 4.

Using the percent by mass data find the empirical formula of dichloroethane.
Carbon 24.27%, Hydrogen 4.07%, Chlorine 71.65%.

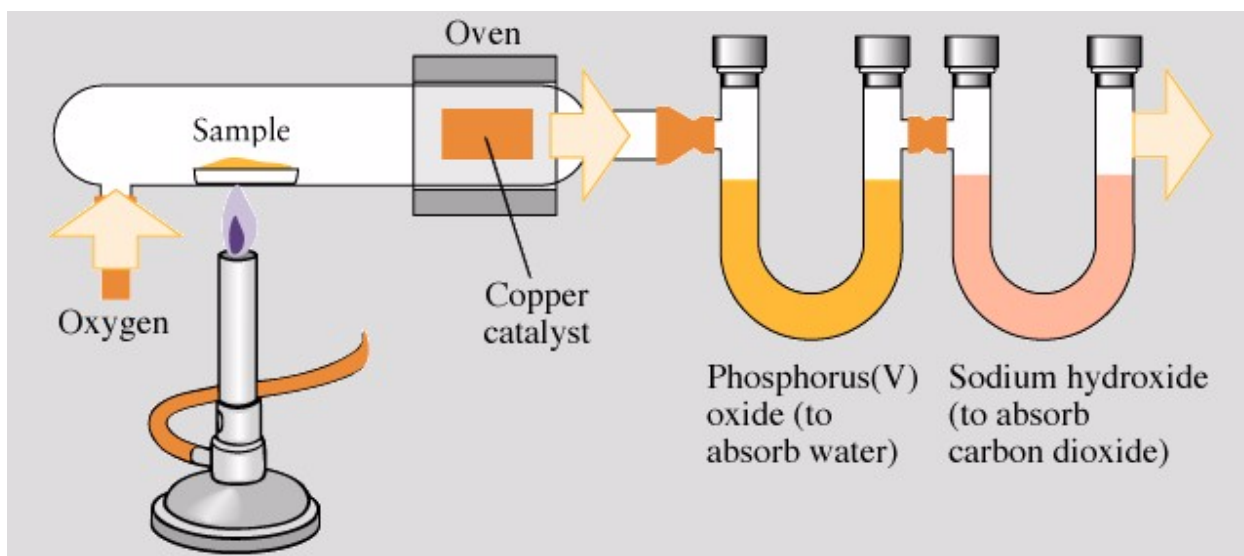
What is the molecular formula for this compound if it has a molar mass of 98.96 grams per mole?

A compound that contains only nitrogen and oxygen is 30.4% nitrogen by mass. The molar mass of the compound is 92. What is the molecular formula of this compound?

A compound is found to be 40.92% Carbon, 4.58% Hydrogen, and 54.50 % Oxygen.
What is the empirical and molecular formula for it if it has a molar mass of 176 g/mole?

Combustion Problems

What happens when something burns?



1) A 1.000g sample of an alcohol containing only carbon, hydrogen, and oxygen was combusted to produce 1.913 g of carbon dioxide and 1.174 g of water. The molar mass of the compound is 92 g/mole. What is the molecular formula of the compound?

2) Cumene is a hydrocarbon that is used in organic synthesis. Combustion of 47.6 grams of it produces 156.8 g of carbon dioxide and 42.8 g of water. If it has no elements besides hydrogen and carbon and it has a molar mass between 115 and 125 what is its molecular formula?

3) A compound contains only carbon, hydrogen, and oxygen. A 10.68 g sample is combusted producing 16.01 g of carbon dioxide and 4.37 grams of water. Mass spectroscopy reveals that it has a molar mass of 176.1 grams per mole. What is its molecular formula?

4) A compound contains carbon, hydrogen, oxygen, and nitrogen. Combustion of 0.157 g of the compound produced 0.213 g of carbon dioxide and 0.0310 g of water. In a second experiment 0.103 g of the compound produces 0.0230 g of NH_3 . What is the empirical formula of the compound?

Lab Based Problems

From the 1990 exam

An experiment is performed to determine the empirical formula of a copper iodide formed by direct combination of elements. A clean strip of copper metal is weighed accurately. It is suspended in a test tube containing iodine vapor generated by heating solid iodine. A white compound forms on the strip of copper, coating it uniformly. The strip with the adhering compound is weighed. Finally, the compound is washed completely from the surface of the metal and the clean strip is dried and reweighed.

DATA TABLE

Mass of clean copper strip	1.2789 grams
Mass of copper strip and compound	1.2874 grams
Mass of copper strip after washing	1.2748 grams

- (a) State how you would use the data above to determine each of the following.
(Calculations not required.)

(1) The number of moles of iodine that reacted

(2) The number of moles of copper that reacted

- (b) Explain how you would determine the empirical formula for the copper iodide.

- (c) Explain how each of the following would affect the empirical formula that could be calculated.

(1) Some unreacted iodine condensed on the strip.

(2) A small amount of the white compound flaked off before weighing.

The very classic lab for stoichiometry is either the copper cycle or the synthesis of Alum. This question deals with Alum and was asked in 2005.

Former AP Question

In an experiment, a student synthesis alum, $2\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}(s)$, by reacting aluminum metal with potassium hydroxide and sulfuric acid, as represented in the balanced equation above.

(a) In order to synthesize alum, the student must prepare a 5.0 M solution of sulfuric acid. Describe the procedure for preparing 50.0 mL of 5.0 M H_2SO_4 using any of the chemicals and equipment listed below. Indicate specific amounts and equipment where appropriate.

10.0 M H_2SO_4

50.0 mL volumetric flask

Distilled water

50.0 mL buret

100 mL graduated cylinder

25.0 mL pipet

100 mL beaker

50 mL beaker

(b) Calculate the minimum volume of 5.0 M H_2SO_4 that the student must use to react completely with 2.7 g of aluminum metal.

(c) As the reaction solution cools, alum crystals precipitate. The student filters the mixture and dries the crystals, then measures their mass.

(i) If the student weighs the crystals before they are completely dry, would the calculated percent yield be greater than, less than, or equal to the actual percent yield? Explain

(ii) Cooling the reaction solution in an ice bath improves the percent yield. Explain.

(d) The student heats crystals of pure alum, $\text{KAl}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$, in an open crucible to a constant mass. The mass of the sample after heating is less than the mass before heating. Explain