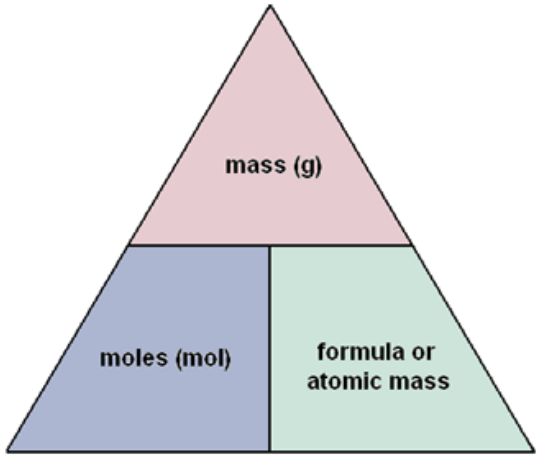
20 Key Calculations

1. Relative Formula Mass
2. Concentration (g/dm^3)
3. Moles to Particles
4. Moles Triangle
5. Concentration (mol/dm^3)
6. Converting Concentration
7. Titration Calculation
8. Empirical Formula from Molecular Formula
9. Molecular Formula from Empirical Formula
10. Empirical Formula from Reacting Masses
11. Empirical Formula from %composition
12. Conservation of Mass
13. Reacting Masses
14. Limiting Reagent
15. % Yield
16. Atom Economy
17. Gas Volume
18. Isotope Calculation
19. Bond Enthalpy

Equation Sheet – Combined Only

Moles Triangle



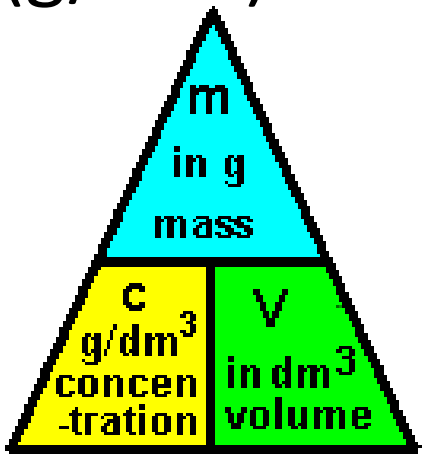
Empirical Formula

Divide mass by RAM and then compare the ratios

Reacting Masses

Convert given mass to moles. Then convert moles to mass of unknown

Concentration (g/dm^3)



Bond Enthalpy

Energy change = bonds broken – bonds made

$$\text{Relative Atomic Mass} = \frac{\%_1 \times \text{mass}_1 + \%_2 \times \text{mass}_2}{100}$$

Equation Sheet – Separate Chemistry Only

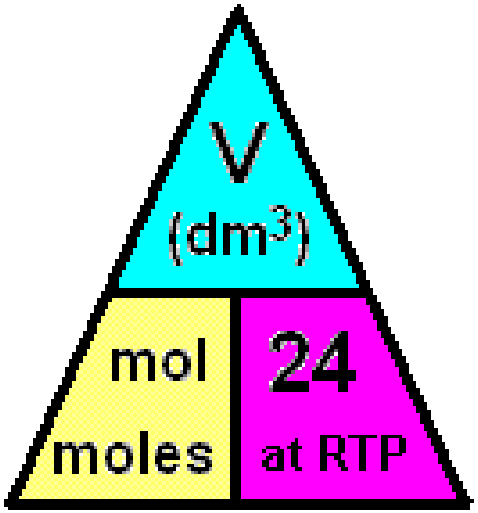
$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Titration Equation

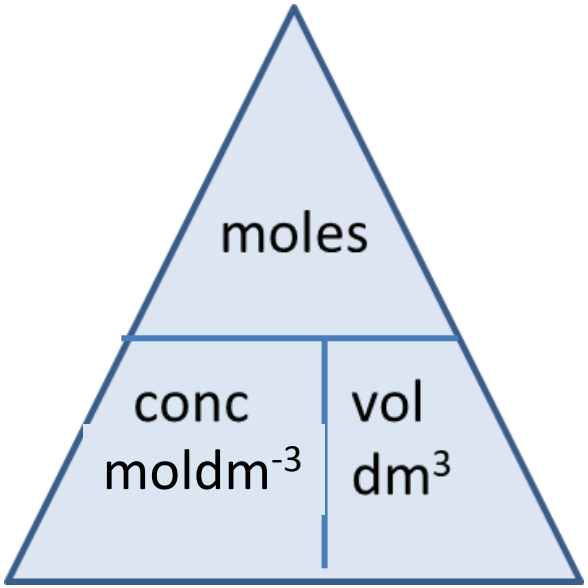
$C_A V_A = C_B V_B$

$$\%\ \text{atom economy} = \frac{\text{total RFM of desired product}}{\text{total RFM mass of all products}} \times 100\%$$

Gas Volume Triangle



Mole of solute=concentration x volume



1-Relative Formula Mass

The relative formula mass (RFM) is calculated by adding together the atomic masses of all the atoms shown in the formula.

Example: Calculate the relative formula mass of ammonia, NH_3. The relative atomic masses are: H=1.0 and N=14.0)

RFM =14.0+ (3x1.0)=17.0

You try:

Bronze: Calculate the relative formula mass of O_2 (The relative atomic mass of O=16.0)

Silver: Calculate the relative formula mass of $NaNO_3$ (The relative atomic mass of Na=23.0, N=14.0, O=16.0)

Gold: Calculate the relative formula mass of $Mg(OH)_2$ (The relative atomic mass of Mg = 24.3, O=16.0, H=1.0)

1-Relative Formula Mass - Answers

Bronze: Calculate the relative formula mass of O_2 (The relative atomic mass of O=16.0)

RFM= 2x16.0=32.0

Silver: Calculate the relative formula mass of $NaNO_3$ (The relative atomic mass of Na=23.1, N=14.0, O=16.0)

RFM= 23.0 + 14.0 + (3x16.0) =62.0

Gold: Calculate the relative formula mass of $Mg(OH)_2$ (The relative atomic mass of Mg = 24.3, O=16.0, H=1.0)

RFM = 24.3+ 2(16.0+1.0)=58.3

1-Calculating Relative Formula Mass

calculation using 1dp for Ar as required by A level

Q1. Calculate the relative formula mass of water, H_2O.

(Relative atomic masses: H = 1.0, O = 16) = (1.0x2) + 16.0 =18.0

Q2. Calculate the relative formula mass of iron chloride, $FeCl_3$.

(Relative formula masses: Cl = 35.5, Fe = 56) = 56 + (35.5x3) =162.5

Q3. Calculate the relative formula mass of $C_{16}H_{12}N_2O$. RAM H=1, C=12, N=14, O=16

= (12.0x16) + (1.0x12) + (14.0x2) + 16.0 = 248.0

Q4. Calculate the relative formula mass of calcium hydroxide, $Ca(OH)_2$.

(Relative atomic masses: Ca = 40; O = 16, H=1) = 40.1 + 2x(16.0+1) =74.1

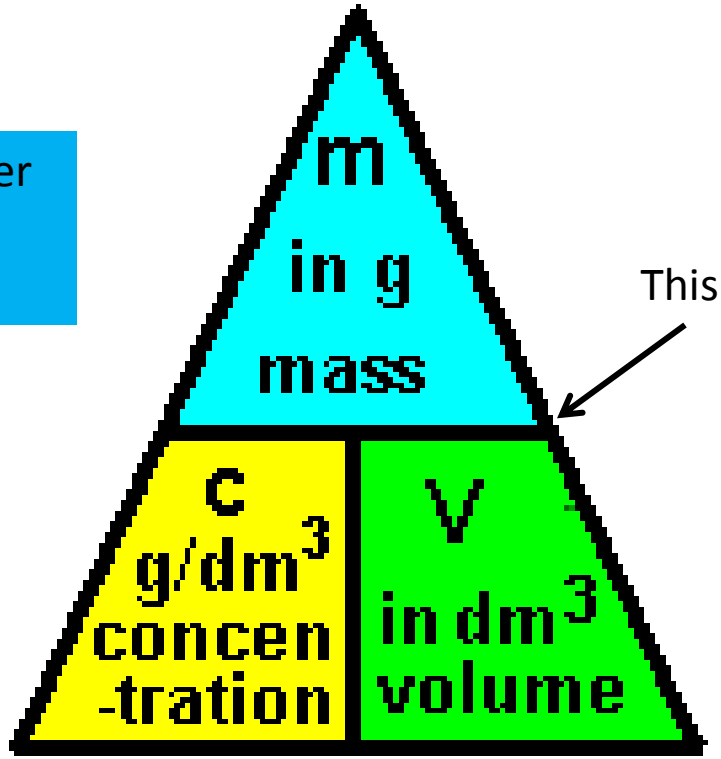
Q5. Calculate the relative formula mass of magnesium nitrate, $Mg(NO_3)_2$.

(relative atomic masses: Mg = 24, N = 14, O = 16) = 24.3 + 2x(14.0+16.0x3) =148.3

2-Calculating Concentration

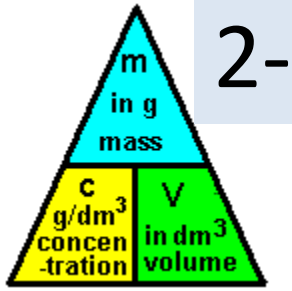
$$\text{Concentration} = \frac{\text{mass}}{\text{volume}}$$

To use this triangle, cover up the part you are looking for.



This line means divide

This line means multiply



2-Calculating Concentration

1000 cm^3 = 1 dm^3

$$\text{Concentration} = \frac{\text{mass}}{\text{volume}}$$

Example:

What is the concentration of a solution made from 25.0g of sodium hydroxide and 100.0 dm^3 of water?

$$\text{Concentration} = \frac{25.0}{100.0} = 0.250 g / dm^3$$

You try:

1. Calculate the concentration of a solution made from 10.0g of sodium hydroxide and 200.0 dm^3 of water.
2. What is the concentration of a solution made from 2.50g of sodium chloride and 12.0 dm^3 of water?
3. How many grams are needed to make 500.0 dm^3 of solution of potassium hydroxide with a concentration of 5.00 g/dm^3 ?
4. What volume of water is needed to make a solution 7.00g/dm^3 solution from 2.00g of sodium carbonate?
5. What is the concentration in g/dm^3 of a solution made from 0.900g of sugar and 25.0 cm^3 of water?

2-Calculating Concentration - Answers

You try:

1. Calculate the concentration of a solution made from 10.0g of sodium hydroxide and 200.0 dm^{-3} of water.

$$\text{Concentration} = \frac{10.0}{200.0} = 0.0500 g/dm^3$$

2. What is the concentration of a solution made from 2.50g of sodium chloride and 12.0 dm^3 of water?

$$\text{Concentration} = \frac{2.50}{12.0} = 0.210 g/dm^3$$

3. How many grams are needed to make 500.0 dm^3 of solution of potassium hydroxide with a concentration of 5.00g/dm^3 ?

$$\text{Mass} = \text{concentration} \times \text{volume} = 5.00 \times 500.0 = 2500\text{g}$$

4. What volume of water is needed to make a solution 7.00g/dm^3 solution from 2.00g of sodium carbonate?

$$\text{Volume} = \frac{\text{mass}}{\text{concentration}} = \frac{2.00}{7.00\text{g}} = 0.285\text{dm}^3$$

5. What is the concentration in g/dm^3 of a solution made from 9.00g of sugar and 25.0 cm^3 of water?

$$\text{Concentration} = \frac{9.00}{25.0} \text{x} 1000 = 360\text{g/dm}^3$$

3-Using Avogadro's Constant

You will be given this constant on your exam

Calculating the number of particles:

To calculate the number of particles, multiply the number of moles by 6.02×10^{23}.

How many particles?

1. 2 moles of carbon

$$\text{particles} = \text{moles} \times 6.02\text{x}10^{23} = 2 \times 6.02 \times 10^{23} = 1.204 \times 10^{24}$$

2. 0.04 moles of CO_2

$$\text{particles} = \text{moles} \times 6.02\text{x}10^{23} = 0.04 \times 6.02 \times 10^{23} = 2.408 \times 10^{22}$$

3. 0.5 moles of HCl

$$\text{particles} = \text{moles} \times 6.02\text{x}10^{23} = 0.5 \times 6.02 \times 10^{23} = 3.01 \times 10^{23}$$

3-Using Avogadro's Constant

You will be given this constant on your exam

Calculating the number of moles:

To calculate the number of moles, divide the number of particles by 6.02×10^{23}.

How many moles?

1. 1000 molecules of oxygen

$$\text{moles} = \frac{\text{particles}}{6.02\times10^{23}} = \frac{1000}{6.02\times10^{23}} = 1.66\times10^{-21}$$

2. 2,000,000 molecules of hydrogen

$$\text{moles} = \frac{\text{particles}}{6.02\times10^{23}} = \frac{2{,}000{,}000}{6.02\times10^{23}} = 3.32\times10^{-18}$$

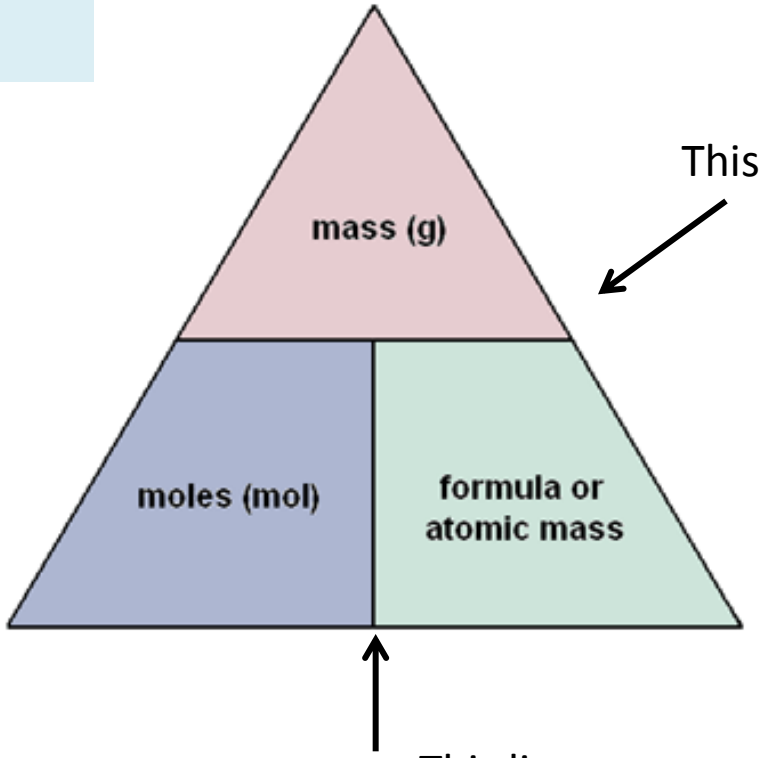
3. 3.00×10^{25} atoms of helium

$$\text{moles} = \frac{\text{particles}}{6.02\times10^{23}} = \frac{3.00\times10^{25}}{6.02\times10^{23}} = 498$$

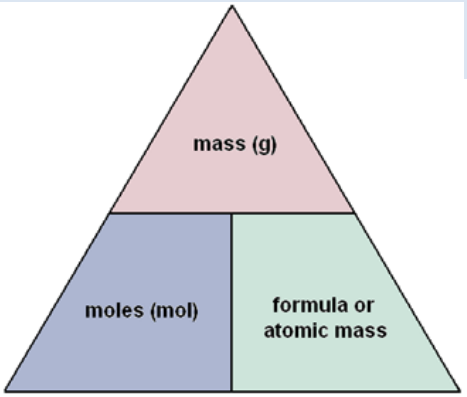
4-The Moles Triangle

$$\text{moles} = \frac{\text{mass}}{\text{RFM/RAM}}$$

To use this triangle, cover up the part you are looking for.



4-Using the Moles Triangle:



You have to learn this triangle

Bronze: How many moles in . . .?

1. 12.0g of Mg (RAM of Mg=24.3)
2. 2.00g of H_2 (RAM of H=1.0)
3. 51.0g of NH_3 (RAM of H=1.0, N=14.0)

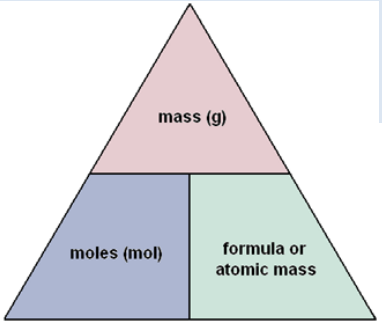
Silver: How many grams in . . .?

1. 1.00 mole of carbon (RAM of C=12.0)
2. 0.200moles of CO_2 (RAM of C=12.0, O=16.0)
3. 0.500 moles of HCl (RAM of H=1.0, Cl=35.5)

Gold: How many particles in . . .?

1. 3.00g of Mg (RAM of Mg =24.3)
2. 0.500g of Water, H_2O (RAM of H=1.0, O=16.0)

4-Using the Moles Triangle:



Bronze: How many moles in . . .?

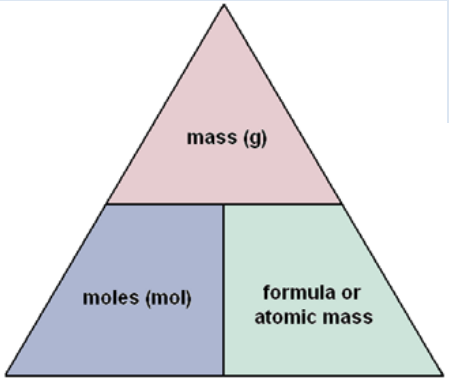
1. 12.0g of Mg (RAM of Mg=24.3) $\text{moles} = \frac{\text{mass}}{\text{RAM}} = \frac{12.0}{24.3} = 0.494$

2. 2.00g of H_2 (RAM of H=1.0) $\text{moles} = \frac{\text{mass}}{\text{RFM}} = \frac{2.00}{2.0} = 1.0$

3. 51.0g of NH_3 (RAM of H=1.0, N=14.0)

$$\text{moles} = \frac{\text{mass}}{\text{RFM}} = \frac{51.0}{17.0} = 3.00$$

4-Using the Moles Triangle:



Silver: How many grams in . . .?

1. 1.00 mole of carbon (RAM of C=12.0)

$$\text{mass} = \text{moles} \times \text{RAM} = 1.00 \times 12.0 = 12.0g$$

2. 0.200 moles of CO_2 (RAM of C=12.0, O=16.0)

$$\text{mass} = \text{moles} \times \text{RFM} = 0.200 \times 44.0 = 8.80g$$

3. 0.500 moles of HCl (RAM of H=1.0, Cl=35.5)

$$\text{mass} = \text{moles} \times \text{RFM} = 0.500 \times 36.5 = 18.3/18.25g$$

4-Using the Moles Triangle:

Gold: How many particles in . . .?

1. 3.00g of Mg (RAM of Mg =24.3)

Step 1: Calculate moles:

$$\text{moles} = \frac{\text{mass}}{\text{RAM}} = \frac{3.00}{24.3} = 0.123$$

Step 2: Use Avogadro's Constant to calculate particles

$$\text{particles} = \text{moles} \times 6.02\text{x}10^{23} = 0.125 \times 6.02 \times 10^{23} = 7.53 \times 10^{22}$$

2. 0.500g of Water, H_2O (RAM of H=1.0, O=16.0)

$$\text{moles} = \frac{\text{mass}}{\text{RFM}} = \frac{0.500}{18.0} = 0.0278$$

$$\text{particles} = \text{moles} \times 6.02\text{x}10^{23} = 0.0278 \times 6.02 \times 10^{23} = 1.67 \times 10^{22}$$

4-Exam Questions

Q3. 1.27 g of copper were produced in an experiment. Calculate the number of moles of copper, Cu, produced in this experiment. (Relative atomic mass: Cu = 63.5)

(1)

amount of copper produced = 0.0200 mol

Q2. Glucose has the formula $C_6H_{12}O_6$. Calculate the number of moles in a 0.250g sample.
(relative atomic masses: H=1.0, C=12.0, O=16.0)

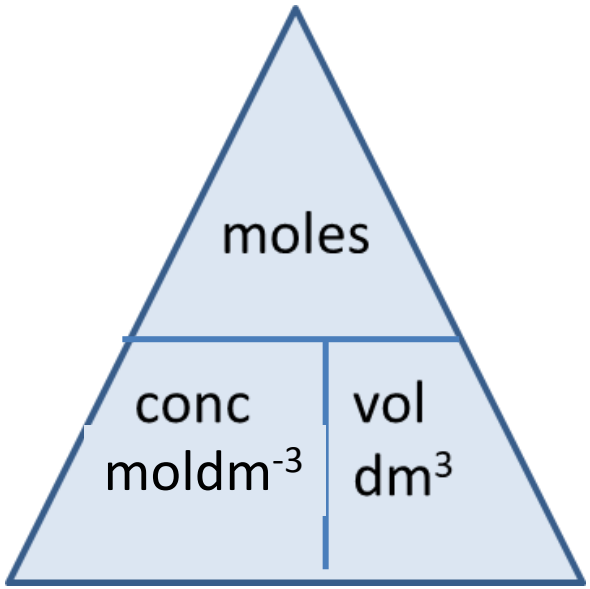
RFM = 180.0

Moles = 0.00139

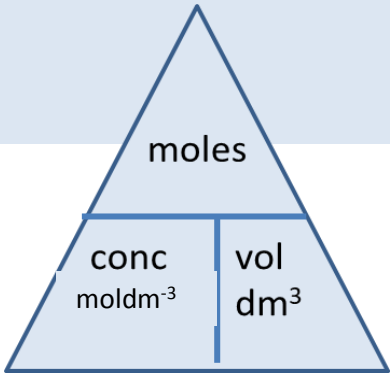
5-The Molarity Triangle

You have to learn this triangle

Concentration in moles/dm^3 (molarity, M) can be calculated using the following triangle:



5-Using the Molarity Triangle:



Bronze: What is the concentration in mol/dm^3 of:

1. 0.500 moles of hydrochloric acid in 1.00 dm^3 of water
2. 1.50 moles of sodium carbonate in 0.500 dm^3 of water

You have to learn this triangle

Silver: How many moles in?

1. 0.250 dm^3 of a 0.0500 mol/dm^3 solution of sulfuric acid
2. 0.500 dm^3 of a 2.00 mol/dm^3 solution of sodium hydroxide?

Gold: Calculate . . .

1. The concentration in mole/dm^3 from 0.750 moles of copper sulphate in 500.0 cm^3 of water.
2. The number of moles of ethanoic acid in 25.0 cm^3 of a 1.50 mol/dm^3 solution

Bronze: What is the concentration in moles/dm^3 of:

1. 0.500 moles of hydrochloric acid in 1.00 dm^3 of water

$$c = \frac{\text{moles}}{\text{volume}} = \frac{0.500}{1.00} = 0.500 \text{ mol/dm3}$$

2. 1.50 moles of sodium carbonate in 0.5 dm^3 of water

$$c = \frac{\text{moles}}{\text{volume}} = \frac{1.50}{0.500} = 3.00 \text{ mol/dm3}$$

Silver: How many moles in?

1. 0.250 dm^3 of a 0.0500 mol/dm^3 solution of sulfuric acid

$$\text{moles} = c \times \text{volume} = 0.0500 \times 0.250 = 0.0125 \text{ moles}$$

2. 0.500 dm^3 of a 2.00 mol/dm^3 solution of sodium hydroxide?

$$\text{moles} = c \times \text{volume} = 2.00 \times 0.500 = 1.00 \text{ mole}$$

Gold: Calculate . . .

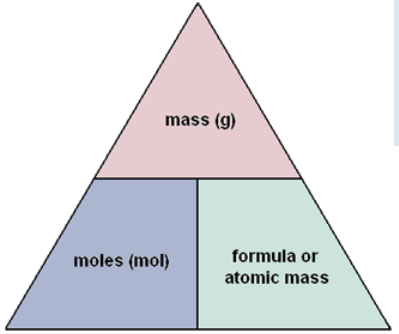
1. The concentration in mole/dm^3 from 0.750 moles of copper sulfate in 500.0 cm^3 of water.

$$c = \frac{\text{moles}}{\text{volume}} = \frac{0.750}{0.500} = 1.50 \text{ mol/dm3}$$

2. The number of moles of ethanoic acid in 25.0 cm^3 of a 1.50 mol/dm^3 solution

$$\text{moles} = c \times \text{volume} = 1.50 \times 0.0250 = 0.0375 \text{ moles}$$

6-Converting from mole/dm^3 to g/dm^3



To convert from mol/dm^3 → g/dm^3

Multiply the concentration by the RFM/RAM.

Tip: keep the dm^3 and treat this as a moles to grams calculation

What is the concentration in g/dm^3 of:

1. A 2.00 mol/dm^3 solution of HCl (RAM H=1.0, Cl=35.5)

$$\text{g}/dm^3 = \text{mol/dm}^3 \times \text{RFM} = 2.00 \times 36.5 = 73.0 g/dm^3$$

2. A 0.750 mol/dm^3 solution of NaOH (RAM H=1.0, O=16.0, Na=23.0)

$$\text{g/dm}^3 = \text{mol/dm}^3 \times \text{RFM} = 0.750 \times 40.0 = 30.0 g/dm^3$$

3. A 0.0500 mol/dm^3 solution of NaCl (RAM Na=23.0, Cl=35.5)

$$\text{g}/dm^3 = \text{mol/dm}^3 \times \text{RFM} = 0.0500 \times 58.5 = 2.93 g/dm^3$$

6-Converting from g/dm^3 to mole/dm^3

To convert the concentration from g/dm^3 → mol/dm^3 :

You divide by the RFM/RAM

Tip: Ignore the dm^3 and treat this as a grams to moles calculation

What is the concentration in mol/dm^3 of ?

1. A 2.00g/dm^3 solution of magnesium chloride, $MgCl_2$ (RAM Mg=24.3, Cl=35.5)

$$\text{mol/dm}^3 = \frac{\text{grams/dm}^3}{\text{RFM}} = \frac{2.00}{95.3} = 0.0210\ \text{mol/dm}^3$$

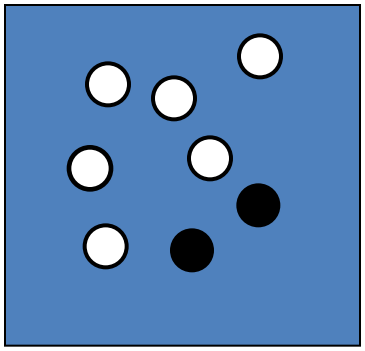
2. A 5.00g/dm^3 solution of KOH (RAM K=39.1, O=16.0, H=1.0)

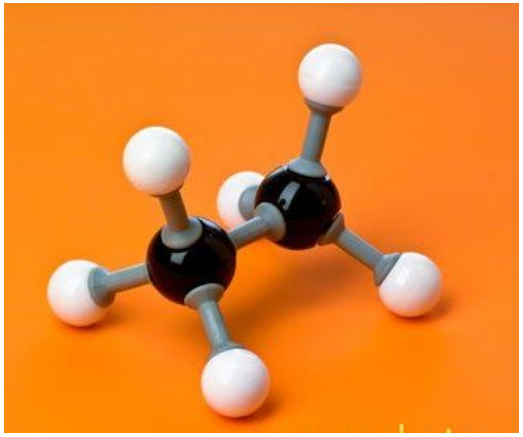
$$\text{mol/dm}^3 = \frac{\text{g/dm}^3}{\text{RFM}} = \frac{5.00}{56.1} = 0.0891\ \text{mol/dm}^3$$

7-Calculating the Concentration of an Unknown

A variety of methods can be used for this. A method will be taught in the first term that develops the understanding needed to tackle the wide variety of A Level calculations

8-Molecular v. Empirical formula





This model shows a molecule of ethane. The black circles represent carbon (C) and the white circles are hydrogen (H).

Key Words:

The molecular formula is the actual number of atoms in the molecule.

The empirical formula is the simplest **whole** number ratio formula of a compound.

Extension: What is the molecular formula of the molecule shown above? What is its empirical formula?

8-Finding the empirical formula

Example: What is the empirical formula of H_2O_2?

Answer: Divide everything by the smallest number in the formula – in this case 2. Dividing through by 2 gives HO. (each element appears once)

TASK: Find the empirical formulas of the following:

1. C_6H_6 CH
2. $C_6H_{12}O_6$ CH_2O
3. C_4H_8 CH_2
4. H_2O H_2O
5. $H_4C_4O_8$ $H_2C_2O_4$
6. P_4O_{10} P_2O_5
7. $Ca(OH)_2$ CaO_2H_2
8. $(NH_4)_2CO_3$ $N_2H_8CO_3$
9. $Mg(NO_3)_2$ MgN_2O_6

8-Exam Questions

Q1. The formula of a molecule of ethane is C_2H_6.

Give the empirical formula of ethane. (1)

To calculate the empirical formula, divide by the smallest number in the formula (2). This gives CH_3.

Q2. The formula of ammonium sulfate is $(NH_4)_2SO_4$. What is the empirical formula of ammonium sulfate?

(1)

- ☒ **A** NHSO
- ☒ **B** NH_2SO_2
- ☒ **C** NH_4SO_4
- ☒ **D** $N_2H_8SO_4$

9-Calculating the molecular formula from the empirical formula

Example: The empirical formulae of a compound is CH_2O. The relative formula mass for the molecular formula is 180.0. What is the molecular formula?

Must show working of the following steps:

Step 1: Calculate the relative formula mass for the empirical formula CH_2O:

RFM = 12.0 + (2x1.0) + 16.0= 30.0

Step 2: Divide the molecular RFM by the empirical RFM

180.0/30.0 = 6

Step 3: Multiply the empirical formula by that number

Molecular formula = $6xCH_2O = C_6H_{12}O_6$

9-Calculating the molecular formula from the empirical formula

1. The empirical formulae of a compound is $AlCl_3$. The relative formula mass for the molecular formula is 267.0. What is the molecular formula? Al_2Cl_6
2. The empirical formula of a hydrocarbon was CH_2. Find the molecular formula is the relative formula mass is 28.0. (RAM H=1.0, C=12.0) C_2H_4
3. The empirical formula of a hydrocarbon was CH. Find the molecular formula is the relative formula mass is 78.0. (RAM H=1.0, C=12.0). C_6H_6
4. The empirical formula of adipic acid is $C_3H_5O_2$. The RFM is 146 g. Calculate the molecular formula. (RAM H=1.0, C=12.0, O=16.0). $C_6H_{10}O_4$

9-Calculating the molecular formula from the empirical formula

Question: The empirical formulae of a compound is $AlCl_3$. The relative formula mass for the molecular formula is 267. What is the molecular formula?

Step 1: Calculate the relative formula mass of $AlCl_3$:

RFM = 27.0 + (3x35.5) = 133.5

Step 2: Divide the molecular RFM by the empirical RFM

267.0/133.5 = 2

Step 3: Multiply the empirical formula by that number

Molecular formula = 2x$AlCl_3$ = Al_2Cl_6

10-Calculating the Empirical Formula

Example: A compound of aluminium chloride contained 0.135g of aluminium and 0.533g of chlorine. What is its empirical formula?
(relative atomic mass (RAM) of Al=27, Cl=35.5)

Substance	**Aluminium Chloride**	
1. Elements	**Al**	**Cl**
2. $\frac{\text{Mass}}{\text{RAM}}$	$\frac{0.135\text{g}}{27} = 0.005$	$\frac{0.533}{35.5} = 0.015$
3. Divide by the smaller or smallest number	$\frac{0.005}{0.005} = 1$	$\frac{0.015}{0.005} = 3$
4. Ratio	**1:3**	
5. Formula *(must give this in the end)*	**$AlCl_3$**	

10-Exam Questions

Q1. In an experiment, 3.1 g of phosphorus reacted with 24 g of bromine to form phosphorus bromide.

Calculate the empirical formula of the phosphorus bromide.

You must show your working. (relative atomic masses: P = 31, Br = 80)

(3)

PBr_3

empirical formula .....

Q2. An oxide of lead was analysed.
0.414 g of lead was combined with 0.064 g of oxygen in this oxide.

Calculate the empirical formula of this lead oxide. (relative atomic masses: O = 16, Pb = 207)

(3)

empirical formula . . . . PbO_2 . . . . . .

Q3. A sample of calcium bromide contains 0.2 g calcium and 0.8 g bromine by mass. **Calculate the empirical formula of calcium bromide.** (relative atomic masses: Ca = 40, Br = 80)

$CaBr_2$

Q4. 14.3 g of an oxide of copper contained 12.7 g of copper.
Calculate the empirical formula of this oxide.
Show your working.
(Relative atomic masses: Cu = 63.5, O = 16)

Cu_2O

answer = ...

10-Writing a Balanced Equation

Example: 3.18g of copper reacted with 0.800g of oxygen to form a copper oxide. (Atomic Mass Cu=63.5: O=16.0) Use this information to determine the balanced equation for this reaction.

1. Calculate the empirical formula of the product.

$$\text{Cu} = \frac{3.18}{63.5} = 0.0501 \qquad \text{O} = \frac{0.800}{16.0} = 0.0500$$

Ratio 1:1

CuO

2. Write a symbol equation for the reaction:

$Cu + O_2 \rightarrow CuO$

3. Balance!

$2Cu + O_2 \rightarrow 2CuO$

10-Exam Question – writing an equation

When iron wool is heated in bromine vapour, it reacts to form iron bromide.

In an experiment, 5.60 g of iron reacted exactly with 24.0 g of bromine, Br_2.

[relative atomic masses: Fe = 56.0, Br = 80.0]

Determine, using this information, the balanced equation for the reaction between iron and bromine. You must show your working.

(4)

Question number	Answer	Additional guidance	Mark
	• calculates mol of Fe (1) • calculates mol of Br^2 (1) • determines simplest ratio/LHS of equation (1) • deduces formula of iron bromide produced/RHS of equation (1) OR • divides mass by relative atomic mass (1) • simplest ratio (1) • empirical formula (1) • deduces LHS to obtain balanced equation (1)	Example of calculation mol Fe = $\frac{5.6}{56} = 0.1$ mol Br_2 = $\frac{24}{(2\times80)}$ = 0.15 ratio Fe:Br_2 = 2 : 3/ $2Fe + 3Br_2$ $2FeBr_3/Fe_2Br_6$ Fe : Br $\frac{5.6}{56}$: $\frac{24}{80}$ 0.1 : 0.3 1 : 3 $FeBr_3$ $2Fe + 3Br_2 \rightarrow 2FeBr_3$	(4)

This is the method you can use based on the prior slides

11-Calculating the Empirical Formula

Example: An oxide of magnesium, X, has the following percentage composition by mass: Mg, 60%; O 40%.

Calculate the empirical formula of X (relative atomic mass (RAM) of O=16, Mg=24)

Tip: Treat the % exactly how you treated the masses in calculation 10

Substance	**Magnesium Oxide**	
1. Elements	**Mg**	**O**
2. $\frac{\text{Mass}}{\text{RAM}}$	$\frac{60}{24.3} = 2.47$	$\frac{40}{16.0} = 2.5$
3. Divide by the smaller number	$\frac{2.47}{2.47} = 1$	$\frac{2.5}{2.47} = 1.01 \approx 1$
4. Ratio	**1:1**	
5. Formula	**MgO**	

11-Calculating the Empirical Formula

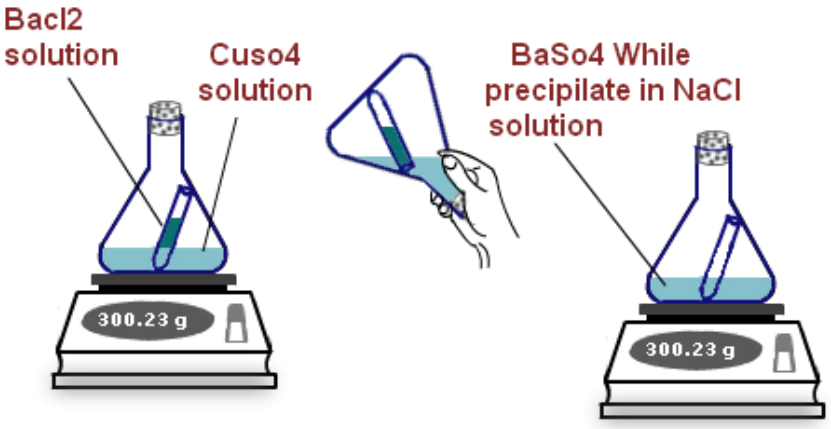
1. An hydrocarbon, **Z**, has the following percentage composition by mass: C, 80%; H 20%. Calculate the empirical formula of **X** (RAM of H=1.0, C=12.0)

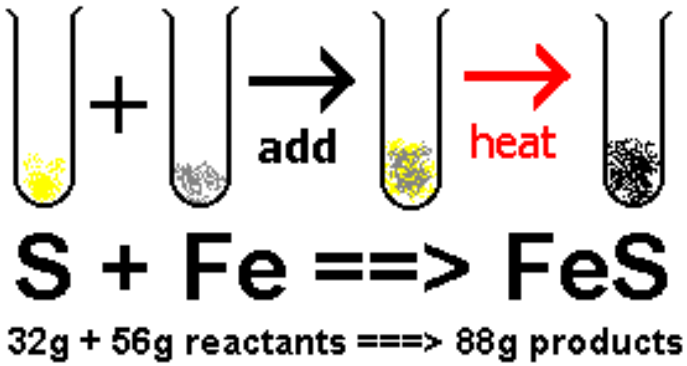
   CH_3

2. Silver nitrate has the following percentage composition by mass: Ag, 63.5%; N, 8.2%; O, 28.3%. Calculate the empirical formula. RAM of Ag=107.9, N=14.0, O=16.0)

   $AgNO_3$

12-The Law of Conservation of Mass

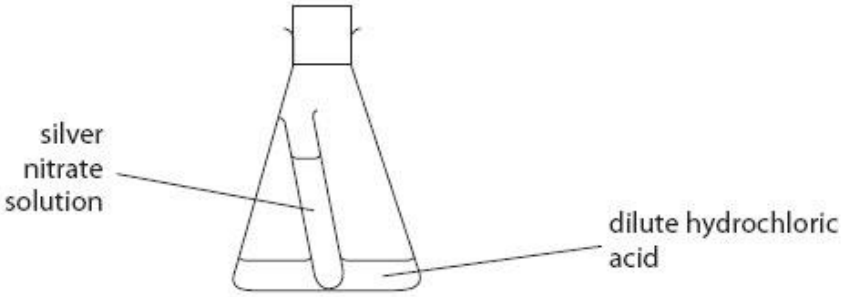




The total *mass* of *products* at the end of the reaction is equal to the total mass of the *reactants* at the beginning.

12-Exam Questions

Q1. Dilute hydrochloric acid reacts with silver nitrate solution to form silver chloride and nitric acid. This apparatus is used to investigate the mass of the reactants and the mass of products in this reaction.



The total mass of this apparatus was measured.
The flask was shaken to allow the silver nitrate solution and dilute hydrochloric acid to react.
After the reaction the total mass of the apparatus was measured again.
State how the total mass of the apparatus after the reaction will compare with the total mass of the apparatus before the reaction.

The mass will stay the same

(1)

.....

Q2. When calcium carbonate is heated strongly it undergoes thermal decomposition.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

2.50 g of calcium carbonate was heated strongly.
1.40 g of solid remained after heating.
Calculate the mass of carbon dioxide produced during this reaction.

The mass of the products = mass of the reactants. So 2.50 = 1.40 + mass of CO_2 (1)

.....

Mass of CO_2 = 2.50-1.40 = 1.10g

.....

12-Exam Questions

Q3. Propene can be made by cracking fractions obtained from crude oil.
This equation shows the cracking of decane to produce propene and butane.

$$C_{10}H_{22} \rightarrow 2C_3H_6 + C_4H_{10}$$

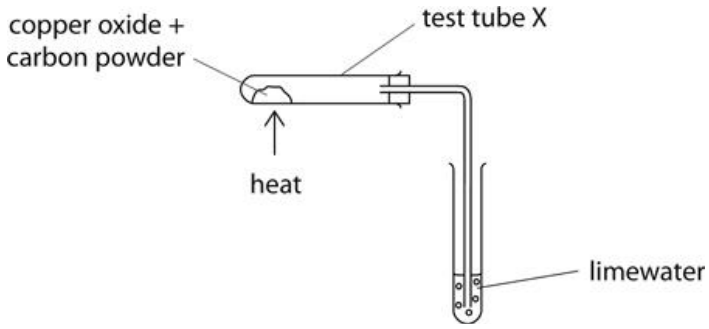
decane propene butane

Give the total mass of products formed if 17 g of decane is cracked in this way.

The mass products = mass of reactants = 17g (1)

.....

Q4. A mixture of copper oxide and carbon powder was heated. Carbon dioxide was produced. It was bubbled into limewater.



The word equation for the reaction is
copper oxide + carbon → copper + carbon dioxide

The mass of test tube X and its contents was measured before heating and after heating.
There was a change in mass.
Explain why the total mass of the test tube and contents changes during the reaction.

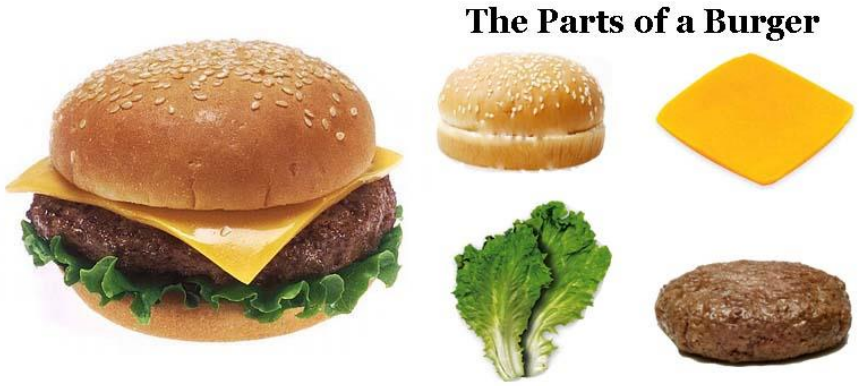
The decreased (got smaller) because the CO_2 escaped from the test tube (2)

.....

.....

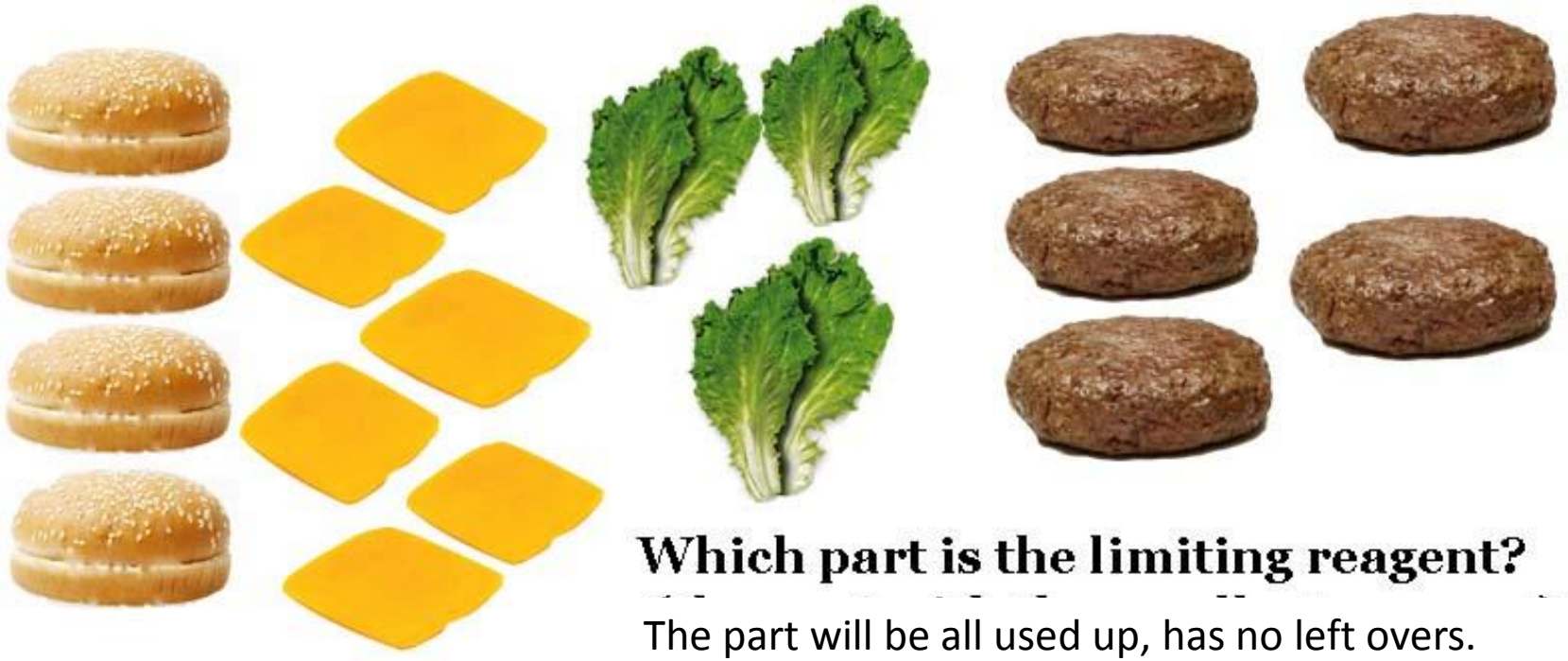
13-Reacting Mass Calculations

A variety of methods can be used for this. A method will be taught in the first term that develops the understanding needed to tackle the wide variety of A Level calculations



- Only 3 burgers can be made.
- The limiting reagent is the lettuce leaves.

Question 1
How many burgers can be made?



Which part is the limiting reagent?

The part will be all used up, has no left overs.

14-What is a limiting reagent?

The limiting reagent is the reactant will all react. The amount of product you make depends of how much of the limiting reagent you have.

It is not necessarily the least amount you have, such as the lettuce, there are six pieces but you need two in each burger, despite there are less burgers and buns, the amount of burger made is still limited by the amount of lettuce.

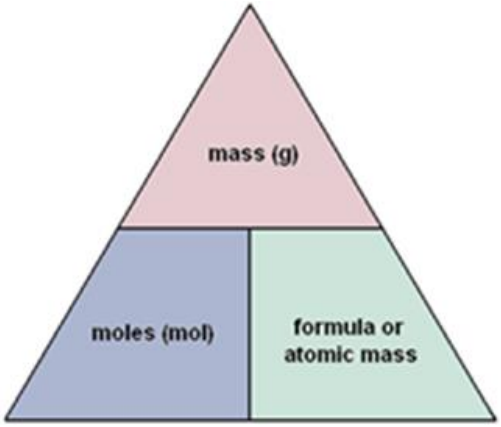
The reactant that has some left after a reaction is said to be in excess.

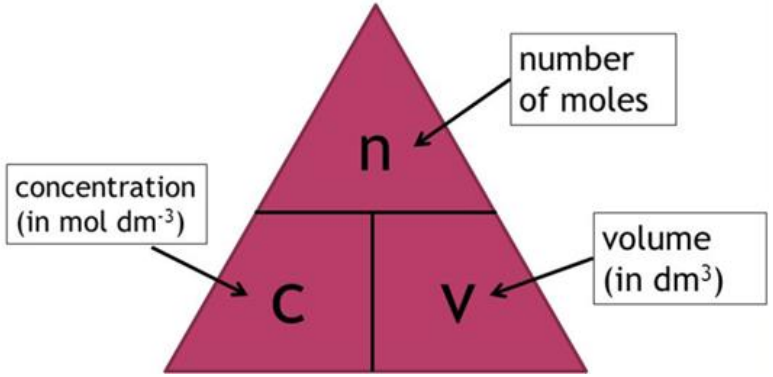
You could be asked to find the limiting reagent or the reagent in excess.

14-Finding the Limiting Reagent

To find the limiting reagent (or the reagent in excess) compare the moles of both reactants; divide the moles by the mole ration number in the balanced equation and see which is greater.

To find the moles, you will have to use one of the triangles at the beginning of this powerpoint such as:



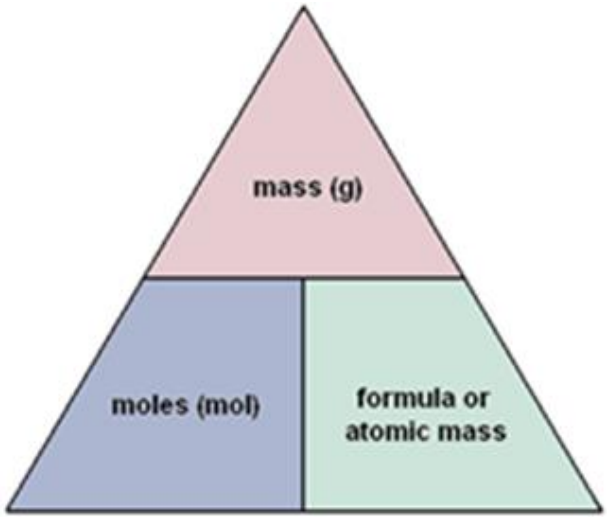


14-Finding the Limiting Reagent

Example 1: Carbon reacts with oxygen to form carbon dioxide:

If 1.00g of carbon reacts with 1.00g of oxygen, which is the limiting reagent? Which is in excess?

	C	+	O_2 →	CO_2
Mole ratio	1		1	
Moles	$\frac{1.00}{12.0} = 0.083$		$\frac{1.00}{32.0} = 0.0300$	
$\frac{\text{mole}}{\text{mole ratio}}$	0.083		0.0300	



Use this triangle to find moles

Answer: Oxygen is the limiting reagent (smaller/smallest mole/mole ratio value) and Carbon is in excess

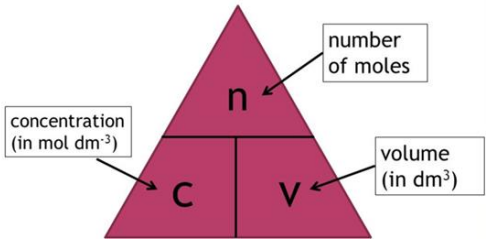
14-Finding the Limiting Reagent

Example 2: Sulphuric acid and sodium hydroxide neutralise each other.

If 1.0 dm^3 of 0.30 mol/dm^3 H_2SO_4 reacts with 1.0dm^3 of 0.50 mol/dm^3 NaOH, which is the limiting reagent? Which is in excess?

$$H_2SO_4 + 2NaOH \rightarrow NaCl + H_2O$$

Mole ratio	1	2
Mole	1.0x0.30=0.30	1.0x0.50=0.50
$\frac{\text{mole}}{\text{mole ratio}}$	0.30/1=0.30	0.50/2=0.25



Use this triangle to find moles

Answer: NaOH is the limiting reagent and HCl is in excess

NB: Discard mole/mole ratio value for further calculations, go back to use the moles above it for further calculations! This is crucial!

14-Finding the Limiting Reagent

1. What is the limiting reagent when 5.09g of Fe reacts with 5.00g of S to form iron sulphide?

Fe + S → FeS (RAM of Fe=55.8, S=32.1)

Answer: Iron is the limiting reagent and sulfur is in excess

2. What is the limiting reagent when 5.0 dm^3 of 0.25 mol/dm^3 of HCl reacts with 2.0 dm^3 of 0.5 mol/dm^3 NaOH?

$HCl + NaOH \rightarrow NaCl + H_2O$

Answer: NaOH is the limiting reagent and HCl is in excess

15-Calculating Percent Yield

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Questions:

1. What is the percentage yield of a reaction where the theoretical yield was 75 kg but the actual yield was 68 kg?

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{68}{75} \times 100 = 91\%$$

2. During a practical a student made 30g of product, but the theoretical yield was 40g. What was the percentage yield?

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{30}{40} \times 100 = 75\%$$

16-What is atom economy?

The atom economy of a chemical reaction is a measure of the amount of starting materials that become useful products.

$$\%\,\text{atom economy} = \frac{\text{total RFM of desired product}}{\text{total RFM mass of all products}} \times 100\%$$

16-How to calculate atom economy

Example: What is the atom economy for making hydrogen by reacting coal with steam?

$C(s) + 2H_2O(g) \rightarrow CO_2(g) + 2H_2(g)$

STEP 1: Calculate the total RFM of the desired product (H_2):

$C(s) + 2H_2O(g) \rightarrow CO_2(g) + 2H_2(g)$

$2 \times 2.0 = 4.0$

STEP 2: Calculate the total RFM mass of products

$C(s) + 2H_2O(g) \rightarrow CO_2(g) + 2H_2(g)$

$44.0 + 2 \times 2.0 = 48.0$

STEP 3: Put values into equation

% atom economy = $^{4.0}/_{48.0} \times 100 = 8.30\%$

16-Calculating atom economy

Questions:

1. Calculate the atom economy for making hydrogen from methane:

$CH_4 + H_2O \rightarrow CO + 3H_2$ (RAM H=1.0, C=12.0, O=16.0)

STEP 1: Total RFM of desired product = 3 x 2.0 = 6.0
STEP 2: Total RFM of all products = 28.0 + 6.0 = 34.0
STEP 3: Atom economy = 6.0/34.0 x 100 = 17.6 %

2. What is the atom economy of this process to make ethanol?

$C_2H_4 + H_2O \rightarrow C_2H_5OH$ (RAM H=1.0, C=12.0, O=16.0)

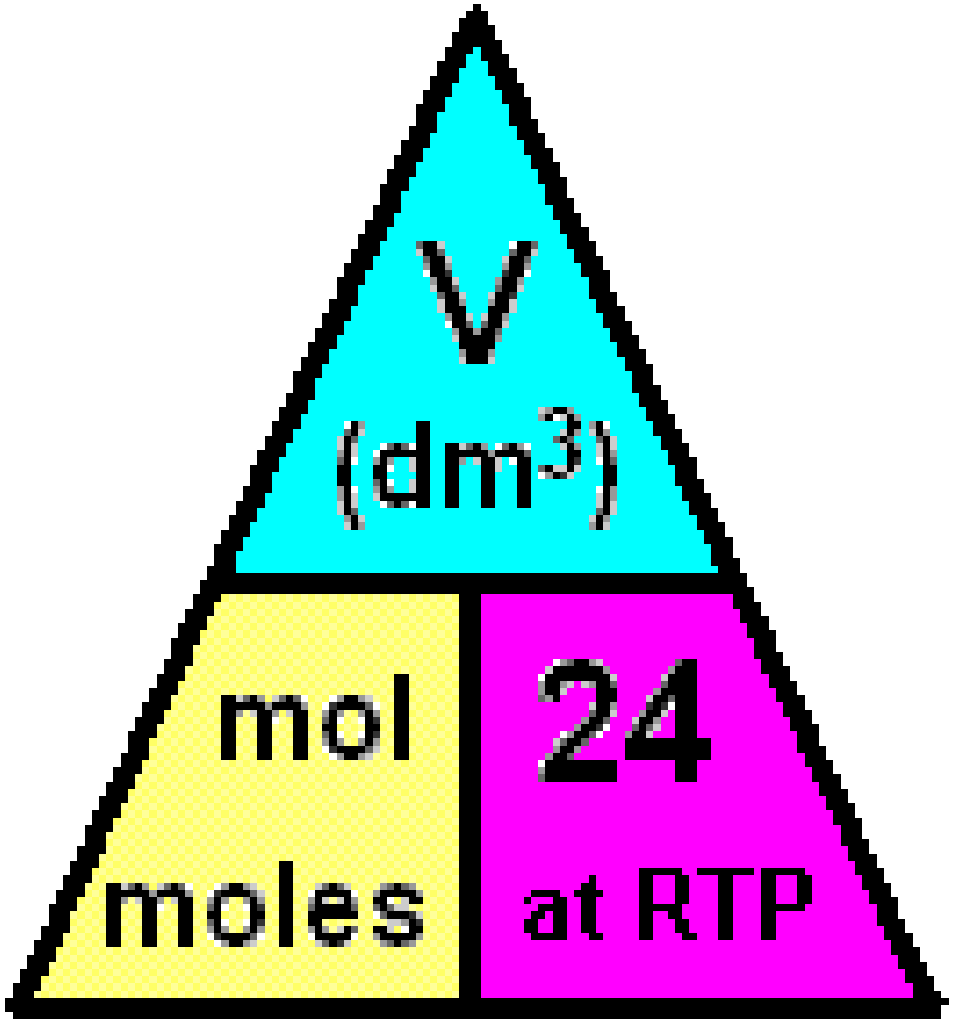
Because there is only one product the atom economy will be 100%

3. What is the atom economy of extracting iron from its ore?

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$ (RAM Fe=55.8, C=12.0, O=16.0)

STEP 1: Total RFM of desired product = 2 x 56.0 = 111.8
STEP 2: Total RFM of all products = 111.8 + 3x44.0 = 243.8
STEP 3: Atom economy = 111.8/243.8 x 100 = 45.9 %

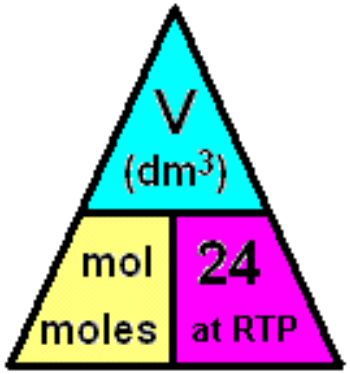
17-Using the molar volume



17-Calculating volume from moles

(NOT CORRECT SIG FIGS ARE GIVEN IN ANSWERS)

For all questions assumes it is room temperature and pressure (RTP), so the molar volume is 24 dm^3.



1. What is the volume of 1.5 moles of H_2 gas?

   $volume = moles \times 24 = 1.5 \times 24 = 36\ dm^3$

2. What is the volume of .25 moles of O_2 gas?

   $volume = moles \times 24 = 0.25 \times 24 = 6\ dm^3$

3. How many moles of CO_2 are there in 48 dm^3 of gas?

   $$moles = \frac{volume}{24} = \frac{48}{24} = 2 \text{ moles}$$

4. How many moles of Cl_2 are there in 2 dm^3 of gas?

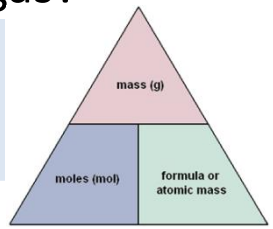
   $$moles = \frac{volume}{24} = \frac{2}{24} = 0.083 \text{ moles}$$

<u>Extension</u>:

1. How many grams of nitrogen are there in 10 dm^3 of nitrogen (N_2) gas? (RAM N = 14)

   STEP 1: $moles = \frac{volume}{24} = \frac{10}{24} = 0.417$ moles;

   STEP 2: mass= moles X RFM =11.7

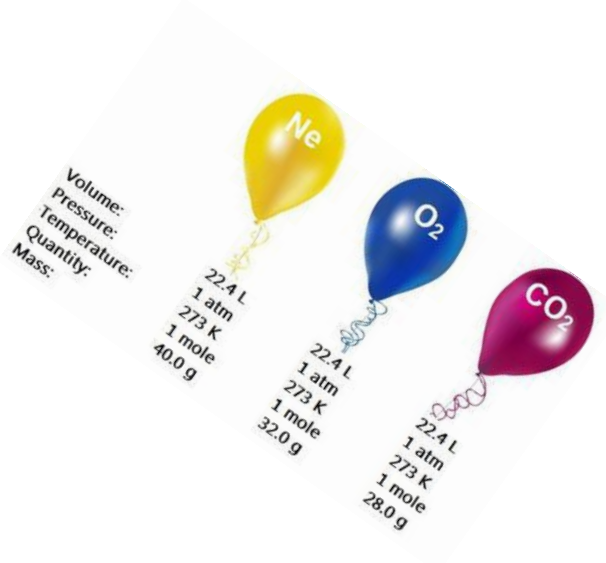


2. What is the volume of 1.2 g of Ne gas? (RAM Ne = 20)

   STEP 1: $\text{moles} = \frac{mass}{RAM} = \frac{1.2}{20} = 0.06$ STEP 2: $\text{Volume} = \text{moles} \times 24 = 1.44\ dm^3$

17-Gas Volumes and Reacting Masses

- Because the volume of gases is directly linked to the number of moles (and volume is the same for each gas), volumes can be used instead of moles in reacting mass calculations.



<u>Example</u>:

$Cl_2(g) + H_2(g) \rightarrow 2HCl\ (g)$

If there is 10.0 dm^3 of Cl_2, then there needs to be 10.0 dm^3 of H_2 to react completely with it. There would be 20.0 dm^3 of HCl made because the ratio is 2 to 1.

17-Exam Questions

Q1. Sulfur trioxide is produced by reacting sulfur dioxide with oxygen.

$$2SO_2 + O_2 \rightleftharpoons 2SO_3$$

What volume of oxygen, in cm^3, would react completely with 500 cm^3 sulfur dioxide?

(1)

- ☒ **A** 500 ÷ 2
- ☒ **B** 500
- ☒ **C** 500 × 2
- ☒ **D** 500 × 32

Q2. When nitrogen and hydrogen react to form ammonia, the reaction can reach a dynamic equilibrium.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Calculate the minimum volume of hydrogen required to completely convert 1000 dm^3 of nitrogen into ammonia.

(1)

.....

..... 3000

volume of hydrogen = ... dm^3

Q3. Hydrogen reacts with oxygen to form water vapour.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

If 200 cm^3 of hydrogen react completely with 100 cm^3 of oxygen, what is the maximum volume of water vapour formed, if all volumes are measured at the same temperature and pressure?

- ☒ **A** 100 cm^3
- ☒ **B** 200 cm^3
- ☒ **C** 300 cm^3
- ☒ **D** 400 cm^3

18-Isotopes & Calculating Relative Atomic Mass

How to Calculate Relative Atomic Mass.

Example. 80% of Boron atoms are the Boron-11 isotope. 20% of Boron atoms are the Boron-10 isotope. What is the relative atomic mass of Boron?

Step 1: (80 × 11) + (20 × 10) = 1080

Step 2: 1080 ÷ 100 = **10.8**

18-Isotopes & Calculating Relative Atomic Mass

1. 75% of chlorine atoms are the ^{35}Cl isotope. 25% of chlorine atoms are the ^{37}Cl isotope. What is the relative atomic mass of chlorine?

35.5

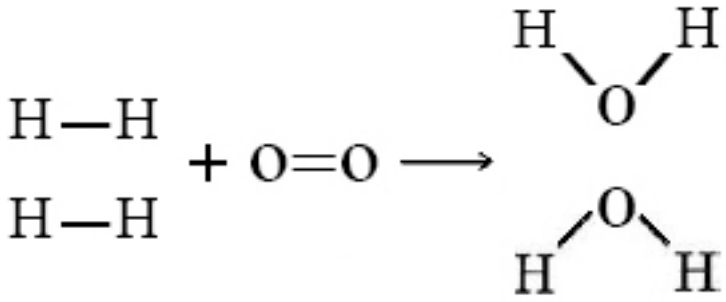
2. Lithium has an atomic number of 3. A sample of lithium is 7.6% Lithium-6 and 92.4% Lithium-7. Calculate the relative atomic mass of lithium. 6.9

3. Neon has an atomic number of 10. A sample of neon is 90.5% Neon-20. The rest of the sample is Neon-22. Calculate the relative atomic mass of neon. 20.2

4. A sample of iron contains 6% Iron-54, 92% Iron-56 and 2% Iron-57. What is the relative atomic mass of iron in this sample? 55.9

19-Bond Energy Calculations

Example: Calculate the energy change when water is formed from H_2 and O_2.



STEP 1 Bonds Broken

2 x (H-H) = 2 x 436 = 872

1 x (O=O) = 498

Total = 872 + 498 = 1370

Bond	Bond Energy
H-H	436
H-O	464
O=O	498

STEP 2 Bonds formed

4 x (O-H) = 4 x 464 = 1856

STEP 3

Energy change = bonds broken – bonds formed (BB-BF)

= 1370 – 1856 = -486

The negative sign means its exothermic.

19-Exam Question

Q5. The energies of some bonds are shown in Figure 13.

bond	energy of bond / kJ mol^{-1}
H—H	436
Cl—Cl	243
H—Cl	432

Figure 13

Hydrogen reacts with chlorine to form hydrogen chloride.

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

Calculate the energy change, in kJ mol^{-1}, for the reaction of 1 mol of hydrogen gas, H_2, with 1 mol of chlorine gas, Cl_2, to form 2 mol of hydrogen chloride gas, HCl.

(4)

19-Exam Question -working

STEP 1 Bonds Broken

1 x (H-H) = 436

1 x (Cl-Cl) = 243

Total = 436 + 243 = 679

STEP 2 Bonds Made

2 x (H-Cl) = 2 x 432 = 864

STEP 3

Energy change = bonds broken – bonds made

= 679 – 864 = -185 exothermic