**Unit 5: Chapter 10**

 **The Mole**

**Mole Explanation**

1. The mole is the central unit of measurement in chemistry; it is quantity like pair (2), dozen (12), case (24), and ream (500).
2. Like there are 12 items in a dozen, there are a specific number of particles in a mole.
	1. For elements, it is the number of atoms
	2. For covalent compounds it is the number of molecules
	3. For ionic compounds it is the number of formula units
3. The number is called Avagadro’s number and it is 6.02 x 1023 particles
	1. For instance there is 6.02 x 1023 atoms in one mole of Carbon

**Conversion from Particles to Moles and Moles to Particles**

 It’s pretty simply really if you want particles you multiply by 6.02 x 1023

If you desire to have the number of moles when you have particles you merely need to divide by 6.02 x 1023

**Steps for completing Mole problems**

1. Known
2. Unknown
3. Conversion
4. Problem Set-up
5. Answer (sig. fig, label)

 **Ex:** How many atoms are in 1.50 moles of Li?

1. 1.50 moles Li
2. Number of atoms of Li
3. 1mole/ 6.02 x 10 23 atoms

|  |  |  |
| --- | --- | --- |
| 1.50 moles Li | 6.02 x 1023 atoms Li | = |
|  | 1 mole Li |

= 9.03 x 1023 atoms of Li

 **Ex:** How many moles are in 15.06 x 1023 molecules of CH4?

1. 15.06 x 1023 molecules of CH4
2. Number of moles of CH4
3. 1mole/ 6.02 x 10 23 molecules of CH4

|  |  |  |
| --- | --- | --- |
| 15.06 x 1023 molecules CH4 | 1 mole CH4 | = |
|  | 6.02 x 1023 molecules CH4 |

= 2.501 moles CH4

**Practice:**

How many moles are there in 8.43 x 1023 formula units of KOH?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

How many moles are in 4.0x 1023 molecules of PBr3?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

How many formula units are in .033 moles of NaCl?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

How many moles are in 3.79 x 1023 atoms of Au?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

**Molar Mass**

1. **Molar mass** = the mass of one mole of a substance
	* Atomic mass for an element
	* Formula mass for an ionic compound
	* Molecular mass for covalent compounds
2. The mole has been defined in such a way that one mole of an element has a mass that is equal to it’s atomic mass.
3. *Finding molar mass for Elements*

Simply look on the periodic table and find the atomic mass; this is also the molar mass.

Ex: C, 1 mole of C has a mass of 12.0 g/1mole

Now you try. What is the atomic mass of Na?

1. *Finding molar mass for Compounds*

This is very similar to elements the only difference is you have to add up different atomic masses to find the total.

**Ex**: CCl4

 C = 1 x 12.0 g = 12.0g

 Cl = 4 x 35.5 g = 142.0g

154.0 g

 Carbon tetrachloride has a molar mass of 154.0/1mole CCl4

**Ex:** Ca(C2H3O2)2

Ca = 1 x 40.1g = 40.1g

 C = 4 x 12.0g = 48.0g

 H = 6 x 1.0g = 6.0g

 O = 4 x 16.0g = 64.0g

 158.1g/1mole Ca(C2H3O2)2

**Practice**

1. C4H10O
2. Al(OH)3
3. (NH4)3PO4

**Conversion from Mass to Moles and Moles to Mass**

All you have to do here is

* Find the molar mass of the compound in question and set up the problem correctly.
* You start with your known
* Now knowing that 1 mole is equal to the molar mass of your compound you set up the fraction  so the units from your known cancel out.
* Make sure your answer is in the units of your unknown and you are done.

**Steps for completing Mole problems**

* 1. Known
	2. Unknown
	3. Conversion
	4. Problem Set-up
	5. Answer (sig. fig, label)

**Completed Examples**

*How many moles are in 10.0g of Na?*

1. 10.0g Na
2. Moles of Na
3. 1 mole/23.0grams Na

|  |  |  |
| --- | --- | --- |
| 10.0g Na | 1 mole Na | = |
|  | 23.0g Na |

4

 5. 5. 0.435 moles Na

*How many grams are in 15.4 moles of CaCO3?*

1. 15.4 moles of *CaCO3*
2. Mass (grams) of *CaCO3*
3. 1mole/100.1 g *CaCO3*

|  |  |  |
| --- | --- | --- |
| 15.4 mole CaCO3 | 100.1 g CaCO3 | = |
|  | 1 mole CaCO3 |

#  = 1540g CaCO3

# Practice Problems

How many moles are in 16.3g of LiCl?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

How many grams are in 4.09 moles of C?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

How many moles are in .065g of Mg(NO3)2?

|  |  |  |
| --- | --- | --- |
|  |  | = |
|  |  |

**Molar Volume: Converting from liters to moles and moles to liters**

There are some substances that are difficult to mass out, they are gases and it is much easier to deal with them using volume rather than mass. Fortunately, there is a way in which we can go from moles to volume.

*History time*

 A long time ago there was a man named Gay-Lussac and he did several experiments that involved gases. He found that when certain gasses combined and formed compounds they would no longer occupy the same volume of space. They would change by very specific whole number ratios.

 He could not explain this phenomenon so he did what any good chemist would do and he repeated his experiments took meticulous data and then reported his findings.

 A young gentleman by the name of Amedeo Avagadro came across his data and then made the startling conclusion that…

### *Equal volumes of gas contain equal number of particles.*

This means that 100 liters of H2 would have the same number of molecules as 100 liters of O2 even though the molecules are very different in size.

 This explained why when Gay-Lussac combined 2 liters of H2 with 1 liter of O2 he ended up with 2 liters of H2O vapor! Normally you would expect that when you combine 2 liters with 1 liter you would get 3 liters. But the atoms combined

2 H2 + O2 🡪 2 H2O

The number of H2O molecules in 2 liters had to be the same number of H2 molecules in 2 liters of H­2, and twice as many molecules as there are in 1 liter of O2. And if you look you see that there are 4 Hydrogen atoms and 2 Oxygen on each side. It is the number of molecules that adds up properly not the volumes.

### So once we new that all we had to do is find the volume 6.02 X 1023 particles of a given gas occupy and that would be the same volume for 6.02 x 1023 particles (1 mole) of all gases.

### In order to make this consistent we had to choose a specific pressure and temperature. (If you remember back to physical science volumes of a gas change with the temperature and pressure)

### We decided that *S*tandard *T*emperature and *P*ressure would be a good choice. “What is Standard Temperature and pressure?” you ask, I’m glad you asked…

Standard Temperature = 00 C

Standard Pressure = 1 atmosphere (the pressure of the atmosphere at sea level)

### So we did the tests at S.T. & P. and found that 6.02 X 1023 particles of a given gas occupy 22.4 liters.

And now whenever we want to convert from moles to liters we know that a mole of any gas at **STP**takes up 22.4 liters and we can multiply or divide as we choose.

### Examples:

* 1. 11.2 L of a gas are collected at STP. How many moles of gas are present?
1. 11.2 Liters of a gas
2. Number of moles
3. 1mole / 22.4 liters

|  |  |  |
| --- | --- | --- |
| 11.2 liters | 1 mole | = 2.00 moles |
|  | 22.4 liters |

* 1. What volume would 5.47 moles of gas occupy at STP?
	2. How many moles would be in 8.500L of ammonia (NH3) at STP?

**Combination Problems-moles, particles, mass, and volume.**

**The MOLE BOX:**

**1 mole 6.02 x 1023 particles**

 **Mass Volume**

 **(# of grams) 22.4 liters**

1. How many grams on dry ice (CO2) will occupy a volume of 3.50 Liters at STP?

 b. How many molecules of dry ice (CO2) are in this same sample?

**Percent Composition**

1. Percent composition is the percentage by mass of each element in a compound
2. Can be determined through experiment and used to determine the compound’s formula.
3. If the formula of the compound is known then the percent composition can be calculated.
4. To calculate

|  |
| --- |
| X 100% = % Composition |

Mass of an element in sample

 Mass of the total sample

1. Two main types of problems
	* Where the **Masses** are given

Ex: A sample of a compound containing carbon and oxygen had a mass of 88.0 gr. If 24.0 gr. of the sample was C, what is the % composition?

|  |  |  |
| --- | --- | --- |
| Carbon = | 24.0g | X 100 = 27.3 % C |
| 88.0g |

|  |  |  |
| --- | --- | --- |
| Oxygen = | 64.0g | X 100 = 72.7 % 0 |
| 88.0g |

* + Where you are given the **formula**

Ex: Find the % composition of Na2O First find the molar mass

First find the molar mass : Na = 2 x 23.0g = 46.0g

 O = 1 x 16.0g = 16.0g

 62.0 Na2O

Next you need to find the % of each element based on the molar mass

|  |  |  |
| --- | --- | --- |
| Sodium = | 46.0 | X 100 = 74.2 % Na |
| 62.0g |

|  |  |  |
| --- | --- | --- |
| Oxygen = | 16.0g | X 100 = 25.8 % O |
| 62.0g |

**Practice Problem:**

* What is the % composition of Hydrogen in 1 mole of water?

**Empirical Formula**

1. The empirical formula is the simplest ratio of atoms in a chemical compound.
2. If the compound can be simplified by dividing you do

Ex:Butyric acid has a molecular formula of HC4H7O2, its empirical formula is C2H4O

Ex: Benzene, C6H6 has the empirical formula of CH.

1. The empirical formula can be determined from the atomic masses and the percent composition of the compound.
	1. Step 1: **Change % to grams**. Assume 100 g.
	2. Step 2: Calculate the moles of each element in the compound from the given data: **Change grams to moles**.
	3. Step 3: **Divide** the number of moles of each element found in Step 1 **by the smallest number of moles**, to obtain whole-number subscripts for the empirical formula
	4. If Step 4: does not give **whole numbers** (within ± 0.1), the decimal portion of the number should be close to fraction ( 0.5 = ½, 0.33 = 1/3 0.67 = 2/3 and so on) If this is the case then simply multiply each number by the denominator.

**Ex:** A compound contains carbon, hydrogen, and oxygen is found to contain 9.10 % hydrogen and 54.5% carbon what its empirical formula? (When given % simply assume a 100g sample and treat them as masses)

1. 9.10g H, 54.5g C, 36.4g O

|  |  |  |
| --- | --- | --- |
| 9.10 g H | 1 mole | = 9.10 moles H |
| 1 | 1.0g H |

|  |  |  |
| --- | --- | --- |
| 54.5 g C | 1 mole | = 4.54 moles C |
| 1 | 12.0 g C |

|  |  |  |
| --- | --- | --- |
| 36.4 g O | 1 mole | = 2.28 moles O |
| 1 | 16.0 g  |

|  |  |  |
| --- | --- | --- |
| 9.10 moles | = |  H’s subscript 4 |
| 2.28 moles |

|  |  |  |
| --- | --- | --- |
| 4.54 moles | = |  C’s subscript 2 |
| 2.28 moles |

|  |  |  |
| --- | --- | --- |
| 2.28 moles | = |  O’s subscript 1 |
| 2.28 moles |

**H4C2O**

**Ex:** What is the empirical formula of a compound that contains 36% calcium and 64% chlorine?

**Molecular Formula**

1. The molecular formula is the actual formula of a given compound.
2. For ionic compounds the empirical formula is the molecular formula so if you have the empirical formula you are done yeah!
3. For molecular compounds, hence molecular formula, it may be determined by creating a whole number multiple of the empirical formula
4. The molecular mass of the compound has to be known
5. Then you simply divide the molecular mass by the mass of the empirical formula, you should get a small whole number.
6. Each subscript is then multiplied by the small number you obtain.

**Ex:**

A compound has the empirical formula CH2O, and its molecular mass is 180.00 grams. What is the molecular formula of this compound?

First find the empirical mass

 C = 1 x 12.0g = 12.0g

H = 2 x 1.0g = 2.0g

O = 1 x 16.0g = 16.0g

 30.0g

Next simply divide the molecular mass by the empirical mass

|  |  |
| --- | --- |
| 180.0g | = 6  |
| 30.0g |

Now all you have to do is multiply the subscripts by that number

(C1H2O1) x 6 = C6H12O6

(I added the subscripted ones to help make the connection with the multiplication; they are usually not there.)

**Now you try:**

The empirical formula is C3H7 and the molecular mass is 86.0 grams.

The empirical formula is CH2ON and the molecular mass is 176.0 grams.