|  |
| --- |
| UNIT 5  The Mole |

|  |  |
| --- | --- |
| **KEY IDEAS** | |
| Vocabulary | What does it mean? |
| relative mass |  |
| atomic mass |  |
| Avogadro's Number |  |
| mole |  |
| molar mass |  |
| standard temperature and pressure (STP) |  |
| molar volume |  |
| percent composition |  |
| empirical formula |  |
| molecular formula |  |

|  |
| --- |
| **5.0 - John Dalton & Relative Mass** |

John Dalton found mass of atoms by finding **relative mass:** mole ratio between two elements

1. assign the lightest element (H) mass of \_\_\_\_\_\_\_\_
2. compare all other elements to H

|  |
| --- |
| **Example**: The reaction between 2.74 g of hydrogen gas and 97.26 g of chlorine gas makes 100 g of hydrogen chloride. What is the relative mass of a chlorine atom to a hydrogen atom? |

**Atomic mass:** average mass of an atom measured in **atomic mass units (amu)**

6

C

Carbon

12.0

Dalton’s mistake: he didn’t know that some elements are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

|  |  |  |
| --- | --- | --- |
| Element | Compare to H | Atomic Mass (amu) |
| H | 1 |  |
| C | 6 x as heavy |  |
| O | 16 x as heavy |  |

We still have a problem: atoms are so tiny. How do we make sense of amu? Just how much is 1 amu? How do we measure amu? We need a unit that we can use!

|  |
| --- |
| **5.1 - Avogadro's Number** |

**Mole**: a number to describe the amount of matter

* the mole is also called Avogadro's Number

|  |
| --- |
| **Practice**: Perform the following conversions.   1. How many moles of eggs do I have if I have 144 eggs? 2. How many moles of carbon atoms do I have if I have 845 carbon atoms? 3. How many moles of bicycles are 6 x 1025 bicycles? How many moles of bicycle wheels are in 6 x 1025 bicycles? 4. How many moles of oxygen atoms are in 1 mole of carbon dioxide |

|  |
| --- |
| **5.2 - Molar Mass** |

Where did Avogadro’s number come from???

* a scientist named Avogadro found that the number of carbon atoms in 12 g is 6.02 x 1023
* in other words, carbon is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_or \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
* **molar mass** = mass in grams per mole
* can be found on the periodic table

6

C

Carbon

12.0

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Practice**: Find the molar mass of   |  |  |  | | --- | --- | --- | | C = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ | Li = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ | O = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ | | Ca = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ | H = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ | S = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ | |

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Practice**: Find the molar mass of the following compounds.   |  |  | | --- | --- | | 1. carbon monoxide | 1. sodium sulfate | | 1. calcium nitrate | 1. magnesium acetate | |

|  |
| --- |
| **5.3 - Molar Mass as a Conversion Factor** (P. 82 #8-10, P.84 #18-19) |

We can use molar mass to convert mass to moles, and vice versa.

|  |
| --- |
| **Practice**: Convert mass to moles and moles to mass.   1. How many moles of CaCO3 are there in 5.6 g? 2. How many moles of NaCH3COO are there in 11.2 g? 3. How many grams of NaCl are there in 9.1 moles? 4. How many grams of H2O are there in 9.1 moles? 5. If 0.140 mol of acetylene gas has a mass of 3.64 g, what is the molar mass of acetylene? |

|  |
| --- |
| **5.4 - Multiple Conversions** (P. 86 #22) |

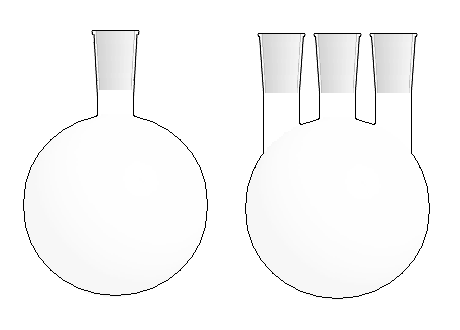
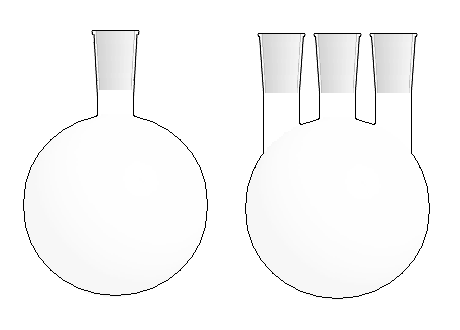
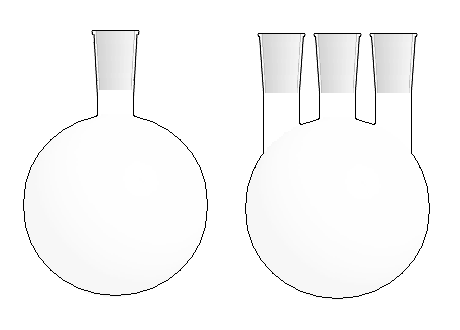
Using more than one conversion factor, Avogadro’s number and molar mass, we can find the number of molecules and atoms in a given amount of substance.

|  |
| --- |
| **Practice**: Use Avogadro’s number and molar mass to answer the following.   1. How many molecules of CuSO4 are there in 40.1 g of CuSO4? 2. How many molecules of MgCl2 are there in 3.55 g of MgCl2? 3. How many oxygen atoms are there in 14.1 g of Na2SO4? 4. How many Cl atoms are there in 0.55 g of AlCl3? 5. What is the mass of 5.99 x 1020 molecules of CuSO4? |

|  |
| --- |
| **5.5 - Gases at STP** (P. 83 #11-12, P. 84 #15-17, P.86 #22-24) |

How do we find the mass of gases?

**Avogadro’s hypothesis**: equal volumes of different gases, at the same temperature and pressure, contain the same number of particles

**Standard temperature and pressure (STP):** 0°C and 101.3 kPa

In other words: equal numbers of moles of any gas at STP will have the same volume

**Molar volume:** the volume of 1 mole of any gas is equal to \_\_\_\_\_\_\_\_\_\_\_ at STP

|  |
| --- |
| **Practice**: Find the volume or number of moles of the following gases.   1. What is the volume occupied by 0.350 mol of SO2(g) at STP? 2. How many moles of gas are there inside a balloon with volume 10.0 L at STP? 3. How many molecules of O2 are there inside an empty 250 mL water bottle at STP? |

|  |
| --- |
| **5.6 - Density and Gases** (P. 88 #25-34) |

We can find the density of gases too. Remember that

|  |
| --- |
| **Practice**: Find the density of the following gases.   1. What is the density of O2(g) at STP? 2. What is the density of C2H6(g) at STP? |

Similarly, we can use density to find volume or moles of gas.

|  |
| --- |
| **Practice**: Use density to find the volume or moles of gas.   1. What is the volume of a 3.00 mol sample of ethanol gas CH3CH2OH (density = 0.790 g/mL)? 2. How many moles of CO2 are contained in 100.0 mL (density = 0.00198 g/mL)? 3. What is the volume of 5.00 x 1023 atoms of silver (density = 10.49 g/mL)? |
| PART 2  Formula Determination |

|  |
| --- |
| **5.7 - Percent Composition** (P.91 #44-45) |

**Percent composition:** the percent by mass of a species in a compound

Steps:

1. Assume there is 1 mole of the given compound
2. Find the compound’s molar mass
3. Find the total mass of each species (element) in the compound
4. Divide the total mass of each species by the molar mass of the compound and multiply by 100

|  |
| --- |
| **Practice**: Find the percent composition of the following compounds.   1. Find the percent composition of CH4 2. Find the percent composition of H2SO4 3. What is the percentage of water in CuSO4 ∙ 5H2O |
| **5.8 - Empirical Formulas** (P.93 #46) |

**Empirical formula:** the smallest whole number ratio of atoms in a compound

Use percent composition to find the empirical formula.

1. Assume you have 100 g of a compound
2. Find the mass of each element in the compound
3. Convert these masses to moles
4. Divide the number of moles of each element by the smallest number of moles found to obtain a ratio

|  |
| --- |
| **Practice**: Find the empirical formula.   1. What is the empirical formula of a compound consisting of 80.0% C and 20.0% H? 2. A compound of nitrogen and oxygen is analysed. 2.000 g of nitrogen and 4.57 g of oxygen are found. What is its empirical formula? 3. A compound contains 58.8% C, 7.3% H, and 34.1% N. What is the empirical formula of the compound? |

Sometimes you will get decimals for the ratio.

* change these to whole numbers since we can’t have less than a whole atom

|  |  |  |
| --- | --- | --- |
| Decimal | Fraction Equivalent | Multiply by |
| 0.5 |  |  |
| 0.25 |  |  |
| 0.75 |  |  |
| 0.33 |  |  |
| 0.67 |  |  |

|  |
| --- |
| **Practice**: Find the empirical formula.   1. What is the empirical formula of a compound containing 81.8% C and 18.2% H? 2. A sample containing N and O was found to contain 25.99% N and 74.1% O. Find the empirical formula. |

Recall Avogadro’s hypothesis:

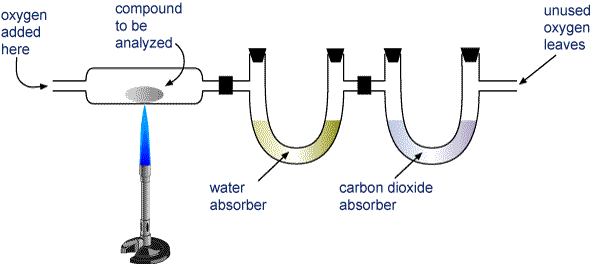
* equal volumes of different gases, at the same temperature and pressure, contain the same number of particles

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| **Example**: An experiment was done to test Avogadro’s hypothesis:  Nitrogen and oxygen gas are combined to form three different covalent compounds. Apply Avogardo’s Hypothesis and the concept of relative mass and empirical formulas to determine the chemical formulas for the following compounds.  http://upload.wikimedia.org/wikipedia/commons/b/be/Round-bottom_flasks.PNG http://upload.wikimedia.org/wikipedia/commons/b/be/Round-bottom_flasks.PNG http://upload.wikimedia.org/wikipedia/commons/b/be/Round-bottom_flasks.PNG   |  |  |  |  |  | | --- | --- | --- | --- | --- | |  | Volume of Nitrogen  (mL) | Volume of Oxygen  (mL) | Whole Number Ratio of N:O Atoms | Formula | | Compound 1 | 100 | 49.5 |  |  | | Compound 2 | 100 | 102.9 |  |  | | Compound 3 | 100 | 201.7 |  |  | |

|  |
| --- |
| **5.9 - Combustion Analysis** |

For compounds containing C, H, and O, the percent composition can be determined by combustion analysis

* burn the compound to produce \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
* find the % of C, H, and O from analyzing CO2 and H2O, and other products



|  |
| --- |
| **Practice**: Find the empirical formula.   1. A 1.00 g C-H sample was analyzed and found to produce 5.72 g CO2 and 4.69 g H2O. Find the empirical formula. 2. A 18.21 g C, H, O sample was found to produce 39.64 g CO2 and 14.62 g H2O. Find the empirical formula. 3. A 20.0 g sample was analyzed and found to produce 29.32 g CO2 and 11.95 g H2O. If the sample was made of C, H, and O, find the empirical formula. |

|  |
| --- |
| **5.10 - Molecular Formulas** (P.95 #47-55) |

|  |
| --- |
| **Consider This:**   1. Find the percent composition of CH2 and C2H4. 2. What do you notice about the percent composition for these two compounds? 3. If a scientist does percent composition analysis on the compound C3H6 and uses the data to calculate the empirical formula, will the empirical formula represent the formula of the actual compound at hand? What formula would the scientist derive? 4. Does the empirical formula represent the formula of the actual compound analyzed? |

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Practice**: What is the empirical formula for the following group of compounds?   |  |  |  |  |  | | --- | --- | --- | --- | --- | | C5H10 | C2H4 | CH2 | C3H6 | C4H8 | |

**Molecular formula (or chemical formula):** the actual chemical formula for a compound

* a whole number multiple of the empirical formula

|  |
| --- |
| **Example**: What is the molecular formula of a CH2 compound when N=3? |

To find N, we can rewrite the above equation:

|  |
| --- |
| **Practice**: Find the molecular formula.   1. A molecule has an empirical formula for HO and a molar mass of 34.0 g/mol. What is the molecular formula? 2. The empirical formula of a compound is SiH3. If 0.0275 mol of a compound has a mass of 1.71 g, what is the compound’s molecular formula? 3. A gas has the empirical formula POF3. If 0.350 L of the gas at STP has a mass of 1.62 g, what is the molecular formula for the compound? 4. A 1.00 g of propene, a C-H compound, was analyzed and found to produce 3.14 g CO2 and 1.24 g H2O. Find the empirical formula and the molecular formula if the molar mass of propene is 42.1 g/mol. |