Percent Composition and Empirical Formulas
Content Objectives

• SWBAT calculate the percent composition by mass of each element in a compound.

• SWBAT calculate the empirical formula of a compound based on the percent composition by mass of each element in a compound.
What is the % composition by weight of each element for CO$_2$?

- A) 73% C and 27% O
- B) 50% C and 50% O
- C) 33% C and 67% O
- D) 27% C and 73% O
What is the % composition by weight of each element for CO$_2$?

- A) 73% C and 27% O
- B) 50% C and 50% O
- C) 33% C and 67% O
- D) 27% C and 73% O
What is the % composition by weight of each element for CO$_2$?

- How do we figure it out?

- The mass of 1 mole of CO$_2$ is 44 g/mol

- Each atom contains 1 carbon and 2 oxygens.
  - 1 mole of carbon = 12 g
  - 2 moles of oxygen = 32 g
  - 12 g/44 g = 27% for carbon
  - 32 g/44 g = 73% for oxygen
What is the % composition for each element in $\text{H}_2\text{SO}_4$?

- The molar mass is $2(1)+1(32.07)+4(16) = 98.07 \text{ g/mol}$
  - For H it is $2/98.07 = 2.0\%$
  - For S it is $32.07/98.07 = 32.7\%$
  - For O it is $64/98.07 = 65.3\%$
A Molecule Only Contains Nitrogen and Oxygen

• Analysis shows that 30.4% is Nitrogen by mass and 69.6% is Oxygen by mass.

• What molecule is it?

• \( \text{N}_2\text{O}_4 \) or \( \text{NO}_2 \) would both work out the same!
Empirical Formula

• Simplest number of each element, regardless of actual molecular geometry, number of bonds, or chemical formulas.

• Example: $\text{N}_2\text{O}_4$ can be simplified to $\text{NO}_2$
Empirical formulas are always used for ionic compounds

- Sodium chloride is not just one Na and one Cl.
- It is a crystal of ionically bonded Na’s and Cl’s.
- Thus, the empirical formula is NaCl
- In reality, a single crystal it is unlimited
Network Solids (Covalent Crystals)

• SiO$_2$ does not normally exist as a discrete 3-atom compound, but rather a network of alternating Si and O atoms that average out to SiO$_2$ over the entire crystal.
Example of Empirical Formulas (CH$_2$O)

- CH$_2$O  Formaldehyde
- C$_2$H$_4$O$_2$  Acetic Acid
- C$_6$H$_{12}$O$_6$  Glucose
Why Use Empirical Formulas

• Ionic compounds always have empirical formulas because their crystals can contain unlimited numbers of each.

• Sometimes you may come across an unknown compound (maybe an illegal drug or toxic chemical), and there is no way to definitively tell what it is just by looking.
Say you received a letter with a suspicious white power


• How many different tests are you going to try to figure it out?
Using Mass Spectrometers

• With a mass spectrometer you can determine what elements in what relative composition are in the substance.

• Then, you can simplify your other tests to focus on just the possible compounds with that empirical formula.
Example 1: Which Antimony Chloride Is It?

- Antimony forms two chlorides with very different properties.
- One chloride is a solid and melts at 73 °C and is 46.8% chlorine by mass.
- The other is a liquid that boils at 140 °C and contains 59.3% chlorine by mass.

- Determine the empirical formula for each.
Which Antimony chloride is it?

• Pretend you have 100 g of it.
• If one compound is 46.8% chlorine, then in a 100 gram sample, you have 46.8 g of Cl and 53.2 g of Sb.
• Then, determine the number of moles of each to see what the ratio is.

• $46.8 \text{ g of Cl} / 35.45 \text{ g/mol of Cl} = 1.32 \text{ moles of Cl}$

• $53.2 \text{ g of Sb} / 121.76 \text{ g/mol of Sb} = 0.44 \text{ moles of Sb}$
Which Antimony chloride is it?

• Then, determine the number of moles of each to see what the ratio is.
  
  • 46.8 g of Cl/35.45 g/mol of Cl = 1.32 moles of Cl
  • 53.2 g of Sb/121.76 g/mol of Sb = 0.44 moles of Sb
  
  • A ratio of 1.32:0.44 can be simplified to whole numbers by dividing both by the smallest number. 1.32/0.44 = 3 Cl and 0.44/0.44 = 1 Sb
  
  • This compound is SbCl$_3$
The other Antimony chloride

- A compound that is 59.3% Cl and 40.9% Sb

- 59.3 g of Cl/35.45 g/mol of Cl = 1.72 moles of Cl
- 53.2 g of Sb/121.76 g/mol of Sb = 0.34 moles of Sb

- A ratio of 1.72:0.34 can be simplified to whole numbers by dividing both by the smallest number. 1.72/0.34 = 5 Cl and 0.34/0.34 = 1 Sb

- This compound is SbCl$_5$
Example 2: Iron and Oxygen

• Say you were given a compound containing Fe at 70% of the mass and O at 30%.

• Fe = 55.85 \( \frac{g}{\text{mol}} \)
• O = 16.00 \( \frac{g}{\text{mol}} \)
Iron and Oxygen

• Say you were given a compound containing Fe at 70% of the mass and O at 30%.

• $\frac{70}{55.85} = 1.253$ moles Fe
• $\frac{30}{16} = 1.875$ moles O

• $\text{Fe}_{1.253}\text{O}_{1.875}$
Iron and Oxygen

• Say you were given a compound containing Fe at 70% of the mass and O at 30%.

• \( \text{Fe}_{1.253}\text{O}_{1.875} \)

• Divide both sides by the lower number

• \( 1.875/1.253 = 1.5 \) moles for O

• \( 1.253/1.253 = 1 \) mole for Fe

• \( \text{Fe}_{1}\text{O}_{1.5} \)

Now what?
What Happens if it Doesn’t Make a Whole Number?

• What if what you end up with is $\text{Fe}_1\text{O}_{1.5}$?

• Just like with balancing equations, find a common factor and multiply out to get both into whole numbers.

• $2(\text{Fe}_1\text{O}_{1.5}) = \text{Fe}_2\text{O}_3$
Empirical Formulas of Compounds With More Than Two Elements

• Find the empirical formula of a compound that is 48.38% carbon, 8.12% hydrogen, and 53.5% oxygen by mass.

• Pretend you have 100 grams of this compound.
Pretend You Have 100 grams

- Carbon: \((0.4838) \times 100 \text{ g} = 48.38 \text{ g C}\)
- Hydrogen: \((0.0812) \times 100 \text{ g} = 8.12 \text{ g H}\)
- Oxygen: \((0.5350) \times 100 \text{ g} = 53.38 \text{ g O}\)
Convert Mass to Moles

• Convert the masses to moles

• \((48.38 \text{ g C}) \times \left( \frac{1 \text{ mol}}{12.0 \text{ g C}} \right) = 4.028 \text{ mol C} \)

• \((8.12 \text{ g H}) \times \left( \frac{1 \text{ mol}}{1.01 \text{ g H}} \right) = 8.056 \text{ mol H} \)

• \((53.38 \text{ g O}) \times \left( \frac{1 \text{ mol}}{16.0 \text{ g O}} \right) = 3.336 \text{ mol O} \)
Find the Ratios of Each by Dividing Into the Smallest One

- \( \frac{3.336 \text{ mol O}}{3.336} = 1 \text{ mol O} \)
- \( \frac{4.028 \text{ mol C}}{3.336} = 1.2 \text{ mol C} \)
- \( \frac{8.056 \text{ mol H}}{3.336} = 2.4 \text{ mol H} \)

- Find the Common Factor
  - \( 5(1 \text{ O: } 1.2 \text{ C: } 2.4 \text{ H}) = 5\text{O} : 6\text{C} : 12 \text{ H} \)

- Thus it is \( \text{C}_6\text{H}_{12}\text{O}_5 \)