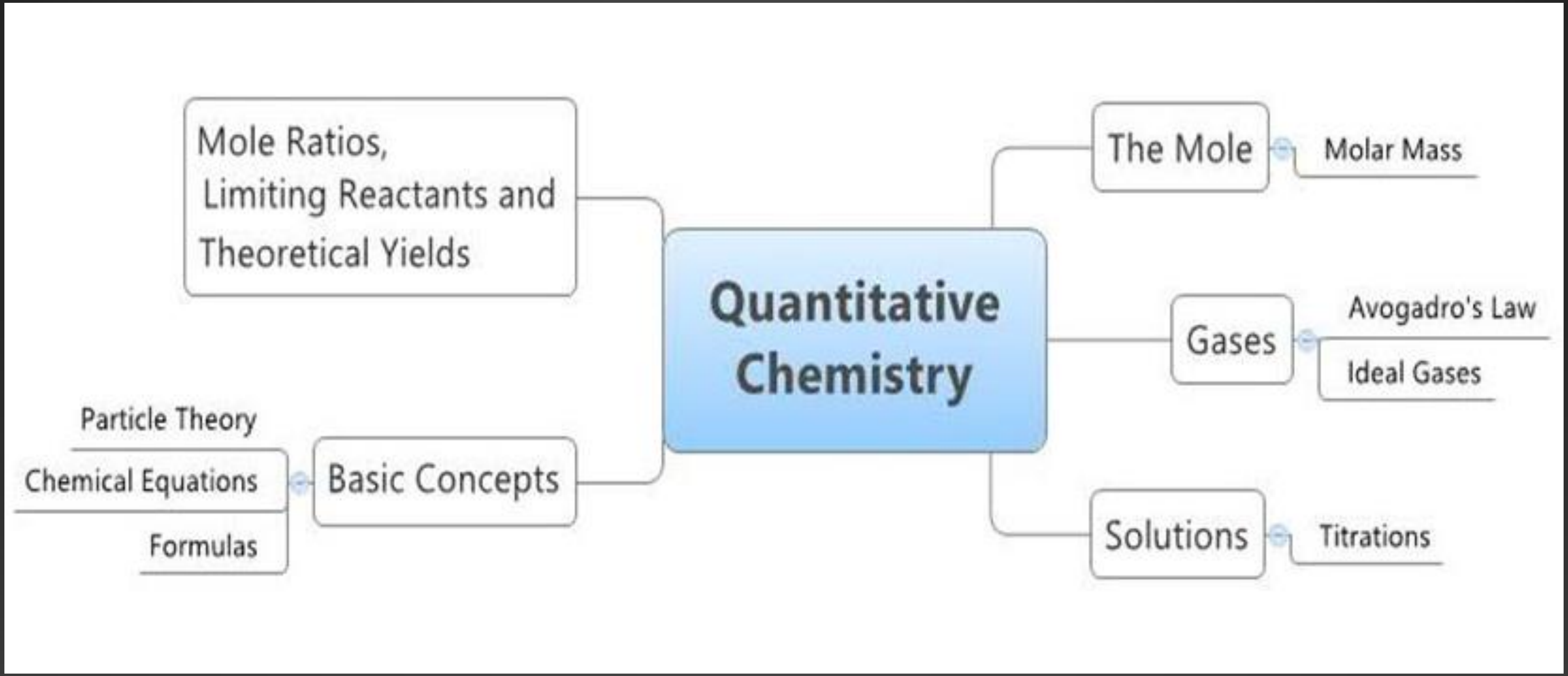
QUANTITATIVE CHEMISTRY



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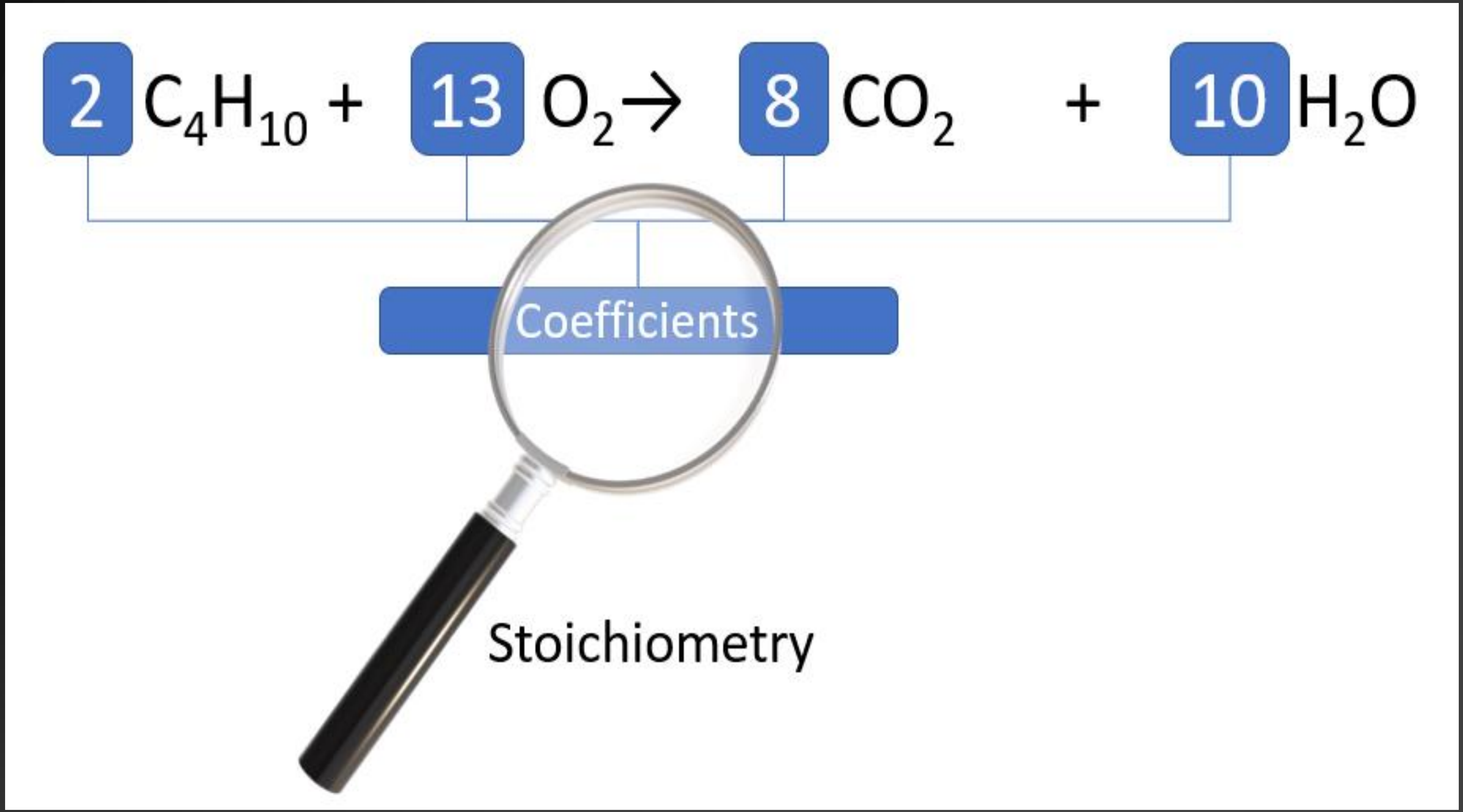
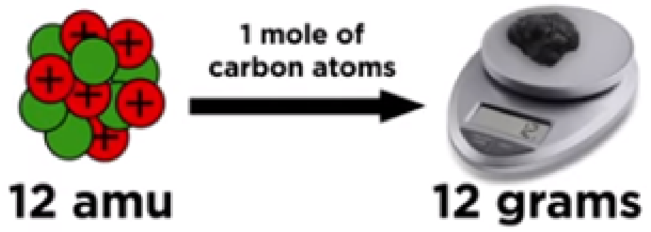


It is not practical to quantify on the atomic/molecular level as we cannot see or measure at this scale

We cannot see atoms, but we can see moles

Moles

Mass



Avogadro's Number

Atomic Mass Units

Molar Mass

Molecular Mass

quintillions

602,200,000,000,000,000,000,000

sextillions

quadrillions

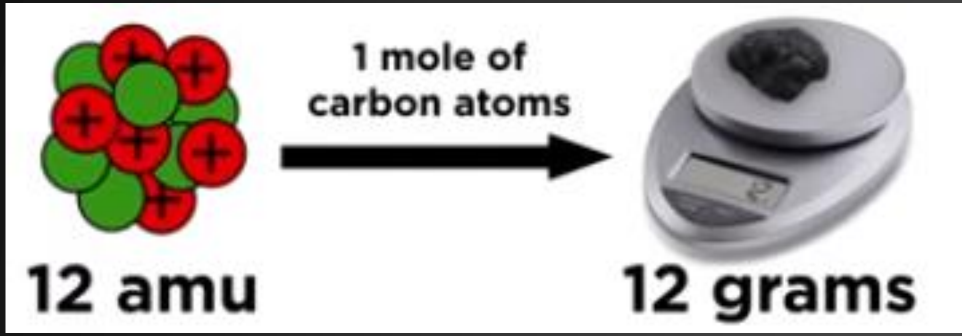
trillions

billions

millions

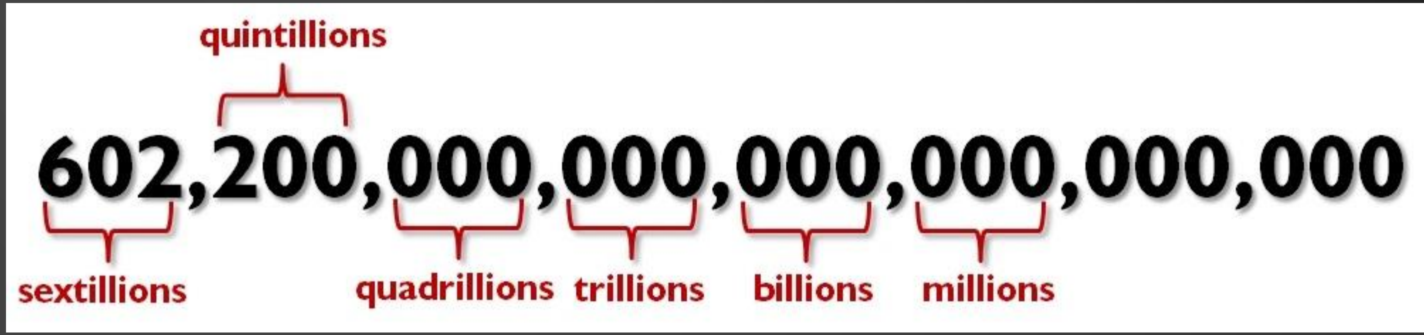
Combining the relative masses of the atoms that make up a substance

MOLES



A mole is the number of atoms in 12g of $Carbon_{12}$

1 mole of $Carbon_{12}$ has a mass of 12g. What is the mass of 1 mole of $Nitrogen_{14}$?

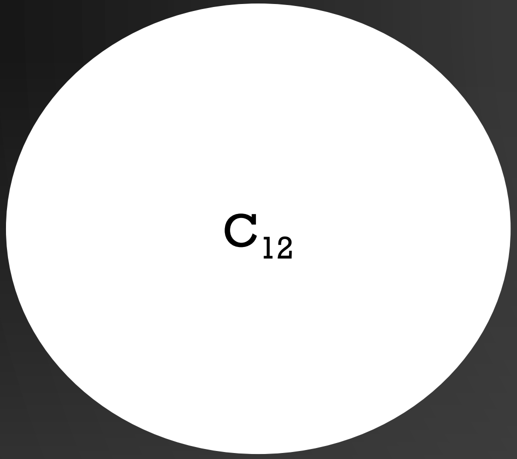


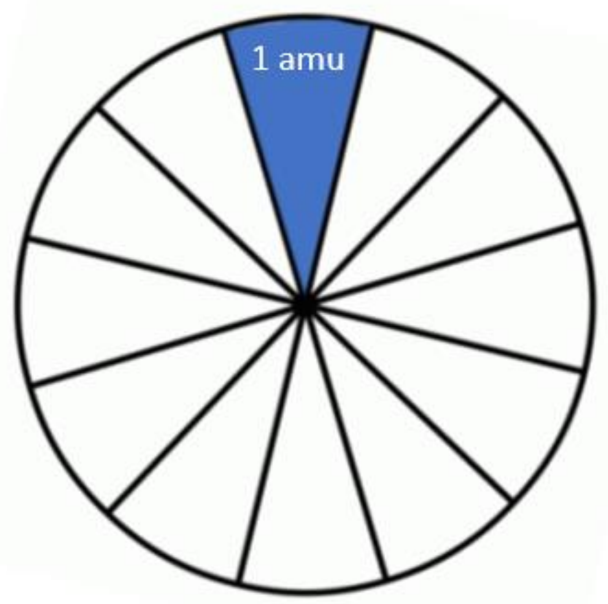
Avogadro's Number: $6.02x10^{23}$

RELATIVE ATOMIC MASS (A_r)

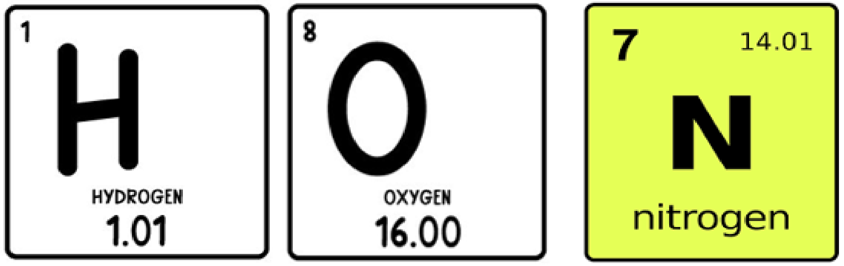
- Because the mass of a single atom is so small, we use scales of relative mass
- The masses of atoms/molecules are compared to the mass of an atom of carbon-12, which has an assigned mass of 12.00
- Units = atomic mass units (amu)

$$1 \text{ amu} = \frac{1}{12} \text{ of a carbon}_{12} \text{ atom} = 1\text{g}$$



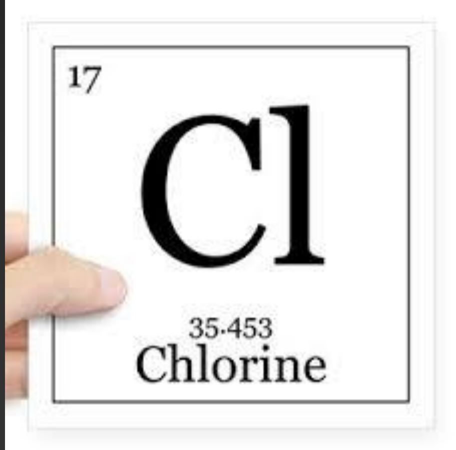


Examples	1 Mole
Hydrogen = $\frac{1}{12}$ of a carbon$_{12}$ atom = 1 amu	1g
Oxygen = $\frac{16}{12}$ of a carbon$_{12}$ atom = 16 amu	16g
Nitrogen = $\frac{14}{12}$ of a carbon$_{12}$ atom = 14 amu	?



- Isotopes are factored in when assigning an A_r value
  - Example: Chlorine has 2 isotopes: chlorine_{35} and chlorine_{37}

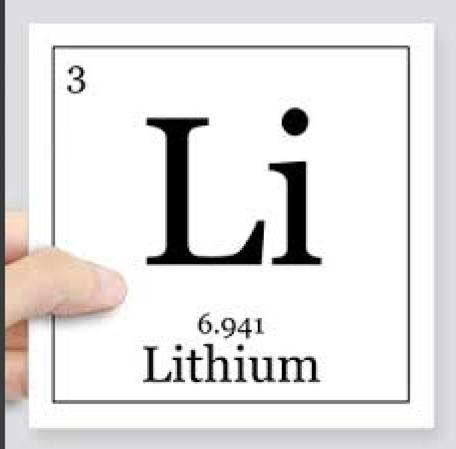
Abundance: 77.35% 22.65%



$$A_r = (35 \text{x} 0.7735) + (37 \text{x} 0.2265)$$

$$= 35.453$$

Calculate the A_r for Lithium (Li). Isotopes: $\text{Lithium}_6 = 5.92\%$, $\text{Lithium}_7 = 94.08\%$



$$(6 \text{ x } 0.0592) + (7 \text{ x } 0.9408) = 6.941$$

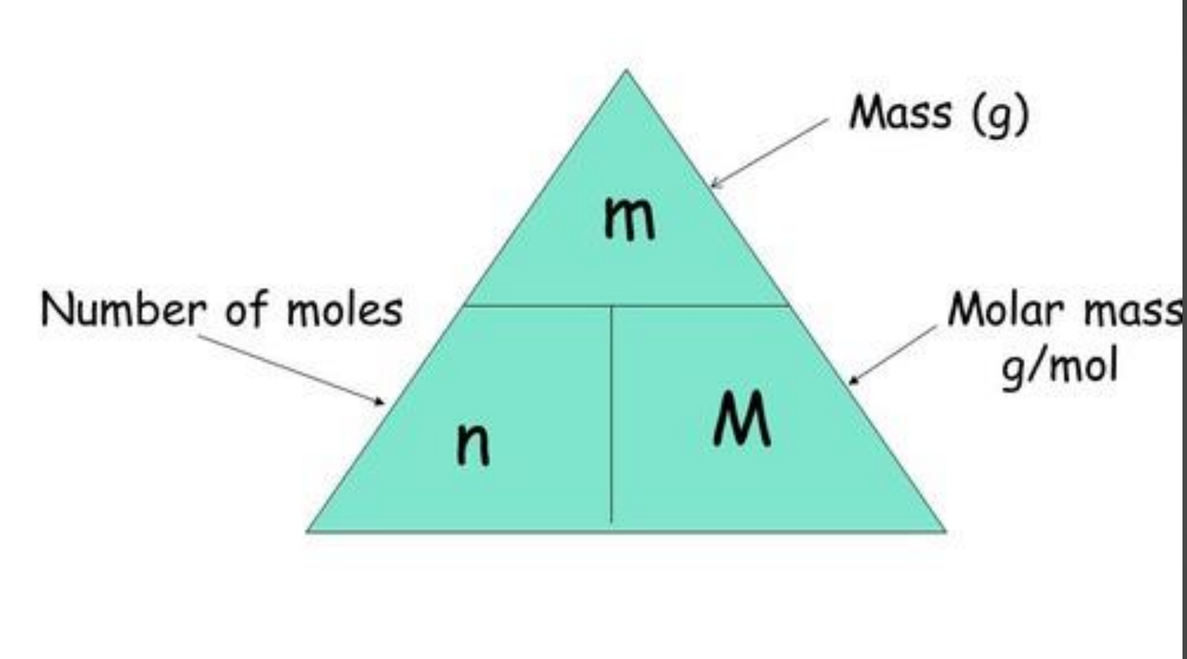
RELATIVE MOLECULAR MASS (M_r)

- The relative molecular mass is the sum of the A_r for the atoms making up a molecule

- The M_r of methane (CH_4) is:
  - 12.01 (A_r of C) + 4x1.01 (A_r of H) = 16.05

- Calculate M_r for glucose ($C_6H_{12}O_6$)
  - A_r of C = 12.01
  - A_r of H = 1.01
  - A_r of O = 15.99

USING MOLES, MASS AND MOLAR MASS



- **Carbon** 6 x 12.011 g/mol = 72.066
- **Hydrogen** 12 x 1.008 g/mol = 12.096
- **Oxygen** 6 x 15.999 g/mol = 95.994

180.156 Molar mass (g/mol)

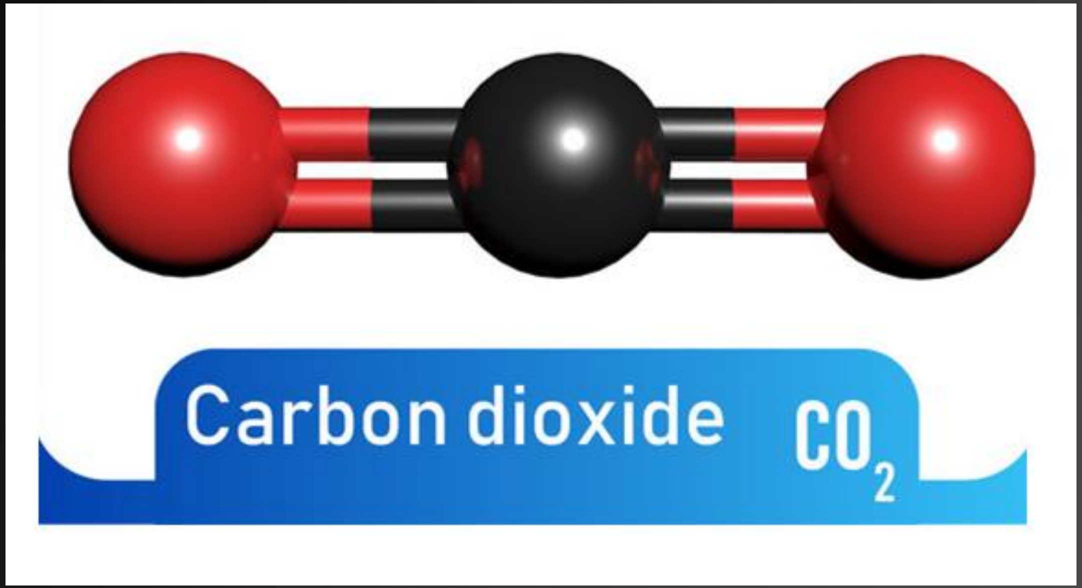
- Calculate the number of moles for 116.422g of glucose

116.422 / 180.156 = 0.646 moles

Video explaining the relationship between Avogadro's number, mass in grams and relative masses (molar mass)

https://www.youtube.com/watch?v=M5q_YMN4CtY

4 min video



Molecular Mass

12.01 + 16 +16 = 48.01

If there are 6.02×10^{23} molecules of CO2 then there is 1 mole of CO2

1 mole of CO2 has a mass of 48.01 grams

EMPIRICAL FORMULAS

An empirical formula is the simplest whole number ratio of the elements present in a compound

- Example CH_2 – this ratio tells us that for every carbon atom there are 2 hydrogen atoms

- An empirical formula can be calculated based on the % mass of each element present in a compound.

$$\text{The \% by mass of an element} = \left(\frac{\text{number of atoms of the element x } A_r}{M_r}\right) \text{x } 100$$

M_r $C_6H_5NO_2$ = 123.10

A_r C = 12.01 A_r O = 15.99 C = 58.54% O = 25.98%

A_r H = 1.01 A_r N = 14.01 H = 4.10% N = 11.38%

An unknown sample has been analysed and found to contain 47% potassium, 14.5% carbon and 38.5% oxygen. What is its empirical formula?

Step 1 – convert to g

- If you have been asked to calculate an empirical formula based on % composition, then you assume that the mass is 100g. This way 1%=1g.

K = 47.00% = 47.0g
C = 14.50% = 14.5g
O = 38.50% = 38.5g

A_r K = 39.09 A_r C = 12.01 A_r O = 16

Step 2 – convert to moles

Number of moles = mass of substance (g) / A_r

K = 47.0/39.09 = 1.20 moles
C = 14.5/12.01 = 1.21 moles
O = 38.5/16.00 = 2.41 moles

Step 3 – Divide each value by the smallest, and round to a whole number

K = 1.20/1.20 = 1
C = 1.21/1.20 = 1
O = 2.41/1.20 = 2

Empirical formula = KCO_2

- A compound was analysed and found to contain 53.4% Ca, 43.8% O, and 2.8% H.

  What is the empirical formula of the compound?

- Step 1 – convert to grams
- Step 2 – calculate # of moles
- Step 3 – divide each value by the smallest and round to a whole number
- Step 4 – write the empirical formula

A_r Ca = 40.1 A_r O = 16.00 A_r H = 1.01

- A compound was analysed and found to contain 57.14% C, 6.16% H, 9.52% N and 27.18% O. What is the empirical formula of the compound?

C = 57.14%= 57.14g
H = 6.16% = 6.16g
N = 9.52% = 9.52g
O = 27.18% = 27.18g

C = 57.14 / 12.01 = 4.76 mol
H = 6.16 / 1.01 = 6.10 mol
N = 9.52 / 14.01 = 0.68 mol
O = 27.18 / 15.99 = 1.70 mol

C = 4.76 / 0.68 = 7
H = 6.10 / 0.68 = 9
N = 0.68 / 0.68 = 1
O = 1.70 / 0.68 = 2.5 Too far to round

Multiply each by the lowest factor that gives a whole number
C = 4.76 / 0.68 = 7 x 2 = 14
H = 6.10 / 0.68 = 9 x 2 = 18
N = 0.68 / 0.68 = 1 x 2 = 2
O = 1.70 / 0.68 = 2.5 x 2 = 5

Empirical Formula = $C_{14}H_{18}N_2O_5$

MOLECULAR FORMULAS

- A molecular formula is the total number of atoms of each element present in a molecule of the compound.
  - It is a multiple of the empirical formula (simplest whole number ratio).

Substance	Molecular Formula	Empirical Formula
Benzene 78.12g/mol	C_6H_6	CH
Acetylene 26.04g/mol	C_2H_2	CH
Glucose 180.156g/mol	$C_6H_{12}O_6$	CH_2O
Water 18.01g/mol	H_2O	H_2O

To calculate the molecular formula from an empirical formula you need the M_r and the empirical formula mass.

Example: The empirical formula of benzene is CH; the molar mass is 78.12 g mol^{-1}. What is the molecular formula?

Mass of empirical formula unit = (12.01+1.01) = 13.02
N = 78.12 / 13.02 = 6

The empirical formula unit occurs 6 times (n)
Therefore, the molecular formula = C_6H_6

COMBUSTION REACTIONS

- A 0.250g sample of hydrocarbon undergoes complete combustion to produce 0.845g of CO_2 and 0.173g of H_2O. What is the empirical formula of this compound?

Step 1 - Determine the grams of C in 0.845 g CO_2 and the grams of H in 0.173 g H_2O.

- carbon: 0.845 g x (12.01 / 43.99) = 0.2307g
- hydrogen: 0.173 g x (2.02 / 18.01) = 0.0194g

M_r CO_2 = 43.99
M_r H_2O = 18.01

Step 2 - Convert grams of C and H to their respective amount of moles.

- carbon: 0.2307 / 12.011 = 0.0192 mol
- hydrogen: 0.0194 / 1.01 = 0.0192 mol

Empirical formula = CH

- A 0.250g sample of a compound known to contain carbon, hydrogen and oxygen undergoes complete combustion to produce 0.3664g of CO_2 and 0.150g of H_2O. What is the empirical formula of this compound?

Step 1 - Determine the grams of carbon in 0.3664 g CO_2 and the grams of hydrogen in 0.1500 g H_2O

carbon: 0.3664 x (12.01 / 43.99) = 0.1000 g
hydrogen: 0.15 x (2.02 / 18.01) = 0.0168 g

Grams of oxygen = 0.25 – (0.1 + 0.0168) = 0.1332

Step 2 – Convert grams to moles
carbon: 0.1000 / 12.01 = 0.0083 mol
hydrogen: 0.0168 / 1.01 = 0.0166 mol
oxygen = 0.1332 / 15.99 = 0.0083 mol

Step 3 – Divide by the lowest value
carbon: 0.0083 / 0.0083 = 1
hydrogen: 0.0166 / 0.0083 = 2
oxygen: 0.0083 / 0.0083 = 1

Empirical Formula = CH_2O

C: 12.01
H: 1.01
O: 16.00
H_2O: 18.02
CO_2: 44.01
O_2: 32

PRACTICE PROBLEMS

Use conservation of mass

A hydrocarbon fuel is fully combusted with 18.214g of oxygen to yield 23.118g of CO_2 and 4.729g of H_2O. Find the empirical formula for the hydrocarbon.

12.915g of a substance containing only carbon, hydrogen, and oxygen was combusted to yield 18.942g CO_2 and 7.749g H_2O. Determine the empirical formula of the substance.

IONIC EQUATIONS

- Just like mass, charge is conserved in a reaction
- Although an equation can be balanced with regards to the number of atoms it may be unbalanced with regards to charge.

$$Cr_2O_7^{2-} + Fe^{2+} + H^+ \longrightarrow Cr^{3+} + H_2O + Fe^{3+}$$

- Balance the number of atoms, then balance the total charge of the products and reactants
- Balance all atoms except H and O, add H_2O to the side deficient in O, add H^+ to balance H

- State symbols are often included in chemical equations

  - $H_2O_{(l)}$ — (l) = liquid
  - $CO_{2(g)}$ — (g) = gas
  - $MgCl_{(S)}$ — (s) = solid
  - $HCl_{(aq)}$ — (aq) = aqueous (dissolved in water)

Total and Net ionic equations

Aqueous sodium hydroxide reacts with aqueous copper(II) sulphate to precipitate copper(II) hydroxide

$$2NaOH_{(aq)} + CuSO_{2(aq)} \longrightarrow Na_2SO_{2(aq)} + Cu(OH)_{2(s)}$$

To write a total ionic equation we have to break up aqueous solutions into constituent ions

$$2Na^+ + 2OH^- + Cu^{2+} + SO_2^{2-} \longrightarrow 2Na^+ + SO_2^{2-} + Cu(OH)_{2(s)}$$

To write a net ionic equation we remove the spectator ions from the equation

$$Cu^{2+} + 2OH^- \longrightarrow Cu(OH)_{2(s)}$$

Name	***Ion***	***Name***	***Ion***
Hydride	H^+	Nitrate	NO_3^-
Hydroxide	OH^-	Phosphate	PO_4^{3-}
Carbonate	CO_3^{2-}	Ammonium	NH_4^+
Sulphite	SO_3^{2-}	Oxalate	$C_2O_4^{2-}$
Sulphate	SO_4^{2-}	Permanganate	MnO_4^-

Alkali Metals (Group 1): +1 ions Li^+, Na^+, K^+

Alkali Earth Metals (Group 2): +2 ions Mg^{2+}, Ca^{2+}, Rb^{2+}

Halogens (Group 7): −1 ions F^-, Cl^-, Br^-,

In order to break apart aqueous solutions into constituent ions, it helps to be familiar with the charges of common ions.

Periodic Table knowledge

PRACTICE PROBLEMS

$$K_2CO_{3(aq)} + 2\,AgNO_{3(aq)} \longrightarrow 2\,KNO_{3(aq)} + Ag_2CO_{3(s)}$$

$$H_2SO_{4(aq)} + 2\,NaOH_{(aq)} \longrightarrow Na_2SO_{4(aq)} + 2H_2O_{(l)}$$

$$3\,HCl_{(aq)} + Al(OH)_{3(aq)} \longrightarrow 3\,H_2O_{(l)} + AlCl_{3(aq)}$$

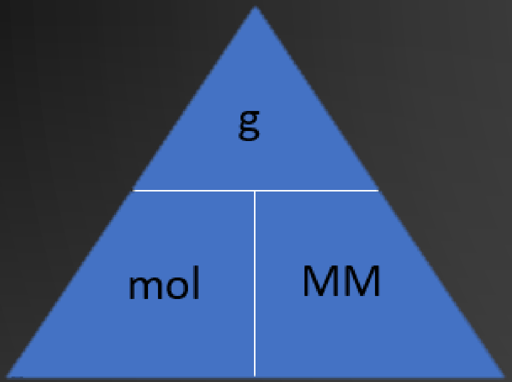
$$2\,Na_3PO_{4(aq)} + 3\,CaCl_{2(aq)} \longrightarrow 6\,NaCl_{(aq)} + Ca_3(PO_4)_{2(s)}$$

CALCULATIONS USING CHEMICAL EQUATIONS

- Chemists need to be able to calculate the mass of each reactant required to produce a certain amount of a product.
- Example: $4\ Na_{(s)} + O_{2(g)} \longrightarrow 2\ Na_2O_{(s)}$

Na: 22.99
O_2: 32.00
Na_2O: 61.98

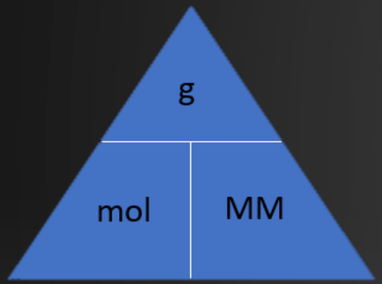
- How much sodium reacts with 3.20g of O_2? What mass of $Na_2O_{(s)}$ is produced?



$$2\,NH_3 + 3CuO \longrightarrow N_2 + 3\,H_2O + 3Cu$$

17.04g/mol — 63.55g/mol

If 2.56g of NH_3 is reacted with excess CuO, calculate the mass of Cu produced



- An alternative method for solving these types of problems

$$\frac{m_1}{n_1 M_1} = \frac{m_2}{n_2 M_2}$$

m_1 = mass of first substance (g)
n_1 = coefficient of first substance
M_1 = molar mass of first substance

$$2\,C_4H_{10(g)} + 13\,O_{2(g)} \longrightarrow 8\,CO_{2(g)} + 10\,H_2O_{(l)}$$

58.14	32	44.01	18.02

If 10.00g of butane is used, calculate:

- The moles of oxygen required
- The mass of CO_2 produced

PRACTICE PROBLEMS

1) How many moles of hydrogen gas are produced when 0.4 moles of sodium react with excess water?

$$Na + H_2O \longrightarrow NaOH + H_2$$

Na: 22.99
H_2: 2.02
H_2O: 18.02
O_2: 32.00
C_3H_8: 44.11

2) How many moles of O_2 react with 0.01mol C_3H_8?

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

3) Calculate the mass of arsenic(III) chloride produced when 0.150g of arsenic reacts with excess chlorine.

$$As \quad + \quad Cl_2 \longrightarrow AsCl_3$$

As: 74.92
Cl_2: 70.91
$AsCl_3$: 181.28
SCl_2: 102.97
NaCl: 58.44

4) What mass of SCl_2 must be reacted with excess NaF to produce 2.25g of NaCl?

$$SCl_2 \quad + \quad NaF \longrightarrow S_2Cl_2 \quad + \quad SF_4 \quad + \quad NaCl$$

CALCULATING THE YIELD OF A CHEMICAL REACTION

- It is important to know the yield of a chemical reaction, for instance there may be two possible methods for synthesising a compound.
  - Route 1 = 95% yield Route 2 = 88% yield

- Yield is usually quoted as a percentage

$$\frac{Actual\ Yield}{Theoretical\ Yield} \text{ X } 100 = \%\ \text{Yield}$$

The theoretical yield will be calculated using coefficient ratios and masses/moles

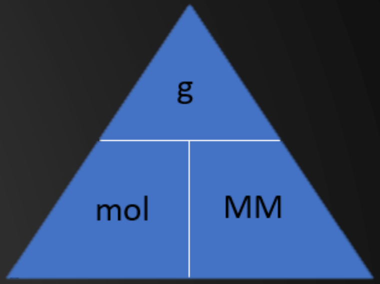
- If the yield of ethyl ethanoate obtained, when 20.0g of ethanol is reacted with excess ethanoic acid, is 30.27g. Calculate the % yield.

$$C_2H_5OH_{(l)} + CH_3COOH_{(l)} \longrightarrow CH_3COOC_2H_{5(l)} + H_2O$$

ethanol 46.08g/mol

ethanoic acid

ethyl ethanoate 88.12g/mol



$$\frac{Actual\ Yield}{Theoretical\ Yield} \text{ X } 100 = \%\ \text{Yield}$$

QUIZ

- Empirical and molecular formulas
- Combustion analysis
- Ionic equations
- Calculations using mass and moles
- Yield

LIMITING REACTANTS

- In most chemical reactions exact quantities are not used, instead an excess of one or more reagents is used.
- Therefore, one reactant is consumed before the others and is called the limiting reactant (also known as the limiting reagent)

Determine the limiting reagent if you have the following:

14 slice of bread 18 slices of Ham 20 slices of cheese

$$2Al_{(s)} + 3Cl_{2(g)} \longrightarrow 2AlCl_{3(s)}$$

26.98g/mol 70.9g/mol

114g of Al is reacted with 186g of $Cl_{2.}$ Which is the limiting reactant?

Step 1 – Find moles

Al = 114/26.98 = 4.23 moles

Cl_2 = 186/70.9 = 2.62 moles

Step 2 –

To use all Al we need $4.23 \times \frac{3}{2}$ = 6.35 mol Cl_2

To use all Cl_2 we need $2.62 \times \frac{2}{3}$ = 1.75 mol Al

We have enough Al to use all the $Cl_{2.}$ = Excess

We don't have enough Cl_2 to use all the Al. = Limiting Reagent

Identify the limiting reactant in each of the following reactions:

0.1mol Sb_4O_6 reacts with 0.5mol H_2SO_4

$$Sb_4O_6 + H_2SO_4 \longrightarrow Sb_2(SO_4)_3 + H_2O$$

291.52 98.08

0.20mol $AsCl_3$ reacts with 0.25mol H_2O

$$AsCl_3 + H_2O \longrightarrow As_4O_6 + HCl$$

181.28 18.02

2.3kg of S_8 is reacted with 3.2kg of O_2

$$S_8 + 12\,O_2 \longrightarrow 8\,SO_3$$

256.52g/mol 32.00g/mol

- Determine the limiting reagents in each of the following reactions:

$AlBr_3$	+	K_2SO_4	$\rightarrow$	KBr	+	$Al_2(SO_4)_3$
266.69g/mol		174.26g/mol				
465g		315g				

N_2	+	H_2	$\rightarrow$	NH_3
28.02g/mol		2.02g/mol		
41g		9.0g		

IDEAL GASES

- Gases are complicated. They're full of billions and billions of energetic gas molecules that can collide and possibly interact with each other.
- Since it's hard to exactly describe a real gas, people created the concept of an Ideal gas as an approximation that helps us model and predict the behaviour of real gases.
- The term ideal gas refers to a <u>hypothetical</u> gas composed of molecules which follow these rules:

Ideal gas molecules do not attract or repel each other. The only interaction between ideal gas molecules would be an elastic collision upon impact with each other or an elastic collision with the walls of the container.

Ideal gas molecules themselves take up no volume. The gas takes up volume since the molecules expand into a large region of space, but the Ideal gas molecules are approximated as point particles that have no volume in and of themselves.

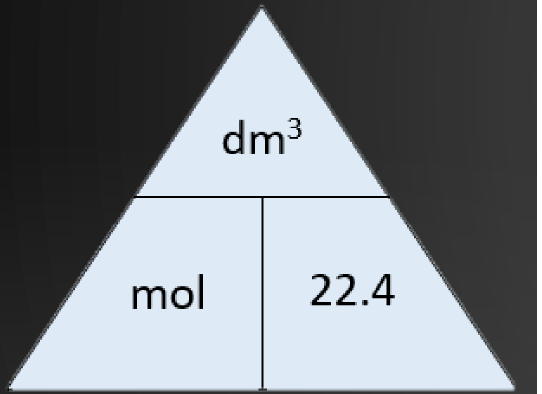
A collision wherein no kinetic energy is converted to other forms of energy during the collision. In other words, kinetic energy can be exchanged between the colliding objects (e.g. molecules), but the total kinetic energy before the collision is equal to the total kinetic energy after the collision

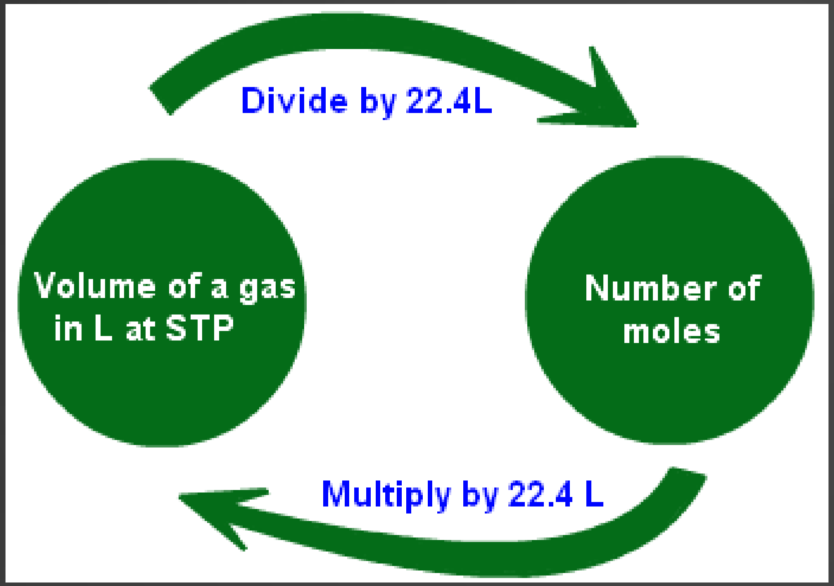
- There are no gases that are exactly *ideal*, but there are plenty of gases that are close enough that the concept of an ideal gas is an extremely useful approximation for many situations.
- In fact, for temperatures near room temperature and pressures near atmospheric pressure, many of the gases we care about are very nearly ideal.
- If the pressure of the gas is too large (e.g. hundreds of times larger than atmospheric pressure), or the temperature is too low (e.g. -200°C) there can be significant deviations from the ideal gas law.

CALCULATIONS INVOLVING GASES

Standard temperature and pressure (STP) = 0°C (273k), 1 atm (100kPa)

The volume of a gas at STP is constant : 1 mole = 22.4 $dm^3\ mol^{-1}$





Avogadros Law

Volumes of gases are often given in dm^3 (litres) or cm^3

1 dm^3 = 1000 cm^3

Calculate the mass of 250cm^3 of O_2 at STP

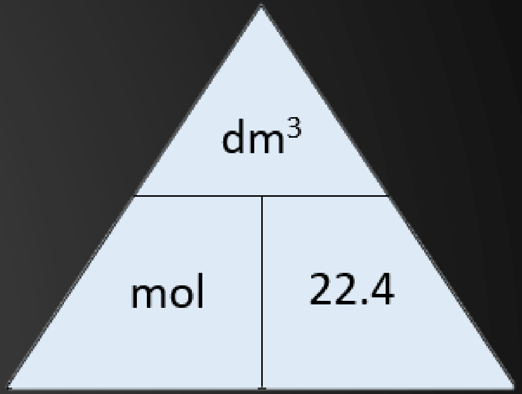
250cm^3 = 0.25dm^3
0.25 / 22.4 = 0.0112 mol
0.0112 x 32 = 0.36g

$1\ dm^3 = 1000\ cm^3$

$$2\ KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$$

122.55g/mol

32g/mol



What mass of $KClO_3$ decomposes to produce $100cm^3$ of O_2 at STP?

ALTERNATIVE FORMULAS

$$\frac{m_1}{n_1 M_1} = \frac{V_2}{n_2 M_v}$$

When a mass is given and the volume of another substance is required

m_1 = mass of first substance (g)
n_1 = coefficient of first substance
M_1 = molar mass of first substance
V_2 = volume of second substance if it is a gas (dm^3)
n_2 = coefficient of second substance
M_v = molar volume of a gas ($22.4 dm^3$ at STP)

$$2As_2S_{3(s)} + 9O_{2(g)} \longrightarrow 2As_2O_{3(s)} + 6SO_{2(g)}$$

246.02g/mol

64.07g/mol

What volume of SO_2 (measured at STP) is obtained when 1kg of As_2S_3 is combusted?

Substance 1 = As_2S_3

$m_1 =$

$n_1 =$

$M_1 =$

Substance 2 = SO_2

$V_2 =$

$n_2 =$

$M_v =$

$$\frac{m_1}{n_1 M_1} = \frac{V_2}{n_2 M_v}$$

$$4NH_{3(g)} + 5O_{2(g)} \longrightarrow 4NO_{(g)} + 6H_2O_{(g)}$$

Ar:

N: 14.01
H: 1.01
O: 16.00

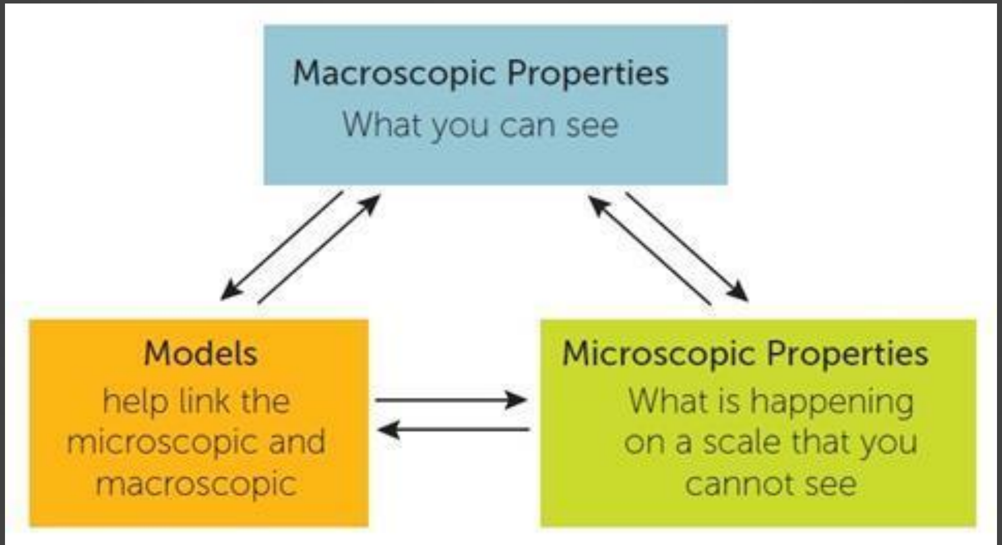
Calculate the volume of $NO_{(g)}$ and $H_2O_{(g)}$ produced when 0.6dm^3 of $O_{2(g)}$ is reacted with excess ammonia.

MACROSCOPIC PROPERTIES OF IDEAL GASES

Macroscopic means 'on a large scale'

- Microscopic properties of gases are the properties of the particles that make up the gas

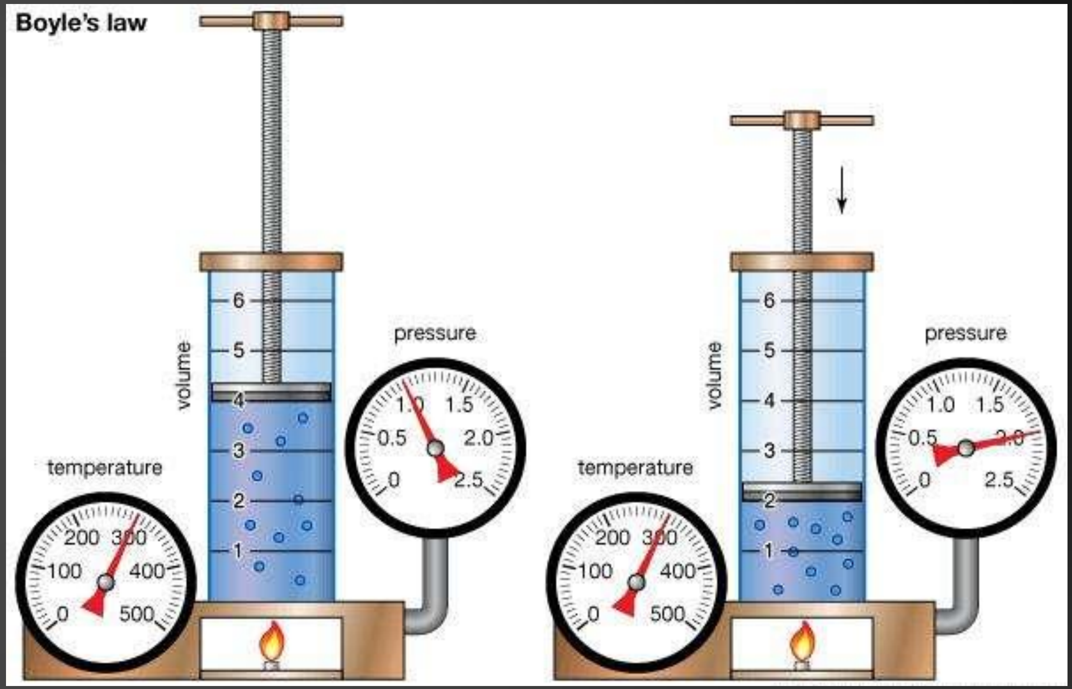
- So far we have dealt with questions involving the volumes of gases at STP. In order to work out volumes of gases under other conditions, we must understand a little about the properties of gases.



BOYLE'S LAW

The relationship between pressure and volume

$$P_1V_1 = P_2V_2$$



At a constant temperature, the volume of a fixed mass of an ideal gas is inversely proportional to its pressure

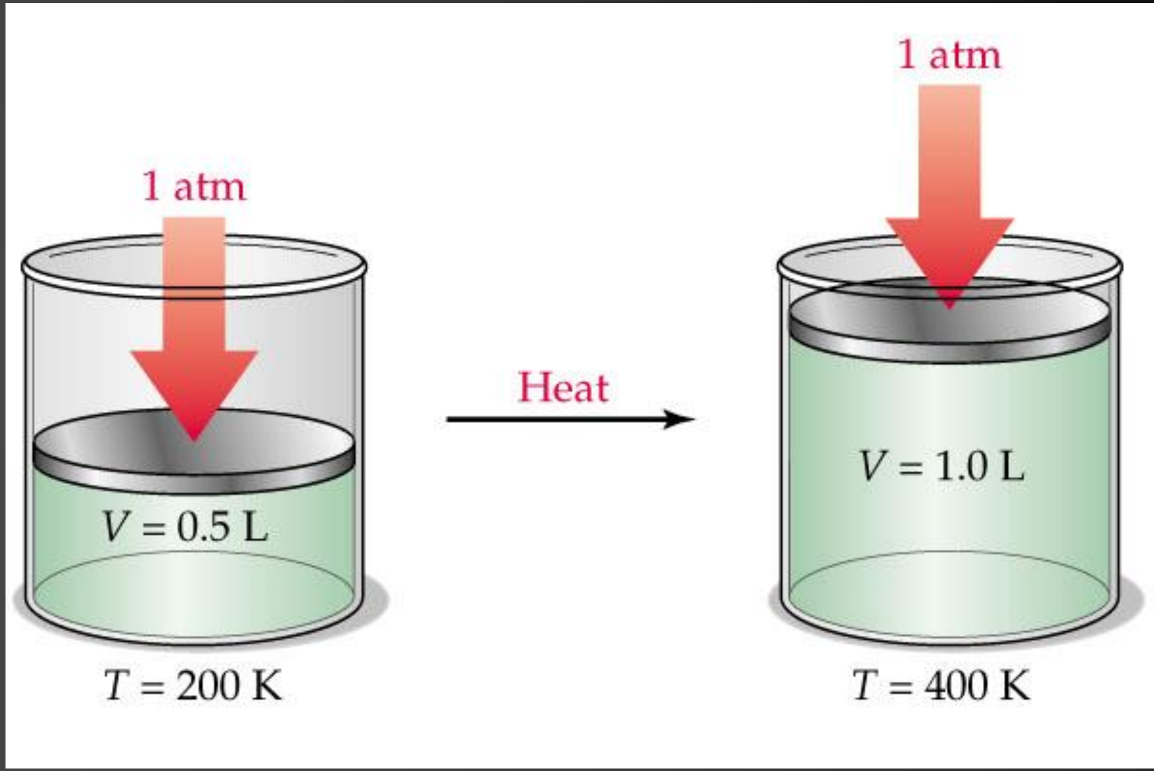
- Example – if the pressure of a gas is doubled at constant temperature, then the volume will be halved and vice versa (double volume = halve the pressure)

CHARLES' LAW

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

The relationship between volume and temperature

The volume of a fixed mass of an ideal gas at constant pressure is directly proportional to its temperature in kelvin.



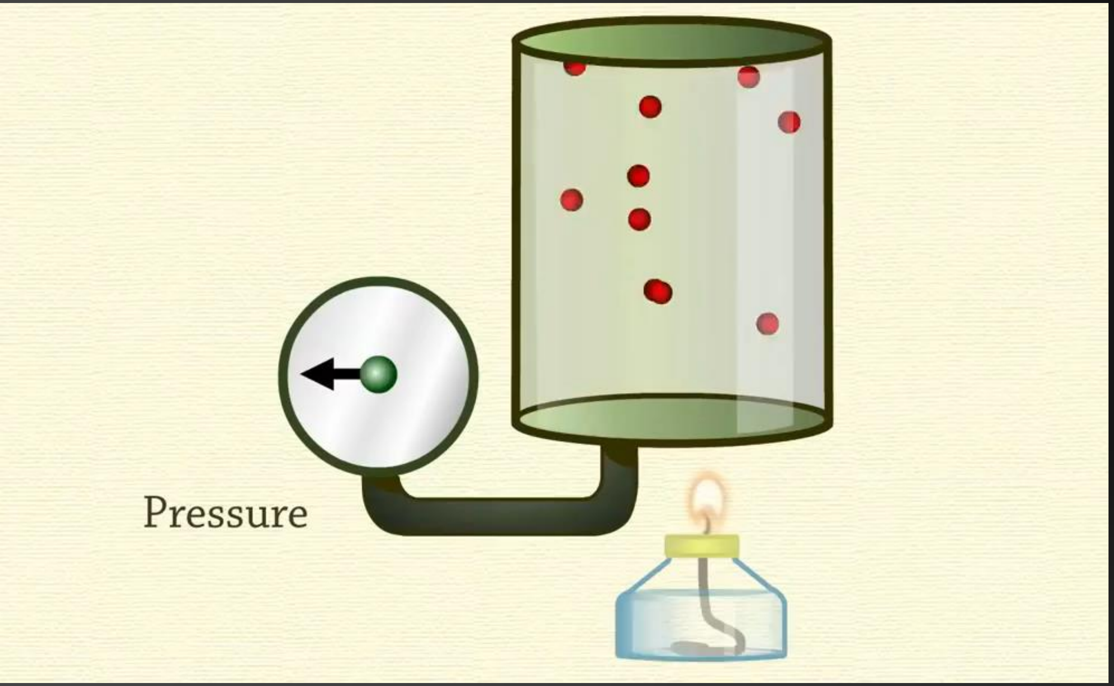
- Example – If an ideal gas has a volume of 200cm^3 at 120K, it will have a volume of 400cm^3 at 240K if the pressure remains constant

GAY-LUSSAC'S LAW

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

The relationship between pressure and temperature

For a fixed mass of an ideal gas at constant volume, the pressure is directly proportional to its absolute temperature (K).



- Example – If the temperature (in Kelvin) of a fixed volume of an ideal gas is doubled, the pressure will also double

- Which gas law requires constant volume?

  a) Boyle's Law
  b) Charles' law
  c) Gay-Lussac's law

- Which gas law requires constant temperature?

  a) Boyle's Law
  b) Charles' law
  c) Gay-Lussac's law

- Which gas law requires constant pressure?

  a) Boyle's Law
  b) Charles' law
  c) Gay-Lussac's law

- Which gas law states that at a constant temperature, the volume of a fixed mass of an ideal gas is inversely proportional to its pressure?

  a) Boyle's Law
  b) Charles' law
  c) Gay-Lussac's law

- Which gas law states that when a fixed mass of an ideal gas has a constant volume, the pressure is directly proportional to its absolute temperature (K)?

  a) Boyle's Law
  b) Charles' law
  c) Gay-Lussac's law

- Which gas law states that the volume of a fixed mass of an ideal gas at constant pressure is directly proportional to its temperature in kelvin?

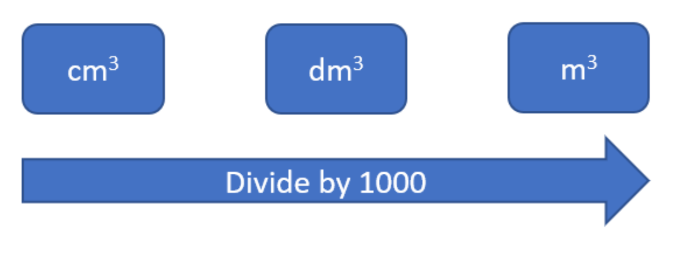
  a) Boyle's Law
  b) Charles' law
  c) Gay-Lussac's law

COMBINED GAS LAW

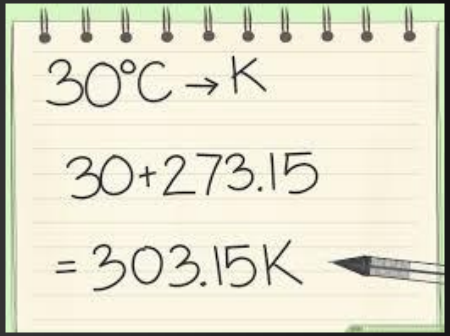
Boyle's law, Charles' law and Gay-Lussac's law can be combined to produce the following equation:

COMBINED GAS LAW

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$



Temperature must be in Kelvin



Any units may be used for P and V, as long as they are the same on both sides

1 atm = 101.32 kPa = 760mmHg = 1.013 bar = 14.7 PSI

- Rearranged to solve for each variable:
  - $P_1=(P_2V_2T_1)/(V_1T_2)$
  - $V_1=(P_2V_2T_1)/(P_1T_2)$
  - $T_1=(P_1V_1T_2)/(P_2V_2)$
  - $P_2=(P_1V_1T_2)/(V_2T_1)$
  - $V_2=(P_1V_1T_2)/(P_2T_1)$
  - $T_2=(P_2V_2T_1)/(P_1V_1)$

What temperature in °C is required for an ideal gas to occupy $1.34dm^3$ at a pressure of 2.05atm if it occupies $756cm^3$ at STP?

P_1 = 2.05atm
V_1 = $1.34dm^3$
T_1 = ?

P_2 = 1.00atm
V_2 = $756cm^3$ = (756/1000) = $0.756dm^3$
T_2 = 273K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{2.05 \times 1.34}{T_1} = \frac{1.00 \times 0.756}{273}$$

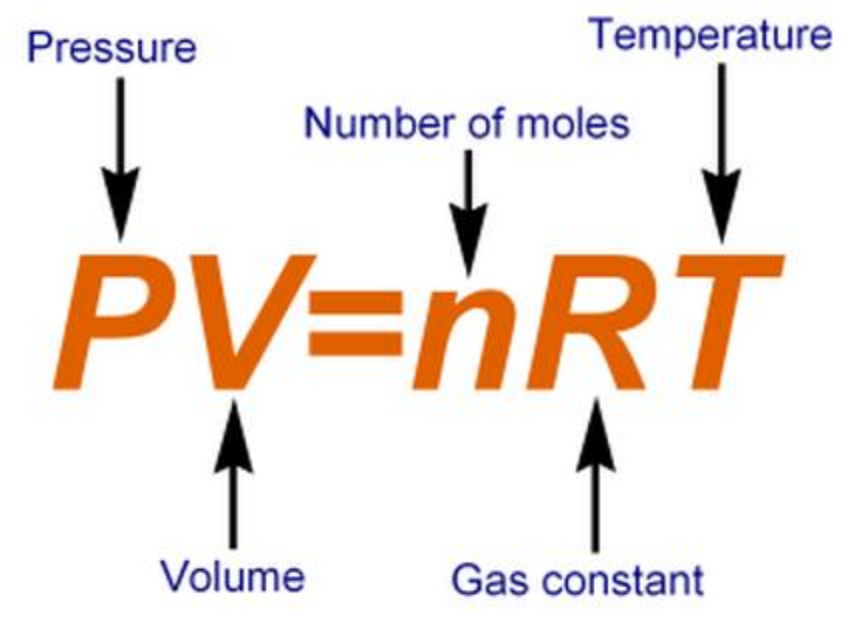
$$T_1 = \frac{2.05 \times 1.34 \times 273}{1.00 \times 0.756} = 992\ K$$

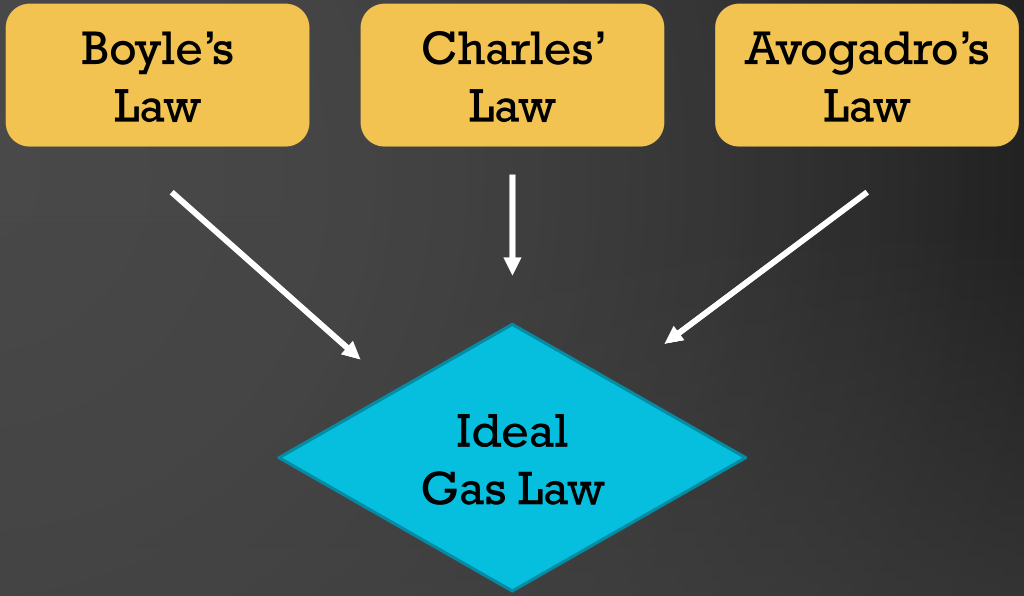
992 – 273 = 719°C

IDEAL GAS LAW

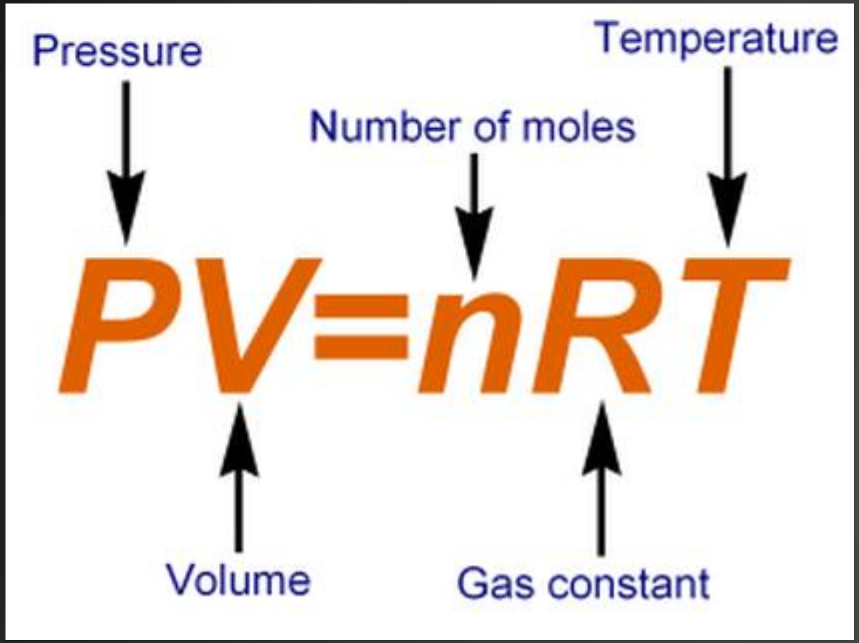
1 mole of any gas
= 22.4L at STP

If the relationship between P, V and T are combined with Avogadro's law the ideal gas equation is obtained.





UNITS



Pressure Units				
	1 atm	1 mmHg	1 PSI	1 Bar
Pa	101,325	133.32	6894.76	100,000

$1m^3 = 1000\ dm^3 = 1,000,000\ cm^3$

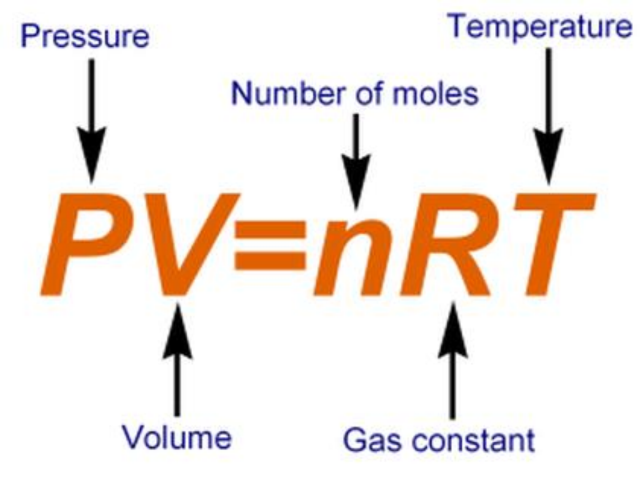
P = Pa
Volume = m^3
R = $8.31\ J\ K^{-1}\ mol^{-1}$
T = K
$1.0x10^{-5}$

What is 800mmHg in Pa?

What is 18PSI in Pa?

What is 6atm in Pa?

An ideal gas occupies $590 cm^3$ at 120°C and 2.00 atm.
What amount of gas (in moles) is present?

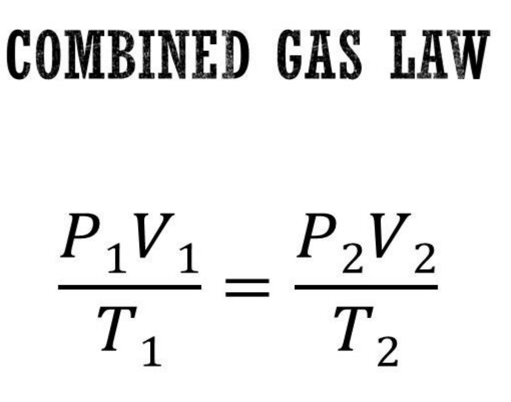


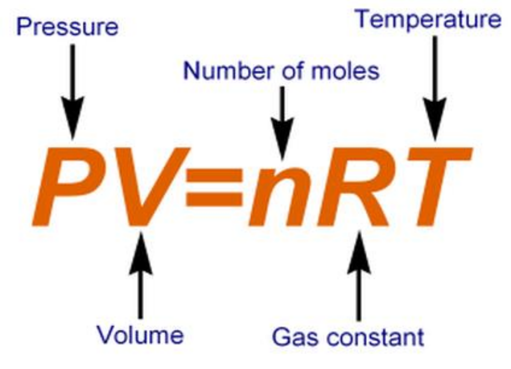
P = Pa
Volume = m^3
R = 8.31 J K^{-1} mol^{-1}
T = K

Pressure Units				
	1 atm	**1 mmHg**	**1 PSI**	**1 Bar**
Pa	101,325	133.32	6894.76	100,000

$1 m^3 = 1000 dm^3 = 1,000,000 cm^3$

COMBINED GAS LAW VS IDEAL GAS LAW





- When should the combined gas law be used?
  - The combined gas law is useful when given two pressures, volumes, or temperatures and asked for an unknown pressure, volume, or temperature. A change in conditions
- When should the ideal gas law be used?
  - The ideal gas law *does not require a change in the conditions of a gas sample*. The ideal gas law implies that if you know any three of the physical properties of a gas, you can calculate the fourth property.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

PRACTICE QUESTIONS

$$PV = nRT$$

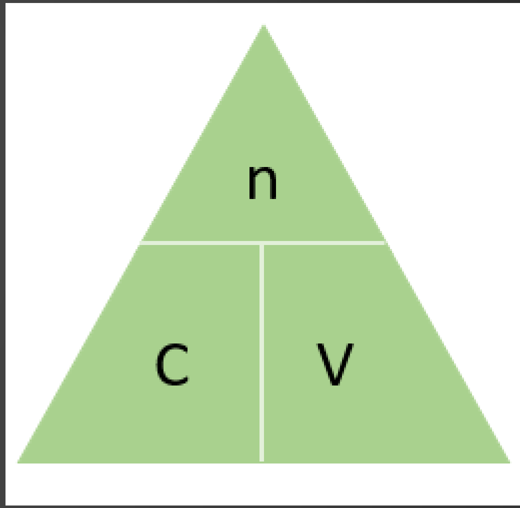
1) If a certain mass of an ideal gas occupies $20cm^3$ at STP, what volume would it occupy at 38°C and $1.06x10^5$ Pa?

2) How many moles of an ideal gas are present in a container if it occupies $1.50dm^3$ at a pressure of 45 PSI and a temperature of 30°C?

3) Calculate the molar mass of an ideal gas if 0.586g of the gas occupies a volume of $282cm^3$ at a pressure of $1.02x10^5$ Pa and a temperature of -18°C.

CALCULATIONS INVOLVING SOLUTIONS

- A solution is a homogenous mixture; containing a solvent and at least one solute
- Solutions in water are given the symbol (aq); in chemical equations aq = aqueous
- The concentration of a solution is the amount of solute dissolved in a unit volume of solution
  - Volume is usually dm^3
  - The amount of solute is usually in grams or moles
  - Therefore, units are $g\ dm^{-3}$ or $mol\ dm^{-3}$
  - $mol\ dm^{-3}$ is also known as **molarity** (M)



Moles can be replaced with mass to calculate concentration in $g\ dm^{-3}$

n = Number of moles
C = Concentration (M)
V = Volume (dm^3)

WORKED EXAMPLES

If 10.00g of NaOH is dissolved in water and the volume is made up to 200ml, calculate the concentration in g dm^{-3} and mol dm^{-3}.

Concentration (g dm^{-3})

$$C = \frac{M}{V}$$

Concentration (mol dm^{-3})

$$C = \frac{n}{V}$$

Calculate the number of moles of HCl present in $50.0cm^3$ of 2.0M hydrochloric acid

How many grams of NaCl are needed to prepare 1.5 L of a 0.20 M solution?

CONCENTRATION OF IONS

$$C = \frac{n}{V}$$

When ionic substances dissolve in water, the substance breaks apart into constituent ions.

$$CuCl_{2(aq)} \rightarrow Cu^{2+}_{(aq)} + 2\ Cl^{-}_{(aq)}$$

Therefore when 0.100 mol $CuCl_2$ dissolves in water, 0.200 mol of Cl^- is produced.

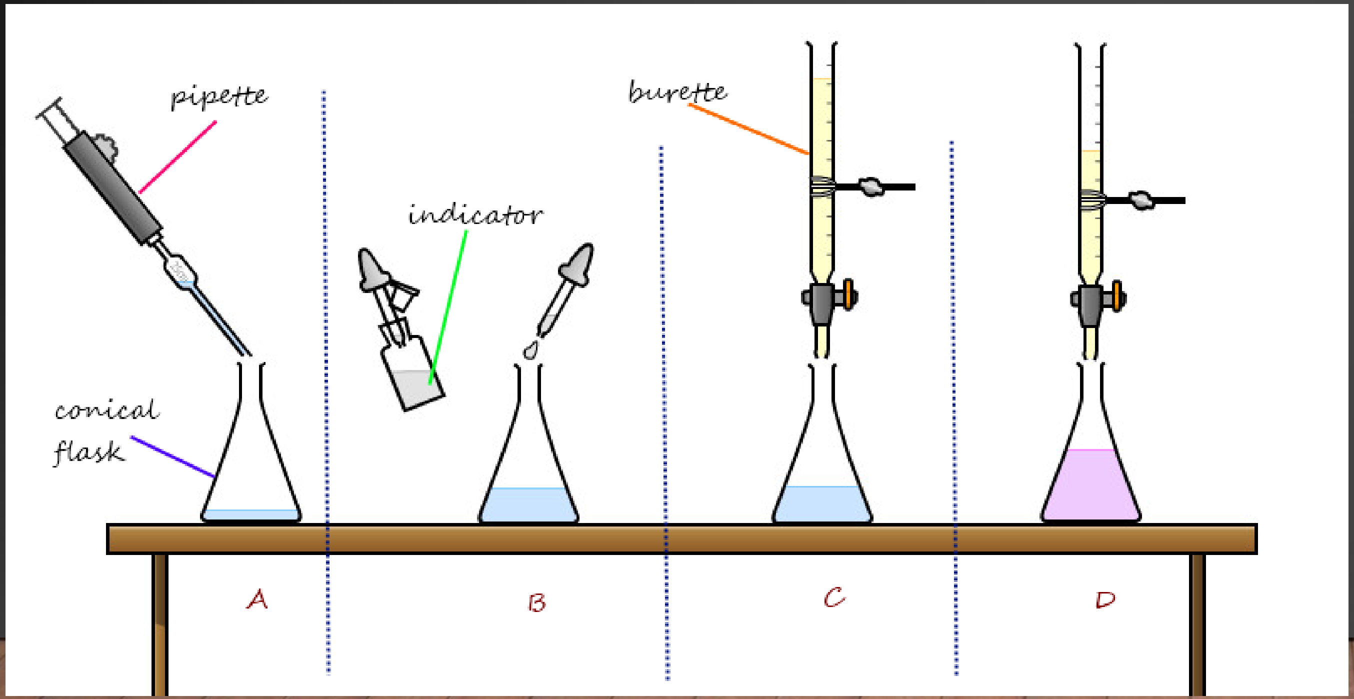
- The [chloride ion], is twice the [$CuCl_2$]

Calculate [OH^-], when 0.500 moles of $Al(OH)_3$ is dissolved in 300ml of water.

$$Al(OH)_{3(aq)} \rightarrow Al^{3+}_{(aq)} + 3\ OH^{-}_{(aq)}$$

TITRATIONS

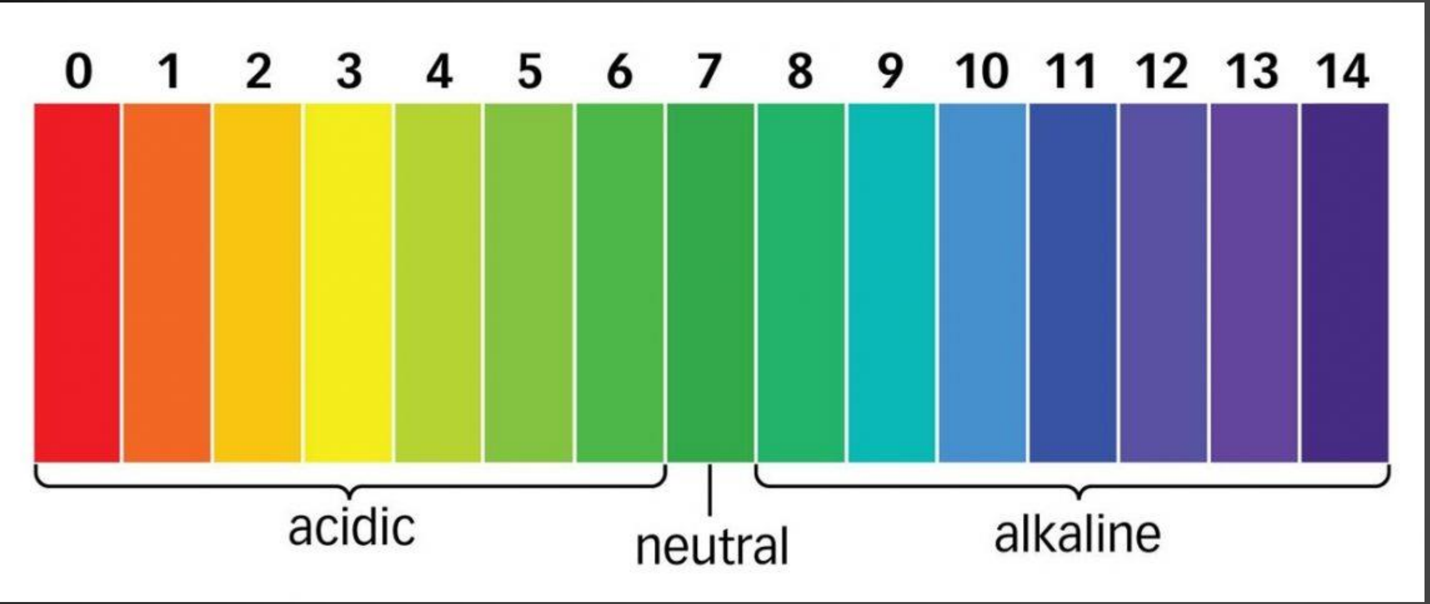
- A titration is a technique where a solution of known concentration is used to determine the concentration of an unknown solution.
- Typically, the titrant (the known solution) is added from a burette to a known quantity of the analyte (the unknown solution) until the reaction is complete.
- The unknown solution is often dispensed using a pipette
- Often, an indicator is used to signal the end of the reaction, the endpoint.

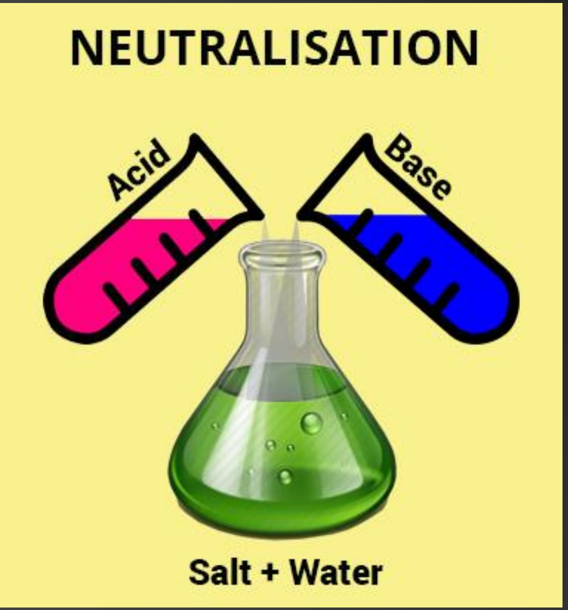


A – Analyte is dispensed into a conical flask using a pipette
B – An indicator is added
C – A titrant is added from a burette until the end point is reached.
D – The volume of the titrant (known concentration) required to reach the end point is used to determine the concentration of the analyte

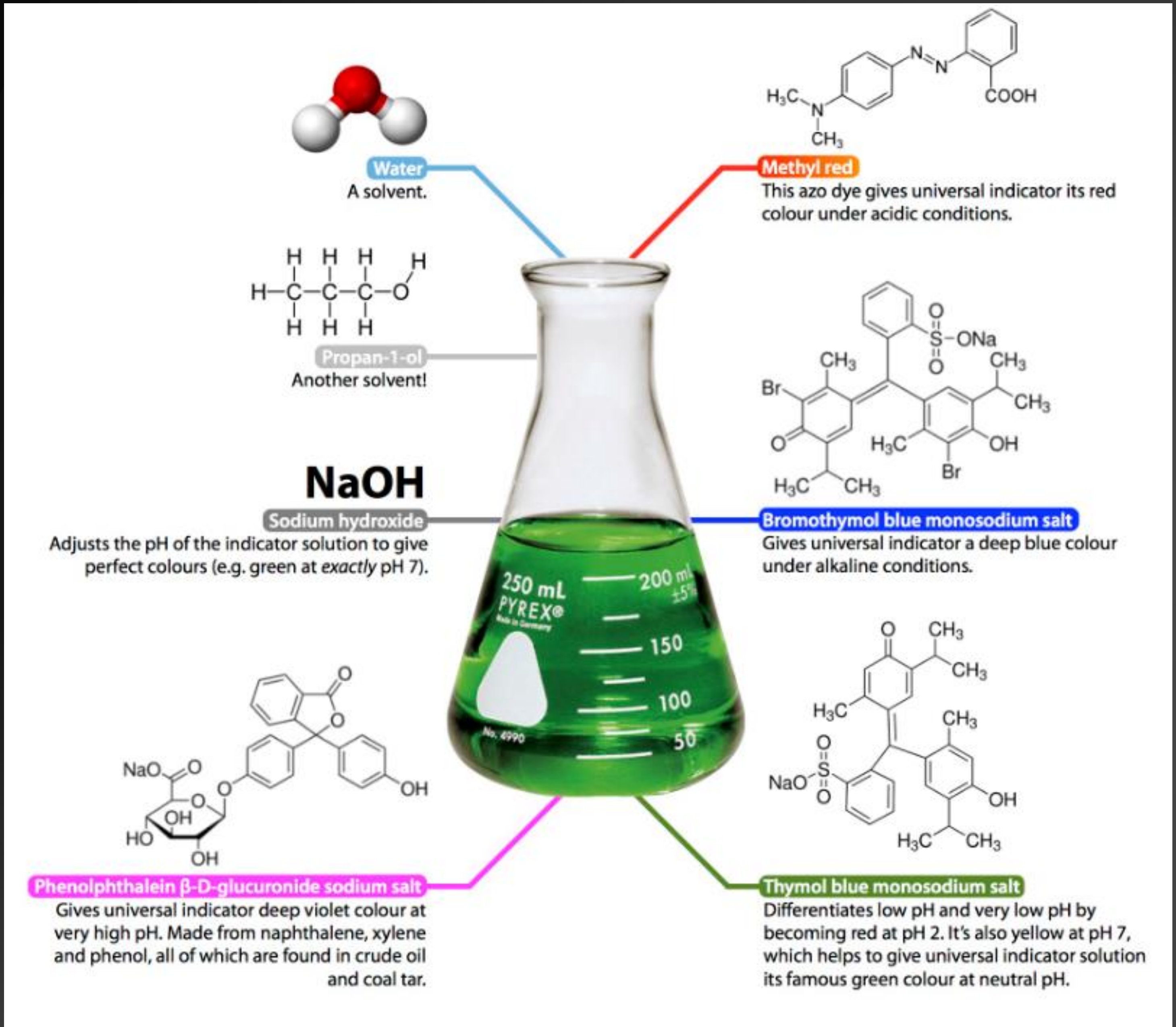
- A common titration is an acid-base titration
- Often the end point is neutralisation

$$\underset{\text{acid}}{H^+} + \underset{\text{base}}{OH^-} \rightarrow \underset{\text{neutral}}{H_2O}$$

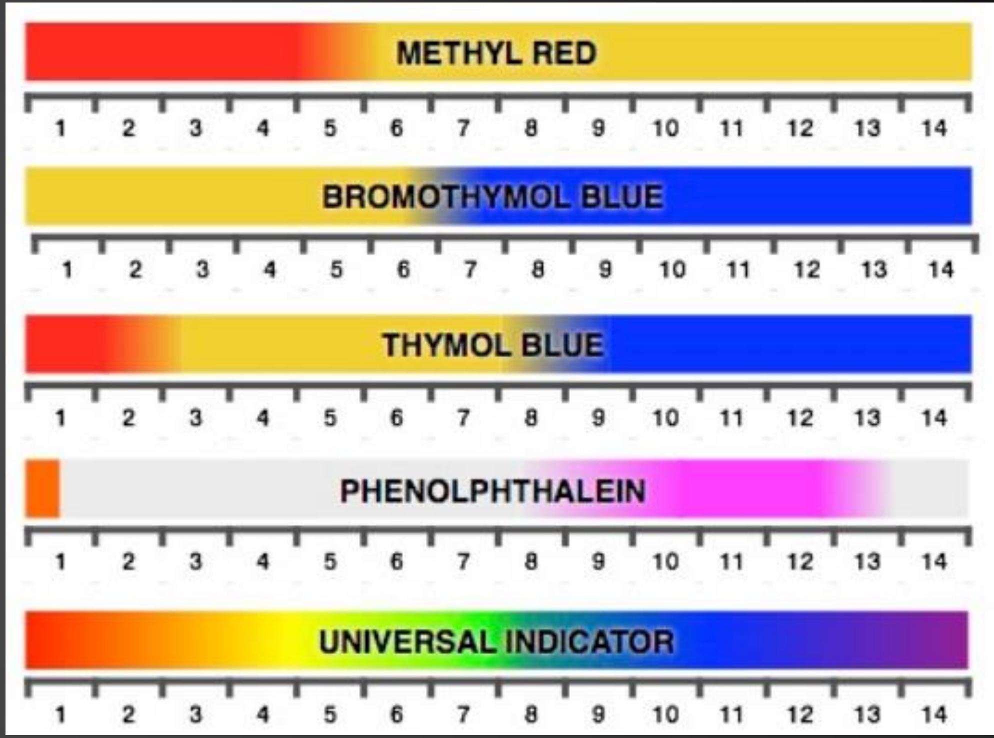




Excess H^+ $[H^+] = [OH^-]$ Excess OH^-



Water
A solvent.
Methyl red
This azo dye gives universal indicator its red colour under acidic conditions.
Propan-1-ol
Another solvent!
NaOH
Sodium hydroxide
Adjusts the pH of the indicator solution to give perfect colours (e.g. green at exactly pH 7).
Bromothymol blue monosodium salt
Gives universal indicator a deep blue colour under alkaline conditions.
250 mL
PYREX
200 ml
150
100
50
Phenolphthalein β-D-glucuronide sodium salt
Gives universal indicator deep violet colour at very high pH. Made from naphthalene, xylene and phenol, all of which are found in crude oil and coal tar.
Thymol blue monosodium salt
Differentiates low pH and very low pH by becoming red at pH 2. It's also yellow at pH 7, which helps to give universal indicator solution its famous green colour at neutral pH.



METHYL RED
1 2 3 4 5 6 7 8 9 10 11 12 13 14
BROMOTHYMOL BLUE
1 2 3 4 5 6 7 8 9 10 11 12 13 14
THYMOL BLUE
1 2 3 4 5 6 7 8 9 10 11 12 13 14
PHENOLPHTHALEIN
1 2 3 4 5 6 7 8 9 10 11 12 13 14
UNIVERSAL INDICATOR
1 2 3 4 5 6 7 8 9 10 11 12 13 14

TITRATION CALCULATIONS

$$2\,NaOH_{(aq)} + H_2SO_{4(aq)} \longrightarrow Na_2SO_{4(aq)} + 2\,H_2O_{(l)}$$

NaOH	H_2SO_4	Na_2SO_4
25ml	23.2ml	Salt
0.200M		
Titrant	Analyte	

Calculate the concentration of the sulphuric acid (H_2SO_4).

$$\frac{C_1 \times V_1}{n_1} = \frac{C_2 \times V_2}{n_2}$$

C = concentration V = volume n = coefficient

- 1) If it takes 54ml of 0.1M $Mg(OH)_2$ to neutralise 125ml of a HCl solution, what is the concentration of the HCl?

- 2) If it takes 250ml 1.2M HCl to neutralise 1.0L of an $Al(OH)_3$ solution, what is the concentration of the $Al(OH)_3$ solution?

- 3) If it takes 50ml of 0.5M KOH solution to completely neutralise 125ml of sulfuric acid solution (H_2SO_4), what is the concentration of the H_2SO_4 solution?

BACK TITRATION

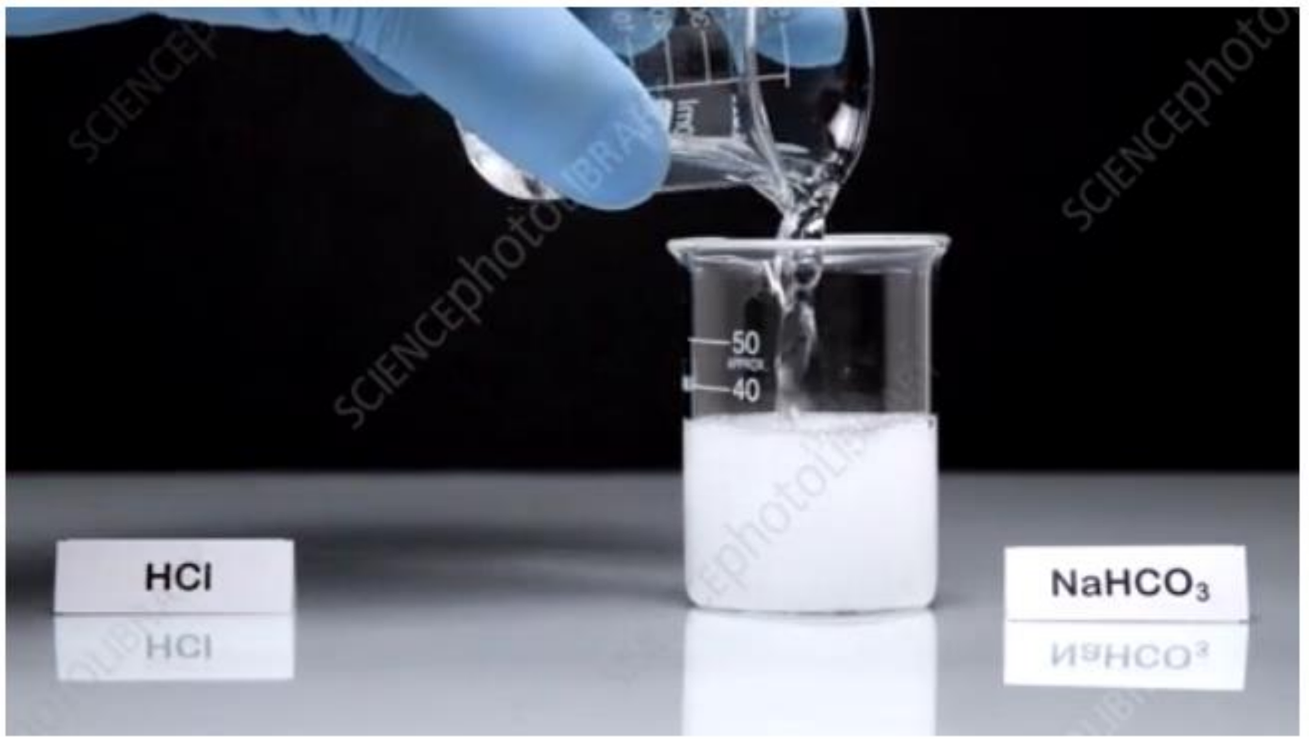
- A technique by which a known excess of a particular reagent (A) is added to a substance (X), so that they react, and then the excess A is titrated against another reagent.
- This enables us to work out how much of reagent A reacted with substance X and therefore how many moles of X were present.
- This technique is useful when X is an impure substance.

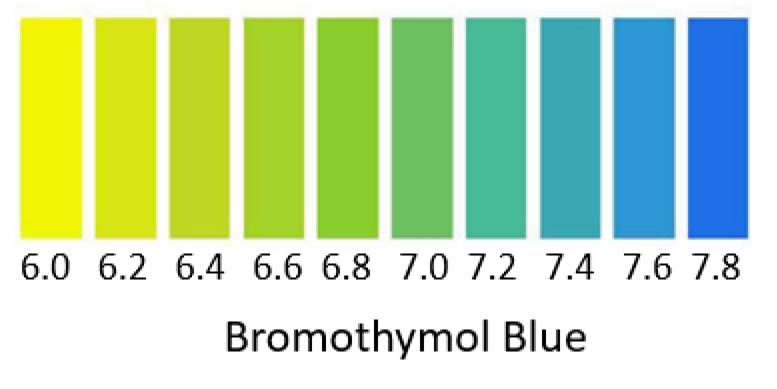
Measure the mass of an impure substance containing sodium bicarbonate ($NaHCO_3$)

The crushed tablet is added to 100ml of 0.1M hydrochloric acid

$$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$$

The sodium bicarbonate neutralises some of the HCl. Now we have an unknown concentration of HCl





NaOH Added

Unknown concentration of HCl

The volume of NaOH needed to neutralise the HCl is used to determine the concentration of the HCl solution.

0.2g tablet containing $NaHCO_3$

Added to 100ml 0.1M HCl

$$0.1 \times \frac{100}{1000} = 0.01 \text{ moles}$$

Indicator added

17ml 0.5M NaOH

$$0.5 \times \frac{17}{1000} = 0.0085 \text{ moles}$$

Moles HCl – Moles NaOH = moles of HCl neutralised by $NaHCO_3$

0.01 0.0085 0.0015

$$NaHCO_3 + HCl \longrightarrow NaCl + H_2O + CO_2$$

1:1

0.0015 moles $NaHCO_3$ (As HCl and $NaHCO_3$ react in a 1:1 ratio)

Moles $NaHCO_3$ x Molar mass $NaHCO_3$ = grams of $NaHCO_3$

0.0015 84.007 0.126

If 0.126g of the 0.2g tablet is $NaHCO_3$; percentage purity = 63%

150ml of 0.2105M nitric acid (HNO_3) was added to 1.3415g of impure $CaCO_3$ (100.09g/mol). The excess acid was back titrated with 0.1055M NaOH, it required 75.5ml to reach the end point. Calculate the % mass of $CaCO_3$ in the sample.

$CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + CO_2 + H_2O$

$HNO_3 + NaOH \rightarrow H_2O + NaNO_3$

4.06g of impure magnesium oxide (MgO, molar mass = 40.30) was completely dissolved in 100ml of 2.0M HCl (in excess). The excess acid required 19.7ml of 0.20M NaOH for neutralisation. Calculate the % purity of the magnesium oxide.

$MgO + 2HCl \rightarrow MgCl_2 + H_2O$

A 1.435g sample of dry $CaCO_3$ and $CaCl_2$ mixture was dissolved in 25.00ml of 0.9892M HCl. What was the $CaCl_2$ percentage in original sample, if 21.48ml of 0.09312M NaOH was used to titrate excess HCl?

(molar mass of $CaCO_3$ = 100.09, molar mass of $CaCl_2$ = 110.98)

$CaCO_3 + CaCl_2 + 2HCl \rightarrow 2CaCl_2 + CO_2 + H_2O$ The $CaCl_2$ does not react

Difficult Problem

A 2.75g sample of dolomite containing $CaCO_3$ and $MgCO_3$ is dissolved in 80ml of 1.0M HCl. The solution is then diluted to 250ml. 25ml of this solution requires 20ml of 0.1M NaOH solution for complete neutralisation. Calculate the % composition of the sample.

(molar mass of $CaCO_3$ = 100.09, molar mass of $MgCO_3$ = 84.31)

$CaCO_3 + MgCO_3 + 4HCl \rightarrow CaCl_2 + MgCl_2 + 2H_2O + 2CO_2$

2.64g of an unknown carbonate (MCO_3) was added to 50ml of 2M HCl. The solution was then made up to 250ml. A 25ml aliquot was neutralised by 37.15ml of 0.1M NaOH.

(molar mass of CO_3 = 60)

$HCl + NaOH \rightarrow NaCl + H_2O$

$MCO_3 + 2HCl \rightarrow MCl_2 + H_2O + CO_2$

What is the unknown metal in the carbonate?

The following website contains worked examples of different types of back titration calculations:

https://www.ibchem.com/IB16/03.55.htm

LAB

- In groups (up to 5) you will complete a back titration before completing the analysis using the mathematical concepts shown on the previous slides.
- This lab will be assessed and will count towards your collective score
- The calculations and analysis questions must be written up and submitted.
- Accuracy is important. Your results will influence the calculated mass of $CaCO_3$