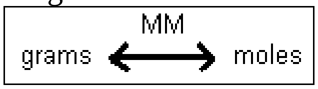
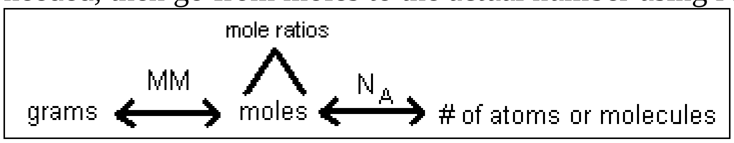
Stoichiometry

I. Stoichiometry,

A. To go between moles and grams we use molar mass (MM) as a conversion factor.



B. Putting it all together we can go from grams to moles using MM, then have a mole ratio if needed, then go from moles to the actual number using N_A.

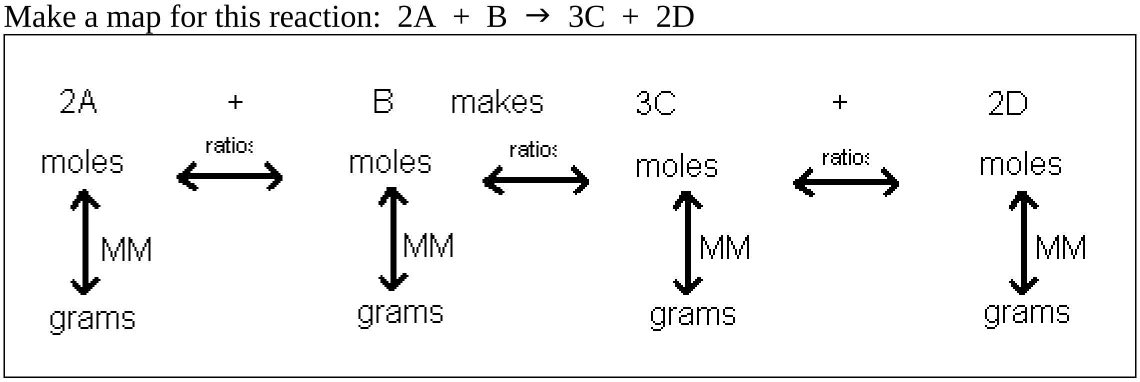


C. Potential questions (note these are all **theoretical** calculations and they are all for chemicals just sitting there, not involved in a reaction):

1. How many H atoms in 3.50 grams of ammonia?
2. How many grams would 3.55×10^{25} molecules of sulfur tetrachloride mass?
3. How many phosphate ions are in 0.444 grams of calcium phosphate?
4. actual examples all grouped together below

D. What about calculations involving chemicals in a reaction? Remember a reaction is balanced to tell us the ratio of moles, not mass. Consider the formation of steam: $2H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$. The reactions says 2 molecules of hydrogen gas react with one molecule of oxygen gas to make 2 molecules of steam. Now we can multiply them all by 6.02×10^{23} to get 2 moles of hydrogen gas react with one mole of oxygen gas to make 2 moles of steam. Balanced reactions tell us molecular or mole ratios, NOT mass ratios. NOT 2 grams plus one gram makes 2 grams!

E. Make a map for this reaction: $2A + B \rightarrow 3C + 2D$



The map shows us that we can NOT compare grams to grams directly. We must always convert to moles first, then use mole ratios, then go back to grams. I often call these "grams to moles to moles to grams" problems and they are SUPER important. (note these are all **theoretical** calculations also) **Don't forget that we can also go to the # using Avogadro's number once we have moles – I just didn't put that additional connection in the picture.**

F. Potential questions:

1. How many grams of steam could be made using 15.0 grams of oxygen gas?
2. How many molecules of steam could be made starting with 35.0 grams of hydrogen gas?
3. practice probs below

II. % Yield,

A. All our calculations so far are theoretical (we are making them up based on a starting variable). The answers assume the reaction was perfect – no spills, every last molecule came out of the beaker and reacted, there is no waste or by-products or secondary side reactions. Of course this is not true. We can never have 100% yield, it must be less.

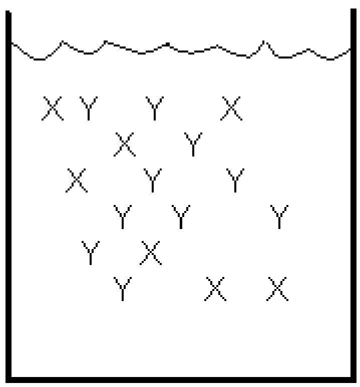
B. % yield = $\left(\dfrac{actual}{theoretical}\right)$ x 100%

III. Limiting Reagents,

A. In reality we must consider both reactants not just one when we run a reaction. In lab we need to know how much of all reactants to add together to make the product(s).

B. Here is an analogy: how many 3 legged bar stools can I make if I have 9 seats and 25 legs? Well 9 seats (1 stool / 1 seat) = could make 9 stools but 24 legs (1 stool / 3 legs) can only make 8 stools. So the real answer is 8 stools with one seat leftover. The seats are in excess and the legs are the limiting reagent.

C. If the product molecule is XY_3 how many can I make given this mixture in a beaker? Pretend X represents one mole of S and Y represents one mole of Y.



Well I have 7 X and 10 Y so

7 X (1 XY_3 / 1 X) = 7 XY_3

and

10 Y (1 XY_3 / 3 Y) = 3.33 XY_3

So I can really make just 3.33 moles of product. Y is the limiting reagent and X is in excess.

IV. Molarity,

A. Most reactions occur in solution – we can't just rub solids together to make them react! So we dissolve solids in liquids, then pour them together. **A solution is defined as a solute dissolved in a solvent.** Typically the solute is a solid or gas and the solvent is water.

B. Molarity is one way to measure concentration and it is defined as moles of solute per liters of solution. Units = mol / L. **M = mol / L.**

C. Molarity M is actually a conversion to go between volume and moles.

$$\boxed{\text{Volume} \xleftrightarrow{\;M\;} \text{moles}}$$

D. Always remember molarity is a ratio. Which is more concentrated? A cup of coffee or the pot it came from? Well the pot may have more volume, but the concentration is the same.

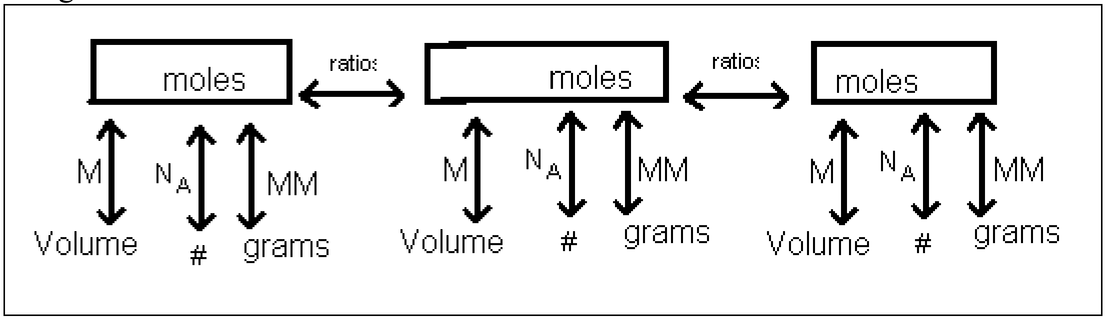
V. Dilutions,

A. To dilute something we add more solvent, usually water. Consider a beaker containing 4 marbles in one liter of water. The concentration is 4 marbles / L. Now add 2 more liters of water to the beaker and we have a concentration of 4 marbles / 3 L = 1.33 marbles / L. Note the # of marbles was constant. Try this – multiply the volume times the liters for each: 1) 4 marbles per liter x 1 liter = 4 marbles and 2) 1.33 marbles per liter x 3 liters = 4. Hum, relating marbles to moles and **molarity times volume = moles!!! ML = mol**

B. Dilution equation: $M_1V_1 = M_2V_2$ BUT remember that V_2 is the total volume. V_1 plus the V added = V_2!!!

VI. **Solution Stoichiometry,**

A. Now apply what we have learned to solutions since most reactions take place in solution. Here is a general map. Remember we have to use moles to use the mole ratios. But from moles we can find grams via molar mass, we can find volume via molarity, and we can find the actual # using Avogadro's number.



B. For example we can start with mL of a reactant but calculate the theoretical grams of a product. Convert mL to moles using the molarity, then use a mole ratio, then convert to grams using molar mass. The possibilities are endless.

VII. **Titrations,**

A. A titration is when we add two solutions together (two reactants dissolved in a solvent) and we add them slowly for the purpose of stopping when exactly enough has been added of each so that there is no reactant in excess. Why? So that there is no waste, nothing is left over, every reactant molecule has reacted. Sounds good but it is time consuming and tedious.

B. Titrations can be used for many types of reactions: acid/base, redox, precipitations as long as both reactants are in solution (aq). The calculations learned above apply. Practice problems below.

VIII. **% Composition,**

A. How do we find a molecular formula if the compound is unknown? Well

1. First analyze the unknown in a machine to find the % of all the elements involved.
2. Pretend you have 100 grams of the sample, so then you know the grams of each element involved and can convert to moles.
3. Get whole number ratios of the elements = empirical formula (kinds like finding a common denominator)
4. Compare the empirical molar mass to the real molar mass to determine how many empirical formulas are in the real formula, and thus get the real molecular formula.

B. Practice problems below.

Practice Problems

1. Calculate how many molecules of NaCl are in 12.5 g?
2. What is the mass of 0.0295 moles of carbon monoxide?
3. If 8.39 moles of water are consumed in the following reaction, how many moles of hydrogen gas are evolved? $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
4. How many grams of propane (C_3H_8) must be completely combusted to produce 12.5 grams water?
5. A solution is made by adding 25.0 mL of water to 10.0 mL of a 7.57M solution. Find the final molarity.
6. Ethylene glycol, a substance used in antifreeze (get some in your car soon!) is composed of 38.7% carbon, 51.6 % oxygen, and 9.7% hydrogen. What is the empirical formula?
7. How many grams of water could be produced if 5.00g butane (C_4H_{10}) were combusted with 5.00g of oxygen?

8. Spock, aboard the starship enterprise, reacted 3.00 g lithium hydroxide with an excess of carbon dioxide and produced 3.00 g of lithium carbonate. Find his percent yield. $2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(l)$
9. How many grams of KNO_3 would be needed to prepare 25.00mL of a 0.750M solution?
10. How many mL of a 1.55 M KOH solution are needed to completely neutralize 8.50 mL of a 2.00 M H_3PO_4 solution?
11. A sample is found to contain 92.3% C and 7.69% H. What is the empirical formula? If the molecular mass is 78 g/mol, what is the molecular formula?
12. If I have 10.0 g of ethane C_2H_6 and combust with 5.00 g oxygen gas, how much carbon dioxide can I produce?

Answers

1. 12.5 g NaCl (1 mole / 58.45 g)(6.02 x 10^{23} molecules / 1 mole) = **1.29 x 10^{23} NaCl molecules**
2. 0.0295 moles CO (28 g/mol) = 0.826 g
3. 8.39 mol water (1 mol H_2 / 2 mol water) = 4.20 moles
4. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$ 12.5 g water (mol/18g)(1 propane / 4 water)(44 g / mol) = 7.64 g
5. 10 mL (7.57 M) = M (35 mL) and M = 2.16
6. 38.7 g C (mol/12g) = 3.225 / 3.225 = 1 C, 51.6 g O (mol/16g) = 3.225 / 3.225 = 1 O, 9.7 g H (mol/1g) = 9.7 / 3.225 = 3 H, So CH_3O
7. 5.00 g butane (mol / 58g)(10 H_2O / 2 butane)(18 g / mol) = 7.76 g water OR 5.00 g O_2 (mol / 32g)(10 H_2O / 13 O_2)(18 g / mol) = **2.16 g water** ***** this is the correct answer
8. 3 g LiOH (mol/23.94 g)(1 Li_2CO_3 / 2 LiOH)(73.88 g / mol) = 4.63 g and 3.00 g / 4.63 g x 100% = 65%
9. (0.75 mol / L)(0.025 L)(101.09 g / mol) = 1.90 g
10. 32.9 mL
11. Assume a 100g sample. Thus % are g. 92.3 g C (mol / 12 g) = 7.69 mol C, 7.69 g H (mol / 1 g) = 7.69 mol H. Now divide each of the mole amounts by the smallest amount which is 7.69. Result - 1 C and 1 H thus empirical formula is CH. Now the empirical mass of CH is 13 g/mol but the real molecular mass is 78 g/mol. 78/13 = 6. Thus the real formula is 6 times the empirical which is C_6H_6.
12. Let's do ethane first. 10.0 g C_2H_6 (mol / 30 g)(4CO_2 / 2C_2H_6)(44 g / mol) = 29.3 g CO_2 Now let's do oxygen gas. 5.00 g O_2 (mol / 32g)(4CO_2 / 7O_2)(44 g/mol) = 3.93 g CO_2 Now which is the real amount? Using ethane I can make 29.3 g but using oxygen I can only make 3.93 g. How can I make 29.3 g when I'll run out of oxygen gas? I can't. I can only make 3.93 g carbon dioxide. Oxygen gas is the limiting reagent and ethane will be leftover in excess.

Use this reaction for the following questions: $Al_2S_3(s) + H_2O(l)$ ---> $Al(OH)_3(s) + H_2S(g)$

1. Balance the reaction.
2. Calculate the molecular weights for all species in the chemical reaction.
3. If I have 5.00 g of aluminum sulfide to react with excess water, how many grams of aluminum hydroxide can I produce?
4. If 4.00 moles of water are used in the above reaction, how many grams of aluminum sulfide was reacted?
5. If 3.00 moles of hydrogen sulfide is produced, how many moles of water were reacted?
6. If 3.00 grams of hydrogen sulfide is produced, how many grams of water were reacted?
7. State the Law of Conservation of Mass..

8. If I start the reaction above with a total mass of the reactants equaling 10.00 grams, what will the final mass be after the reaction?

Answers

1. $Al_2S_3(s) + 6\ H_2O(l) \longrightarrow 2\ Al(OH)_3(s) + 3\ H_2S(g)$
2. $Al_2S_3(s)$ = 150.2 g/mol, $H_2O(l)$ = 18.0 g/mol, $Al(OH)_3(s)$ = 77.98 g/mol, $H_2S(g)$ = 34.1 g/mol
3. 5.00 g Al_2S_3 (1 mol / 150.2 g)(2 mol $Al(OH)_3$ / 1 mol Al_2S_3)(77.98 g / 1 mol) = 5.19 g $Al(OH)_3$
4. 4.00 mol H_2O (1 mol Al_2S_3 / 6 mol H_2O)(150.2 g / 1 mol) = 1.00 x 10^2 g Al_2S_3
5. 3.00 mol H_2S (6 mol H_2O / 3 mol H_2S) = 6.00 mol H_2O
6. 3.00 g H_2S (mol / 34.1 g)(6 mol H_2O / 3 mol H_2S)(18.0 g / mol) = 3.17 g H_2O
7. Matter can neither be created nor destroyed.
8. The final mass must be 10.00 grams since I am not God and can not create matter nor destroy it. Even if all the reactants do not react to form products (thus they are left over) the total mass after reaction must equal the mass before reaction.

There are more here on the help page
http://web.gccaz.edu/~ksmith8/help.htm
on molarity, dilution, Avagadro's number and stoichiometry

Fun Stuff! Study Hard.