

I'm not a bot



Go to Ten Examples Go to a discussion of empirical and molecular formulas. Return to Mole Table of Contents Problem #1: 0.487 grams of quinine (molar mass = 324 g/mol) is combusted and found to produce 1.321 g CO₂, 0.325 g H₂O and 0.0421 g nitrogen. Determine the empirical and molecular formulas. Problem #2: 95.6 mg of menthol (molar mass = 156 g/mol) are burned in oxygen gas to give 26.9 mg CO₂ and 11.0 mg H₂O. What is menthol's empirical formula? Problem #3: 0.1005 g of menthol is combusted, producing 0.2829 g of CO₂ and 0.1159 g of H₂O. What is menthol's empirical formula? (Yes, the answer will be the same as #1,2. Go to answers for 1, 2, and 3 Problem #4: The combustion of 40.10 g of a compound which contains only C, H, Cl and O yields 58.57 g of CO₂ and 14.98 g of H₂O. Another sample of the compound with a mass of 75.00 g is found to contain 22.06 g of Cl. What is the empirical formula of the compound? Solution #1: 1) Get grams of each element. Carbon: (58.57 g) (12.011 / 44.009) = 15.985 g of C in 40.10 g sample Hydrogen: (14.98 g) (2.016 / 18.015) = 1.6764 g of H in 40.10 g sample Oxygen: we leave this to later (see below for an interesting solution path that involves determining the mass of oxygen.) Chlorine: problem gives 22.06 g in 75.00 g sample 2) Let us determine the percent composition: Carbon: 15.985 g / 40.10 g = 39.86% Hydrogen: 1.6764 g / 40.10 = 4.18% Chlorine: 22.06 / 75.00 = 29.41% Oxygen: 100% - (39.86 + 4.18 + 29.41) = 26.55% 3) Let us assume 100 g of the compound. In which case, the percentages above become grams. Now, let us determine the moles of each (I'll skip typing the calcs): C: 3.32 mol H: 4.147 mol O: 1.66 mol Cl: 0.83 mol 4) Divide each by 0.83 C: 4.0 H: 5.0 C: 1. The empirical formula is C₄H₅ClO₂ Solution #2: 1) Determine mass of all four compounds in the 40.10 g sample: Carbon: 15.985 g Hydrogen: 1.676 g Here comes the interesting way that is different from solution #1: Chlorine: (40.10 g) (22.06 g Cl / 75.00 g sample) = 11.795 g Oxygen: 40.10 g - 15.985 g - 1.676 g - 11.795 g = 10.644 g Pretty slick, huh? Notice how the oxygen is determined by subtraction after everything else is calculated. This is a common pattern in combustion analysis. 2) Determine the moles of each (I'll skip typing the calcs): C = 1.331 mol H = 1.663 mol O = 0.3327 mol 0.6653 mol Note that there is no need to assume 100 g of the compound and work from the percent composition. 3) Divide all by the smallest to simplify: C = 4 H = 5 Cl = 1 O = 2 The empirical formula is C₄H₅ClO₂ Problem #5: The empirical formula is C₄H₅ClO₂ Problem #6: The empirical formula of 1.38 grams of a compound which contains C, H, O and N yields 1.72 grams of CO₂ and 1.18 grams of H₂O. Another sample of the compound with a mass of 22.34 grams is found to contain 6.75 grams of O. What is the empirical formula of the compound? 1) Calculate grams of C and H: carbon: (1.72 g) (12.011 / 44.009) = 0.4694 g of C hydrogen: (1.18 g) (2.016 / 18.015) = 0.13205 g of H oxygen: see next step 2) Calculate mass percent of each element: carbon: 0.4694 g / 1.38 g = 34.01% hydrogen: 0.13205 g / 1.38 g = 9.57% oxygen: 6.75 g / 22.34 g = 30.215% nitrogen: 100 - (34.01 + 9.57 + 30.215) = 26.205% 3) Assume 100 g of compound present. Therefore: carbon: 34.01 g hydrogen: 9.57 g oxygen: 30.215 g nitrogen: 26.205 g 4) Calculate moles: carbon: 34.01 g / 12.011 g/mol = 2.832 hydrogen: 9.57 g / 1.008 g/mol = 9.494 oxygen: 30.215 g / 16.00 g/mol = 1.888 nitrogen: 26.205 g / 14.007 g/mol = 1.871 5) Look for smallest, whole-number ratio: carbon: 2.832 / 1.871 = 1.5 (x = 2 = 3) hydrogen: 9.494 / 1.871 = 5 (x = 2 = 10) oxygen: 1.888 / 1.871 = 1 (x = 2 = 10) Empirical formula: C₃H₁₀N₂O₂ Problem #6: The combustion of 3.42 g of a compound is known to contain only nitrogen and hydrogen gave 9.82 g of NO₂ and 3.85 g of water. Determine the empirical formula of this compound. 1) Calculate moles of N and moles of H in the combustion products: Moles N 9.82 g NO₂ / 46.0 g/mol = 0.213 mol NO₂ (0.213 mol N) (1 mol N / 1 mol NO₂) = 0.213 mol N Moles H 3.85 g H₂O / 18.0 g/mol = 0.213 mol H₂O (0.213 mol H₂O) (2 mol H / 1 mol H₂O) = 0.428 mol H 2) Calculate the ratio of moles by dividing both by the smaller: N -> 0.213 / 0.213 = 1 H -> 0.428 / 0.213 = 2 The empirical formula is NH₂ Problem #7: A compound with a known molecular weight (146.99 g/mol) that contains only C, H, and Cl was studied by combustion analysis. When a 0.367 g sample was combusted, 0.659 g of CO₂ and 0.0892 g of H₂O formed. What are the empirical and molecular formulas? Solution: 1) Carbon: 0.659 g of CO₂ has 0.659 / 44 = 0.0150 moles of CO₂ there is 1 mole of C in CO₂ and all the C from the compound becomes CO₂, so moles of C in the compound = 0.0150 moles mass of C = 0.0150 x 12 = 0.1797 g 2) Hydrogen: 0.0892 g of H₂O has 0.0892 / 18 = 0.0050 moles of H₂O there are 2 moles of H in H₂O, so moles of H in the compound = 0.0090 moles mass of H = 0.0099 x 1.0079 = 0.0100 g 3) Chlorine: mass of H + C = 0.1897 g mass of sample = 0.3670 g mass of Cl by difference = 0.1773 g moles of Cl = 0.0050 moles 4) Smallest whole-number ratio: molar ratio of C : H : Cl = 0.0150 : 0.0099 : 0.0050 divide the ratio by the smallest number molar ratio of C : H : Cl = 1.98 : 1.5 : 1 Formulas: empirical formula is C₃H₂Cl₂ this has an "empirical formula weight" of (36+2+35.5) = 73.5 which is 1/2 the molecular mass so the molecular formula is twice the empirical formula C₆H₄Cl₄ Problem #8: A 2.52 g sample of a compound containing carbon, hydrogen, nitrogen, oxygen, and sulfur was burned in excess oxygen gas to yield 4.36 g of CO₂ and 0.892 gram of H₂O and only one other compound residue. The H₂O contained 0.14 g of hydrogen. The other residue was a white, brittle, non-volatile solid that contained 0.27 g of H and 0.1159 g of N. Determine the empirical formula of the compound. Solution: #1: Carbon: # moles = molar mass molar mass of CO₂ = 44.0 g/mol 4.36 g/mol CO₂ has 4.36 / 44 = 0.09909 moles of CO₂ there is 1 mole of C in CO₂ and all the C from the compound becomes CO₂ moles of C in the compound = 0.09909 moles mass of C -> 0.09909 x 12 = 1.1891 g 2) Hydrogen: molar mass of H₂O = 18 g/mol 0.892 g H₂O = 0.04956 moles of H₂O there are 2 moles of H in H₂O so moles of H in the compound = 0.09911 moles mass of H -> 0.09911 x 1.0079 = 0.100 g 3) Sulfur: moles of SO₃ -> 2.60 / 80 = 0.0325 moles of S in 4.14 g of compound = 0.0325 mol moles in 2.52 g of compound -> (0.0325) (2.52 / 4.14) = 0.01978 moles mass of S -> 0.01978 x 32 = 0.6330 g All the sulfur in the SO₃ came from the 4.14 g sample. Notice the scaling from 4.14 g of compound to 2.52 g. 4) Nitrogen: moles of HNO₃ -> 2.80 / 63 = 0.04444 mol moles of N in 5.66 g of sample = 0.04444 moles in 2.52 g of compound -> (0.04444) (2.52 / 5.66) = 0.01979 moles mass of N -> 0.01979 x 14 = 0.2770 g All the nitrogen in the HNO₃ came from the 5.66 g sample. Notice the scaling from 5.66 g of compound to 2.52 g. 5) Oxygen: mass C + H + S + N = 1.1891 + 0.100 + 0.6330 + 0.2770 = 2.1991 mass of O by difference = 2.52 - 2.1991 = 0.3209 g moles of O in 2.52 g = 0.3209 / 16 = 0.0201 moles 6) Now we have all five molar amounts, so do the empirical formula: molar ratio of C : H : S : N : O = 0.09909 : 0.09911 : 0.01978 : 0.0201 divide by the smallest number to get whole-number ratio C : H : S : N : O = 5 : 5 : 1 : 1 : 1 empirical formula is C₅H₅NSO₂ Problem #9: Burning 11.20 mL (measured at STP) of a gas known to contain only carbon and oxygen yields 16.203 grams of water. What is the empirical formula? Solution: 1) Determine moles of H₂O: 16.203 g H₂O / 18.015 g/mol = 0.9000 moles of H₂O. 2) Determine moles of H: 0.9000 moles of H₂O x 2 = 1.8000 moles of H. 3) Determine moles of O: 0.9000 moles of H₂O x 1 = 0.9000 moles of O. 4) Determine moles of C: 11.20 mL / 22.4 L/mol = 0.5000 moles of C. 5) Determine molar ratio: C : H : O = 0.5000 : 1.8000 : 0.9000 divide by 0.5000 to get C : H : O = 1 : 3.6 : 1.8 divide by 1.8 to get C : H : O = 1 : 2 : 1. Empirical formula is CH₂O. Problem #10: The osmotic pressure of a solution containing 2.04 g of an unknown molecular compound dissolved in 175.0 mL of solution at 25.0 °C is 2.13 atm. The combustion of 22.08 g of the unknown compound produced 36.26 g CO₂ and 14.85 g H₂O. Solution: 1) The osmotic pressure will allow us to calculate the molar mass of the substance: n = mRT / PV = (2.13 atm) = (1) (x) (0.175 L) (0.08206 L atm / mol K) (298 K) x = 0.015243 mol 2.04 g / 0.015243 mol = 133.83 g/mol 2) Let us calculate the amount of carbon and hydrogen. Then, by subtraction, we will check for oxygen: carbon -> 36.26 g x (12.011 / 44.01) = 9.8959 g hydrogen -> 14.85 g x (2.016 / 18.015) = 1.6618 g oxygen -> 22.08 minus (9.8959 + 1.6618) = 10.5223 g 3) Calculate moles of each element: carbon -> 9.8959 g / 12.011 g/mol = 0.8239 mol hydrogen -> 1.6618 g / 1.008 g/mol = 1.6486 mol oxygen -> 10.5223 g / 15.999 g/mol = 0.65768 mol 4) Determine a whole number ratio: carbon -> 0.8239 / 0.65768 = 1.25 hydrogen -> 1.6486 / 0.65768 = 2.5 oxygen -> 0.65768 / 0.65768 = 1 Look at it like this: carbon -> 1.25 = 5/4 (times 4 = 5) hydrogen -> 2.5 = 10/4 (times 4 = 10) oxygen -> 1 = 4/4 (times 4 = 4) 5) The empirical formula is: C₅H₁₀O₄ 6) The "empirical formula weight" is 436.6 g/mol. The molecular weight is 1750 g/mol. Divide 1750 g/mol by 436.6 g/mol = 4.0092. Round to 4.0. Multiply 436.6 g/mol by 4.0092 = 1750 g/mol. The molecular formula is C₂₀H₄₀O₁₆. Problem #11: A 1.000 g sample of a compound containing only C, H, and O is burned in excess oxygen to produce 2.146 g of CO₂ and 0.4842 g of H₂O. What is the empirical formula of the compound? Solution: Because everything was measured at equal temperatures and pressures, Avogadro's Hypothesis can be used to solve this problem. The hypothesis is "Equal volumes at equal temperatures and pressures contain equal number of molecules." Just to make it a bit simpler to explain, let us assume a value of T and P such that the 3.0 L of each gas contains 3.0 moles. That means 1.5 moles of the compound was initially present. 3.0 moles of CO₂ contains 3.0 moles of C and 1.5 moles of H₂O contains 3.0 moles of H and 1.5 moles of O. The mass of H₂O contains 3.0 moles of H and 1.5 moles of O. Therefore, 1.0 mol of the compound contains 2.0 mol of C and 2.0 mol of H. The formula of the compound is C₂H₂. Bonus Problem #1: A 6.20-g sample of an unknown compound containing only C, H, and O combusts in an oxygen rich environment. When the products have cooled to 20.0 °C at 1 bar, there are 8.09 L of CO₂ and 3.99 mL of H₂O. The density of water at 20.0 °C is 0.998 g/mL. a) What is the empirical formula of the unknown compound? b) If the molar mass is 168.2 g/mol, what is the molecular formula of the compound? Solution: 1) Mass of H₂O: (3.99 mL) (0.998 g/mL) = 3.982 g 2) Mass of CO₂: PV = nRT (1 bar) (8.09 L) = (n) (0.0831447 L bar / mol K) (293 K) n = 0.3320276 mol (0.3320276 mol) (44.0096 g/mol) = 14.61643 g 3) We need grams of C and H before getting grams of O: (14.61643 g) (12.011 g / 44.009 g) = 3.988646 g (3.988646 g) (2.016 g / 18.015 g) = 0.445613 g 4) Mass of O: 6.20 - (3.988646 + 0.445613) = 1.765741 g 5) Determine moles of all three components: 3.988646 g / 12.011 g/mol = 0.3320283 mol 0.445613 g / 2.016 g/mol = 0.2212074 mol 1.765741 g / 16.00 g/mol = 0.110359 mol 6) Divide each by the smaller: 0.3320283 mol / 0.2212074 mol = 1.5000 mol of C 0.2212074 mol / 0.2212074 mol = 1 mol of H 0.110359 mol / 0.2212074 mol = 0.5000 mol of O 7) Empirical formula is C₃H₂O. Problem #2: Given the following reaction: 8xH₂O(g) + 2(O₂)₃(s) + 4(C₂H₂)₂(g) → 4H₂O(g) + 1.48 g BxHy yields 16.203 grams of water. What is the empirical formula? Solution: 1) Grams of H₂O: 16.203 g H₂O / 18.015 g/mol = 0.9000 moles of H₂O. 2) Determine moles of H: 0.9000 moles of H₂O x 2 = 1.8000 moles of H. 3) Determine moles of O: 0.9000 moles of H₂O x 1 = 0.9000 moles of O. 4) Determine moles of C: 1.48 g BxHy / 12.011 g/mol = 0.12323 mol of C. 5) Determine molar ratio: C : H : O = 0.12323 : 1.8000 : 0.9000 divide by 0.12323 to get C : H : O = 1 : 14.6 : 7.3 divide by 7.3 to get C : H : O = 1 : 2 : 1. Empirical formula is CH₂O. Problem #3: A 1.000 g sample of a compound containing only C, H, and O is burned in excess oxygen to produce 2.146 g of CO₂ and 0.4842 g of H₂O. What is the empirical formula of the compound? Solution: Because everything was measured at equal temperatures and pressures, Avogadro's Hypothesis can be used to solve this problem. The hypothesis is "Equal volumes at equal temperatures and pressures contain equal number of molecules." Just to make it a bit simpler to explain, let us assume a value of T and P such that the 3.0 L of each gas contains 3.0 moles. That means 1.5 moles of the compound was initially present. 3.0 moles of CO₂ contains 3.0 moles of C and 1.5 moles of H₂O contains 3.0 moles of H and 1.5 moles of O. The mass of H₂O contains 3.0 moles of H and 1.5 moles of O. Therefore, 1.0 mol of the compound contains 2.0 mol of C and 2.0 mol of H. The formula of the compound is C₂H₂. Bonus Problem #1: A 6.20-g sample of an unknown compound containing only C, H, and O combusts in an oxygen rich environment. When the products have cooled to 20.0 °C at 1 bar, there are 8.09 L of CO₂ and 3.99 mL of H₂O. The density of water at 20.0 °C is 0.998 g/mL. a) What is the empirical formula of the unknown compound? b) If the molar mass is 168.2 g/mol, what is the molecular formula of the compound? Solution: 1) Mass of H₂O: (3.99 mL) (0.998 g/mL) = 3.982 g 2) Mass of CO₂: PV = nRT (1 bar) (8.09 L) = (n) (0.0831447 L bar / mol K) (293 K) n = 0.3320276 mol (0.3320276 mol) (44.0096 g/mol) = 14.61643 g 3) We need grams of C and H before getting grams of O: (14.61643 g) (12.011 g / 44.009 g) = 3.988646 g (3.988646 g) (2.016 g / 18.015 g) = 0.445613 g 4) Mass of O: 6.20 - (3.988646 + 0.445613) = 1.765741 g 5) Determine moles of all three components: 3.988646 g / 12.011 g/mol = 0.3320283 mol 0.445613 g / 2.016 g/mol = 0.2212074 mol 1.765741 g / 16.00 g/mol = 0.110359 mol 6) Divide each by the smaller: 0.3320283 mol / 0.2212074 mol = 1.5000 mol of C 0.2212074 mol / 0.2212074 mol = 1 mol of H 0.110359 mol / 0.2212074 mol = 0.5000 mol of O 7) Empirical formula is C₃H₂O. Problem #4: A 1.000 g sample of a compound containing only C, H, and O is burned in excess oxygen to produce 2.146 g of CO₂ and 0.4842 g of H₂O. What is the empirical formula of the compound? Solution: Because everything was measured at equal temperatures and pressures, Avogadro's Hypothesis can be used to solve this problem. The hypothesis is "Equal volumes at equal temperatures and pressures contain equal number of molecules." Just to make it a bit simpler to explain, let us assume a value of T and P such that the 3.0 L of each gas contains 3.0 moles. That means 1.5 moles of the compound was initially present. 3.0 moles of CO₂ contains 3.0 moles of C and 1.5 moles of H₂O contains 3.0 moles of H and 1.5 moles of O. The mass of H₂O contains 3.0 moles of H and 1.5 moles of O. Therefore, 1.0 mol of the compound contains 2.0 mol of C and 2.0 mol of H. The formula of the compound is C₂H₂. Bonus Problem #1: A 6.20-g sample of an unknown compound containing only C, H, and O combusts in an oxygen rich environment. When the products have cooled to 20.0 °C at 1 bar, there are 8.09 L of CO₂ and 3.99 mL of H₂O. The density of water at 20.0 °C is 0.998 g/mL. a) What is the empirical formula of the unknown compound? b) If the molar mass is 168.2 g/mol, what is the molecular formula of the compound? Solution: 1) Mass of H₂O: (3.99 mL) (0.998 g/mL) = 3.982 g 2) Mass of CO₂: PV = nRT (1 bar) (8.09 L) = (n) (0.0831447 L bar / mol K) (293 K) n = 0.3320276 mol (0.3320276 mol) (44.0096 g/mol) = 14.61643 g 3) We need grams of C and H before getting grams of O: (14.61643 g) (12.011 g / 44.009 g) = 3.988646 g (3.988646 g) (2.016 g / 18.015 g) = 0.445613 g 4) Mass of O: 6.20 - (3.988646 + 0.445613) = 1.765741 g 5) Determine moles of all three components: 3.988646 g / 12.011 g/mol = 0.3320283 mol 0.445613 g / 2.016 g/mol = 0.2212074 mol 1.765741 g / 16.00 g/mol = 0.110359 mol 6) Divide each by the smaller: 0.3320283 mol / 0.2212074 mol = 1.5000 mol of C 0.2212074 mol / 0.2212074 mol = 1 mol of H 0.110359 mol / 0.2212074 mol = 0.5000 mol of O 7) Empirical formula is C₃H₂O. Problem #5: A 1.000 g sample of a compound containing only C, H, and O is burned in excess oxygen to produce 2.146 g of CO₂ and 0.4842 g of H₂O. What is the empirical formula of the compound? Solution: Because everything was measured at equal temperatures and pressures, Avogadro's Hypothesis can be used to solve this problem. The hypothesis is "Equal volumes at equal temperatures and pressures contain equal number of molecules." Just to make it a bit simpler to explain, let us assume a value of T and P such that the 3.0 L of each gas contains 3.0 moles. That means 1.5 moles of the compound was initially present. 3.0 moles of CO₂ contains 3.0 moles of C and 1.5 moles of H₂O contains 3.0 moles of H and 1.5 moles of O. The mass of H₂O contains 3.0 moles of H and 1.5 moles of O. Therefore, 1.0 mol of the compound contains 2.0 mol of C and 2.0 mol of H. The formula of the compound is C₂H₂. 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Bonus Problem #1: A 6.20-g sample of an unknown compound containing only C, H, and O combusts in an oxygen rich environment. When the products have cooled to 20.0 °C at 1 bar, there are 8.09 L of CO₂ and 3.99 mL of H₂O. The density of water at 20.0 °C is 0.998 g/mL. a) What is the empirical formula of the unknown compound? b) If the molar mass is 168.2 g/mol, what is the molecular formula of the compound? Solution: 1) Mass of H₂O: (3.99 mL) (0.998 g/mL) = 3.982 g 2) Mass of CO₂: PV = nRT (1 bar) (8.09 L) = (n) (0.0831447 L bar / mol K) (293 K) n = 0.3320276 mol (0.3320276 mol) (44.0096 g/mol) = 14.61643 g 3) We need grams of C and H before getting grams of O: (14.61643 g) (12.011 g / 44.009 g) = 3.988646 g (3.988646 g) (2.016 g / 18.015 g) = 0.445613 g 4) Mass of O: 6.20 - (3.988646 + 0.445613) = 1.765741 g 5) Determine moles of all three components: 3.988646 g / 12.011 g/mol = 0.3320283 mol 0.445613 g / 2.016 g/mol = 0.2212074 mol 1.765741 g / 16.00 g/mol = 0.110359 mol 6) Divide each by the smaller: 0.3320283 mol / 0.2212074 mol = 1.5000 mol of C 0.2212074 mol / 0.2212074 mol = 1 mol of H 0.110359 mol / 0.2212074 mol = 0.5000 mol of O 7) Empirical formula is C₃H₂O. Problem #8: A 1.000 g sample of a compound containing only C, H, and O is burned in excess oxygen to produce 2.146 g of CO₂ and 0.4842 g of H₂O. What is the empirical formula of the compound? Solution: Because everything was measured at equal temperatures and pressures, Avogadro's Hypothesis can be used to solve this problem. The hypothesis is "Equal volumes at equal temperatures and pressures contain equal number of molecules." Just to make it a bit simpler to explain, let us assume a value of T and P such that the 3.0 L of each gas contains 3.0 moles. That means 1.5 moles of the compound was initially present. 3.0 moles of CO₂ contains 3.0 moles of C and 1.5 moles of H₂O contains 3.0 moles of H and 1.5 moles of O. The mass of H₂O contains 3.0 moles of H and 1.5 moles of O. Therefore, 1.0 mol of the compound contains 2.0 mol of C and 2.0 mol of H. The formula of the compound is C₂H₂. Bonus Problem #1: A 6.20-g sample of an unknown compound containing only C, H, and O combusts in an oxygen rich environment. When the products have cooled to 20.0 °C at 1 bar, there are 8.09 L of CO₂ and 3.99 mL of H₂O. The density of water at 20.0 °C is 0.998 g/mL. a) What is the empirical formula of the unknown compound? b) If the molar mass is 168.2 g/mol, what is the molecular formula of the compound? Solution: 1) Mass of H₂O: (3.99 mL) (0.998 g/mL) = 3.982 g 2) Mass of CO₂: PV = nRT (1 bar) (8.09 L) = (n) (0.0831447 L bar / mol K) (293 K) n = 0.3320276 mol (0.3320276 mol) (44.0096 g/mol) = 14.61643 g 3) We need grams of C and H before getting grams of O: (14.61643 g) (12.011 g / 44.009 g) = 3.988646 g (3.988646 g) (2.016 g / 18.015 g) = 0.445613 g 4) Mass of O: 6.20 - (3.988646 + 0.445613) = 1.765741 g 5) Determine moles of all three components: 3.988646 g / 12.011 g/mol = 0.3320283 mol 0.445613 g / 2.016 g/mol = 0.2212074 mol 1.765741 g / 16.00 g/mol = 0.110359 mol 6) Divide each by the smaller: 0.3320283 mol / 0.2212074 mol = 1.5000 mol of C 0.2212074 mol / 0.2212074 mol = 1 mol of H 0.110359 mol / 0.2212074 mol = 0.5000 mol of O 7) Empirical formula is C₃H₂O. Problem #9: A 1.000 g sample of a compound containing only C, H, and O is burned in excess oxygen to produce 2.146 g of CO₂ and 0.4842 g of H₂O. What is the empirical formula of the compound? Solution: Because everything was measured at equal temperatures and pressures, Avogadro's Hypothesis can be used to solve this problem. The hypothesis is "Equal volumes at equal temperatures and pressures contain equal number of molecules." Just to make it a bit simpler to explain, let us assume a value of T and P such that the 3.0 L of each gas contains 3.0 moles. That means 1.5 moles of the compound was initially present. 3.0 moles of CO₂ contains 3.0 moles of C and 1.5 moles of H₂O contains 3.0 moles of H and 1.5 moles of O. The mass of H₂O contains 3.0 moles of H and 1.5 moles of O. Therefore, 1.0 mol of the compound contains 2.0 mol of C and 2.0 mol of H. The formula of the compound is C₂H₂. Bonus Problem #1: A 6.20-g sample of an unknown compound containing only C, H, and O combusts in an oxygen rich environment. When the products have cooled to 20.0 °C at 1 bar, there are 8.09 L of CO₂ and 3.99 mL of H₂O. The density of water at 20.0 °C is 0.998 g/mL. a) What is the empirical formula of the unknown compound? b) If the molar mass is 168.2 g/mol, what is the molecular formula of the compound? Solution: 1) Mass of H₂O: (3.99 mL) (0.998 g/mL) = 3.982 g 2) Mass

