Chemistry Notes Module 2 – Introduction to Quantitative Chemistry

Chemical Reactions and Stoichiometry

Inquiry question: What happens in chemical reactions?

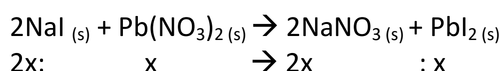
Conduct practical investigations to observe and measure the quantitative relationships of chemical reactions, including but not limited to:

- masses of solids and/or liquids in chemical reactions
- volumes of gases in chemical reactions

Observing and measuring quantitative relationships in chemical reactions

Masses of solids and/or liquids in chemical reactions:

Sodium iodide and Lead (II) nitrate



$2NaI_{(s)} + Pb(NO_3)_{2\,(s)} \rightarrow 2NaNO_{3\,(s)} + PbI_{2\,(s)}$

2x: x → 2x : x

n(NaI) = 15/149.89 = 0.10 (2 sig. fig.)

$n(Pb(NO_3)_2)$ = 16.5/331.22 = 0.050 (2 sig. fig)

- the molar calculations reflect the stoichiometric ratio as the NaI has double the amount of moles to $Pb(NO_3)_2$

$m(PbI_2) = n(PbI_2) \times MM(PbI_2)$
= 0.05 x (207.2 +(2x126.9))
= 23.0g

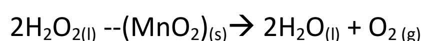
$m(NaNO_3) = n(NaNO_3) \times MM(NaNO_3)$
= 0.10 x (22.99 +14.01 + (3x16.00))
= 8.5g

Sum(reactants) = 15g of NaI + 16.5g of $Pb(NO_3)_2$ = 31.5g

Sum(products) = 23.0g of PbI_2 + 8.5g of $NaNO_3$ = 31.5g

- the law of conservation of mass has been observed as the sum of the mass of the reactants is equal to the sum of the mass of the products

Hydrogen peroxide and Manganese dioxide



$2H_2O_{2(l)}$ --$(MnO_2)_{(s)}$→ $2H_2O_{(l)} + O_{2\,(g)}$

$m(H_2O_2)$ = 138.99g

- the mass of MnO_2 is not important as it is a **catalyst** and only speeds up the reaction, and it is not consumed in the reaction

$V(O_2)$ = 67mL

$m(H_2O)$= 133.58g

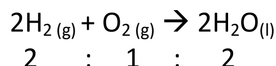
Relate stoichiometry to the law of conservation of mass in chemical reactions by investigating:

- balancing chemical equations
- solving problems regarding mass changes in chemical reactions

Stoichiometry and the Law of Conservation of Mass

Stoichiometry

- The ratio of substances in a chemical reaction- the ratio of reactants to products
- The numbers refer to the relative amounts of moles



$2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(l)}$

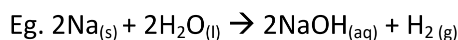
2 : 1 : 2

Law of Conservation of Mass (LFCM)

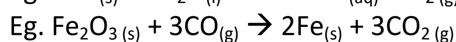
- In a chemical reaction, matter is neither created nor destroyed
  - Therefore, the sum(reactants) = sum(products)
  - Eg. A + B → C
    10g + 10g → 20g
    m(A) + m(B) = m(C)

Balancing equations

- Due to the LFCM, and the consequential fact that the matter cannot be created or destroyed, the same number of each atom must be present on either side of the reaction



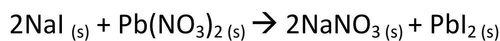
Eg. $2Na_{(s)} + 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2\,(g)}$



Eg. $Fe_2O_{3\,(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2\,(g)}$

Solving problems regarding mass changes in chemical reactions

Sodium iodide and Lead (II) nitrate



$2NaI_{(s)} + Pb(NO_3)_{2\,(s)} \rightarrow 2NaNO_{3\,(s)} + PbI_{2\,(s)}$

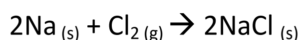
$m(PbI_2) = n(PbI_2) \times MM(PbI_2)$
$= 0.05 \times (207.2 + (2\times126.9))$
$= 23.0g$

$m(NaNO_3) = n(NaNO_3) \times MM(NaNO_3)$
$= 0.10 \times (22.99 + 14.01 + (3\times16.00))$
$= 8.5g$

Sum(reactants) = 15g of NaI + 16.5g of $Pb(NO_3)_2$ = 31.5g
Sum(products) = 23.0g of PbI_2 + 8.5g of $NaNO_3$ = 31.5g

- the law of conservation of mass has been observed as the sum of the mass of the reactants is equal to the sum of the mass of the products

Sodium and chlorine



$2Na_{(s)} + Cl_{2\,(g)} \rightarrow 2NaCl_{(s)}$

$m(Na) + m(Cl_2) = 50g$
$m(NaCl) = 50g$

Mole concept

Inquiry question: How are measurements made in chemistry?

Explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and number of particles in:
- moles of elements and compounds *n* = m/MM (*n* = chemical amount in moles, *m* = mass in grams, *MM* = molar mass in $gmol^{-1}$)
- percentage composition calculations and empirical formulae
- limiting reagent reactions
	<u>Measurements in Chemistry</u>
	
	Measurements
	- Accuracy
	- Error
	- Uncertainty
	- Precision
	- Validity
	- Reliability
	- Relevance
	- Limitations
	
	Analysing (Quantitative) Data- accuracy, precision and uncertainty
	
	- Accuracy: how close the experiment value is to the theoretical value and is often expressed as a % value.
	  - % difference = $\frac{(experimental\ value - theoretical\ value)}{theoretical\ value}$ x 100
	- Precision: how close all the measurements are and the instrument used for the measuring- is sometimes expressed as uncertainty, using a ±
	
	Types of Experimental Error
	
	- Outright mistakes
	  - are not technically experimental as they can be avoided, whereas errors can only be minimised
	- Systematic errors
	  - Errors in the system, bias; an instrument that has not been calibrated correctly will always be incorrect in the same way
	  - Repeating errors will not allow for accuracy
	- Random errors
	  - Errors in any direction- can be ignored
	
	Analysing data- reliability, validity and relevance
	
	- Reliability- how reliable the experiment would be if it was repeated
	  - The repeatability of the data
	- Validity- the credibility of the data
	  - Does the data agree with published trends and observations?
	- Relevance- depends on whether the data collected is appropriate for the experiment

The Mole

The concept

- Avogadro's number (6.022×10^{23}) describes how many atoms there are in one mole of a substance
  - Eg. In one mole of calcium, there are 6.022×10^{23} calcium atoms
    In one mole of hydrogen, there are 6.022×10^{23} hydrogen atoms
- n = $\frac{(no.of\ atoms\ or\ molecules)}{6.022 \times 10^{\wedge}23}$

The moles of elements and compounds

- to find the moles in an element, you use $n = \frac{mass\ (m)}{molar\ mass\ (MM)}$
  where n = chemical amount in moles, m = mass in grams and MM = molar mass in $gmol^{-1}$
  - Eg. 80g of O_2: $n(O_2) = \frac{80}{(2\ x\ 16.00)} = 2.5$
- to find the moles in a compound, you use the mass in grams of the compound divided by the sum of all the elements in the compound's molar weights, including the stoichiometric ratios
  - Eg. 15mL of H_2O: $n(H_2O) = \frac{15}{(2\ x 1.008+16.00)} = 0.8325932504$
- to find the number of molecules in a substance, you use
  $no.molecules = \frac{n}{6.022\ x\ 10^{\wedge}23}$
  - Eg. 80g of O_2: $no.molecules = \frac{n}{6.022\ x\ 10^{\wedge}23} = 1.5 \times 10^{24}$
    no. atoms $= (1.5 \times 10^{24}) \times 2$
    $= 3.0 \times 10^{24}$
- Worked example:

  What is the mass of 0.025 moles of Na_2CO_3?
  First step: calculate the MM of Na_2CO_3
  MM (Na_2CO_3)= (2 x 22.99) + (12.01) + (3 x 16.00)
  = 105.99
  Second step: substitute into the relevant formula
  m(Na_2CO_3) = n x MM = 0.025 x 105.99
  m(Na_2CO_3) = 2.6g

Percentage composition calculations and empirical formulae

- Empirical formula: simplest whole number ratio
- Molecular formula: actual whole number ratio

- Worked example (empirical):

When a piece of copper weighing 2.50g is burned in oxygen, the resulting compound has a mass of 3.13 grams. What is the empirical formula of the compound?

Mass of oxygen= 3.13-2.50
= 0.630g

$n(O) = \frac{m}{MM} = \frac{0.630}{16.00}$
= 0.039375

$n(Cu) = \frac{m}{MM} = \frac{2.50}{63.55}$
= 0.0393391

0.0393391 : 0.039375 = 1 : 1 (given experimental error)
therefore this compound is CuO – Copper (II) oxide

- Worked example (molecular):

Gravimetric analysis of acetic acid shows that it consists of 40.01% carbon, 53.27% oxygen and the remainder is hydrogen. Calculate the empirical formula of acetic acid and, given that its relative molecular mass is 60.05, predict its molecular formula.

	C	O	H
Percentage (%)	40.01	53.27	6.720
Mass in 100g	40.01g	53.27g	6.720g
No. moles in $n=\frac{m}{MM}$	$n(C)=\frac{m}{MM}=\frac{40.01}{12.01}$ = 3.331	$n(O)=\frac{m}{MM}=\frac{53.27}{16.00}$ = 3.330	$n(H)=\frac{m}{MM}=\frac{6.720}{1.008}$ = 6.667
Ratio	1	1	3

Therefore, the empirical formula for acetic acid is CH_2O.

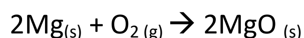
Relative empirical mass = 12.01 + 16.00 (2 x 1.008)
= 30.026

Ratio = 30.026 : 60.05
= 1 : 2

Therefore, the molecular formula is $C_2H_4O_2$.

Limiting reagent reactions

- Worked example 1
  If 1.5g of Mg is burned, is the mass of the magnesium oxide the same?

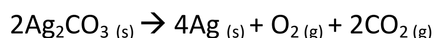
  

  n(Mg) = $\frac{m}{MM} = \frac{1.5}{24.31}$
  = 0.0617… mol
  Mg : MgO = 2 : 2 = 1 : 1

  Therefore, the number moles is the same for both the reactants and the products.

  MM(MgO) = 24.31 + 16.00
  = 40.31
  m = n x MM = 0.0617… x 40.31
  =2.487…
  =2.5g

- Worked example 2
  What mass of silver forms when 25g of silver carbonate decomposes?

  MM(Ag_2CO_3) = (2 x 107.9) + 12.01 + (3 x 16.00)
  = 275.81
  n(Ag_2CO_3) = $\frac{m}{MM} = \frac{25}{275.81}$
  = 0.0906… mol

  Ag_2CO_3 : Ag = 2 : 4 = 1 : 2

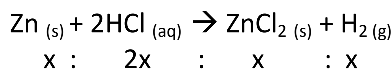
  n(Ag) = 2 x 0.0906…
  = 0.1812…

  m(Ag) = n x MM
  = 0.1812… x 107.9
  = 19.56g
  = 20g

- Steps for limiting reagent reactions
  1. Write a balanced chemical equation
  2. Convert the quantities of the reactants given into moles
  3. Determine the number of moles of each reactant required using stoichiometric ratios
  4. Identify the limiting reagent as the substance present in insufficient amounts
  5. Use the number of moles of the limiting reagent to determine the number of unknown moles
  6. Convert the number of moles to mass, volume or number of molecules as required

- Worked example 3

  If 2.74g of hydrochloric acid in solution is added to 3.27g of zinc, what mass of hydrogen gas is produced?

  $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\,(s)} + H_{2\,(g)}$

  x : 2x : x : x

  n(Zn) = $\frac{m}{MM} = \frac{3.27}{65.38}$
  = 0.0500152952

  n(HCl) = $\frac{m}{MM} = \frac{2.74}{(1.008+35.45)}$
  = 0.07515497285

  Zn : HCl = 1 : 2

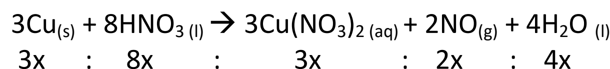
  The moles of the reactants are not in a 1:2 ratio; HCl does not have twice the moles of Zn, therefore making it the limiting reagent. Half of the moles of HCl, 0.03757748642, is all that Zn can react with, causing some Zn to be in excess. Therefore the ratio is:

  Zn : HCl : $ZnCl_2$: H_2 = 1 : 2 : 1 : 1
  0.03757748642 : 0.07515497285 : 0.03757748642 : 0.03757748642
  (with 0.012435 mol excess of Zn)

  m(H_2) = n x MM
  = 0.03757748642 x 2.016
  = 0.0758g

- Worked example 4

  Determine the volume of nitric oxide produces at 25°C and 100kPa when 12.71g of copper reacts with 25.21g nitric acid.

  $3Cu_{(s)} + 8HNO_{3\,(l)} \rightarrow 3Cu(NO_3)_{2\,(aq)} + 2NO_{(g)} + 4H_2O_{(l)}$

  3x : 8x : 3x : 2x : 4x

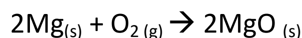
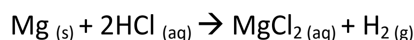
  n(Cu) = $\frac{m}{MM} = \frac{12.71}{63.55}$
  = 0.2 mol

  n(HNO_3) = $\frac{m}{MM} = \frac{25.21}{(1.008)+(14.007)+(3\,x\,16.00)}$
  = 0.4000634769 mol

  Cu : HNO_3 = 3 : 8
  = 0.20 : 0.533

  As there is insufficient HNO_3 it is the limiting reagent. Therefore the ratio according to the limiting reagent is:

  Cu : HNO_3 = 0.15 : 0.40 (with 0.05mol of Cu in excess)

Conduct an investigation to determine that chemicals react in simple whole number ratios by moles	**<u>Experiment- chemicals react in simple whole number ratios by moles</u>** Magnesium to Magnesium Oxide  m(Mg) = 0.1604g m(empty crucible) = 30.0986g m(crucible with MgO) = 30.3791g m(MgO)= 0.2729g m(O) = 0.1125g n(Mg) = $\frac{m}{MM} = \frac{1.604}{24.31}$ = 0.006598… N(O) = $\frac{m}{MM} = \frac{0.1125}{16.00}$ = 0.007031… Ratio = 0.006598… : 0.007031 = 1:1 Therefore the empirical formula is MgO
Conduct a practical investigation to demonstrate and calculate the molar mass (mass of one mole) of: ○ an element ○ a compound	**<u>Investigating the mass of one mole of substances</u>** Determining the molar mass of an element Magnesium + 2M Hydrochloric acid → Hydrogen gas + Magnesium chloride  The number of moles of hydrogen produced equals the moles of magnesium that reacts V($H_{2\,(g)}$) = V(initial burette) – V(final burette) = 50.00mL – 30.01mL = 19.99mL = 0.02L n(H_2) = $\frac{V\,(L)}{MV\,(L)} = \frac{0.02}{24.79}$ (at 25°C and 100 kPa) = 0.000806… Mg : H_2 = 1 : 1 Therefore the moles of H_2 produced equals the magnesium that reacted n(Mg) = 0.000806… MM(Mg) = $\frac{m}{n} = \frac{0.02g}{0.000806}$ = 24.79 % difference to the MM stated on the PT = $\frac{(24.79-24.31)}{(24.31)}$ x 100 = 1.97% error

Determining the molar mass of a compound

- The molar mass of Butane gas will be determined by collecting it from a lighter over water. Since butane is not soluble in water, the volume of water displaced and the decreasing mass of the lighter will be used to calculate its molar mass.
- Experiment method summary: fill a measuring cylinder to the top and invert in a water trough. Release gas from lighter at the base of the measuring cylinder. Gas will displace water at the top of the inverted measuring cylinder. Stop the gas when there is still some water remaining in the cylinder. Replace the glass plate over the open end of the cylinder and remove carefully from the trough. Weigh the lighter again. Measure the amount of water remaining in the measuring cylinder.

m(initial lighter) =
m(final lighter) =

V(butane collected) = 0.047L

n(butane collected) = $\frac{V}{MV}$ = $\frac{0.047\ L}{24.79L}$ (at 25^{O}C and 100kPa)
= 0.0189582577

MM(butane) = $\frac{m(gas\ from\ lighter)}{n(butane\ collected)}$ = $\frac{0.108g}{0.01859582577}$
= 56.96g = 57g

% difference to the MM stated on the PT = $\frac{(58.12-57)}{58.12}$ x 100
= 1.93% error

Concentration and Molarity

Inquiry question: How are chemicals in solutions measured?

Conduct practical investigations to determine the concentrations of solutions and investigate the different ways in which concentrations are measured

<u>Introduction to Concentration and Molarity</u>

Concentration Formulae

- Concentration: amount of solute that is dissolved in a solvent
- Molarity: number of moles per litre of solution (molL^{-1} or M)
  - Formula for molarity: c = $\frac{n}{V}$
  - Percent by weight (% w/w), where c = $\frac{weight\ of\ solute}{weight\ of\ solution}$ x 100
    - Used when both the mass of the solute and the solution are known
    - Expressed as a percentage
  - Percent by volume (%v/v), where c $\frac{volume\ of\ solute}{volume\ of\ solution}$x 100
    - Used when liquids are dissolved in other liquids ie. Ethanol in water
  - Parts per million, where c (ppm) = $\frac{weight\ of\ solute\ (mg)}{weight\ of\ solution\ (kg)}$
  - Parts per million, where c (ppm) = $\frac{volume\ of\ solute\ (mL)}{volume\ of\ solution\ (kL)}$

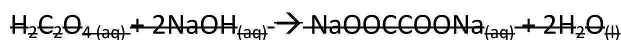
- Parts per million, where c (ppm) = $\frac{mass\ of\ solute\ (mg)}{volume\ of\ solution\ (L)}$

Molality (m)

- When temperate affects the volume of solution
  - Formula: m = $\frac{moles\ (solute)}{kg\ (solvent)}$

~~**Experiment 1 – molarity**~~

- ~~The concentration of the NaOH solution will be determined in moles per litre or molarity by titration.~~
  - ~~Titration: refers to the rank or concentration of a solution with respect to water with a pH of 7 at 25°C. Titration involves volumetric analysis, which is a quantitative technique involving two solutions that react.~~
    - ~~Involved determining the concentration of an unknown solution using an acid-base reaction. In order to determine the concentration of an unknown solution of acid or base, we need to firstly have a solution of known concentration that it can be compared to.~~
    - ~~The solution of a known concentration is known as "the standard" and is:~~
      - ~~A solution of known concentration that is the starting point of a titration, which is prepared by mixing with a solid that is~~
        - ~~Available in a highly pure form~~
        - ~~Has a large molar mass~~
        - ~~Stable in air~~
        - ~~Does not absorb moisture or CO_2 from atmosphere~~
        - ~~Readily soluble in distilled water~~
        - ~~Reacts readily with the solution of unknown concentration~~



~~$H_2C_2O_{4\,(aq)} + 2NaOH_{(aq)} \rightarrow NaOOCCOONa_{(aq)} + 2H_2O_{(l)}$~~

Experiment 2- parts per million

- Calculating salt concentrations is important for determining salinity levels. Commercially it is done using the electrical conductivity or EC of the solution. However, in the laboratory the amount of total dissolved solids will be used to determine the salt concentration.
- Method summary: collect natural salt water, filter using filter paper, and evaporate water from the filtered solution. Weigh the remaining salt to constant mass and calculate its ppm.

m(solution) = 133.47g or 0.13347kg

m(salt) = 5.25g or 5250mg

ppm = $\frac{5250mg}{0.13347\ kg}$

= 39335 ppm

= 39300 ppm

	Experiment 3- percentage composition ○ $NaCl_{(aq)}$ concentration as percent composition, grams per litre and moles per litre, can be determined by performing the one experiment. ○ Prepare an aqueous NaCl solution. Measure out 50mL and transfer to a round bottomed flask. Distill the solution and calculate the mass of the remaining NaCl. Calculate its percentage composition. V(solution) = 50mL m(50mL NaCl solution) = 53.10g m(NaCl) = 9.59g % NaCl = $\frac{9.59\ g}{53.10g}$ x 100 = 18% % H_2O = $\frac{43.51g}{53.10g}$ = 82% Concentration in gL^{-1} V(solution) = 50mL = $\frac{50}{1000}L$ = 0.050L c(gL^{-1}) = $\frac{m(NaCL)}{0.050mL}$ = $\frac{9.59\ g}{0.050L}$ = 191.8gL^{-1} Concentration in $molL^{-1}$ n(NaCl) = $\frac{m}{MM}$ = $\frac{9.59\ g}{(22.99+35.45)}$ = 0.16409… c($molL^{-1}$) = $\frac{0.16409...}{0.0.50L}$ = 3.28 $molL^{-1}$ = 3.28M

Conduct an investigation to make a standard solution and perform a dilution	**Practical Investigation for making a standard solution and performing dilutions**
	
	250mL of a 10% (v/v) solution (standard solution)
	
	○ Equipment
	  ○ Volumetric flasks, Volumetric Bulb pipette (rinsed with distilled water and then methylated spirits), methylated spirits, distilled water
	○ Method
	  ○ calculate what volume of methylated spirits we need to transfer to our volumetric flask
	    ▪ (% v/v) = $\frac{volume\ of\ solute}{volume\ of\ solution}$ x 100
	10% = $\frac{m}{250}$ x 100
	m = 25mL
	Therefore we need 25mLs of methylated spirits to make a 10% (v/v) solution

	100mL of a 2.5% (v/v) solution (dilution) ○ Formula ○ $c_1V_1 = c_2V$ $10.00 \times V_1 = 2.5 \times 100$ $V_1 = \frac{2.5 \times 100}{10.00}$ = 25mL Therefore the volume of the standard solution that needs to be used to make the dilution (V_1) is 25mL 250mL of a 0.0010molL^{-1} solution (standard solution) (V_1= 0.250 L / c_1= 0.0010molL^{-1}) ○ Potassium dichromate $K_2Cr_2O_7$ ○ Formula: $n = \frac{m}{MM} = cV$ $\frac{m}{294.2} = 0.0010 \times 0.25$ mass of solute = (0.01 x 0.25) x 294.2 = 0.7355g ○ Need an electronic balance to measure out something so small accurately 100mL of a 0.001 molL^{-1} solution (dilution) $c_1V_1 = c_2V_2$ $0.001 \times V_1 = 0.001 \times 100$ $V_1 = \frac{0.001 \times 100}{0.01}$ V_1 = 10mL Therefore, you need 10mL of the standard solution to make a 100mL 0.001molL^{-1} solution
Manipulate variables and solve problems to calculate concentration, mass or volume using: ○ ***$c = \frac{n}{V}$* (molarity formula)** ○ **dilutions (number of moles before dilution = number of moles of sample after dilution)**	**<u>Concentration in moles per litre</u>** Formula ○ $c = \frac{n}{V}$ ○ n = number of moles of solute V = volume of solution (L) c = molL^{-1} or M Example Question 1 ○ Calculate the concentration of a solution in moles per litre, when 1 mole of NaCl is dissolved in water to make a solution of final volume 500mL $c = \frac{n}{V}$ $c = \frac{1}{0.500}$ c = 2 molL^{-1}

Example Question 2

- 16.80g of sodium hydrogen carbonate was dissolved in water to produce a solution of final volume 750.0mL. Calculate the concentration of the solution in moles per litre.
  - c = $\frac{n}{V}$
    c = $\frac{n}{0.750L}$
    n($NaHCO_3$) = $\frac{m}{MM}$ = $\frac{16.80}{84.011}$
    = 0.199973813
    c = $\frac{0.199973813}{0.750}$
    c = 0.2666molL^{-1}

Example Question 3- concentration of ions

- Calculate the concentration of magnesium ions and chloride ions in a 0.20 M solution with a volume of 25.0mL.
  - $MgCl_2 \rightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)}$
    c(Mg^{2+}) = 0.20 M
    c(Cl^-) = 0.40 M

Concentration in grams per litre

Formula

- c = $\frac{m(solute)}{v}$ gL^{-1}

Example Question 1

- Calculate the concentration in grams per litre when 0.1241 moles of $CuSO_4.5H_2O$ is made up to a solution volume of 750.0mL.
  - m = n x MM
    m($CuSO_4.5H_2O$) = (159.609 + (5 x 18.01528))
    = 30.98595814
    c = $\frac{30.98595814}{0.750}$
    c = 41.32gL^{-1}

Dilution Calculations

Formula and Foundation Knowledge

- $n_1 = c_1 \times V_1$ and $n_2 = c_2 \times V_2$
  Therefore: n_1 (in concentrate before dilution) = n_2 (after dilution)
  - Think about it like cordial: you pour 10mL of concentrate into a glass, then 300mL of water. In the 310mL solution, there is still only 10mL of concentrate even though it now has a different mass. Hence, when there are 0.2 moles of a substance in a concentrate, the dilution wherein all the concentrate has been used to make the dilution still has 0.2 moles of the formerly concentrated substance.
  - As $n_1=n_2$, we can say that $c_1V_1 = c_2V_2$

Example Question 1

- Calculate the volume of 0.468 M potassium chloride solution that would be required to make 250.0mL of 0.121 M potassium chloride solution.
  - $c_1V_1 = c_2V_2$
    $0.468 \times V_1 = 250.0 \times 0.121$
    $V_1 = \frac{250.0 \times 0.121}{0.468}$ 2
    $= 64.64mL = 0.0646\ L$
    Therefore the volume of 0.468 M KCl required to make a 250mL of a 0.121 M solution is 64.6mL

Gas Laws

Inquiry question: How does the Ideal Gas Law relate to all other gas laws?

Conduct investigations and solve problems to determine the relationship between the Ideal Gas Law and:

- **Gay-Lussac's Law (temperature)**
- **Boyle's Law**
- **Charles' Law**
- **Avogadro's Law**

The Ideal Gas Law Theory

Formula

- $pV = nRT$
- p = pressure
  V = volume
  n = the number of moles
  R = the universal gas constant (8.31 joules per K per mole)
  T = temperature

Early experiments using air

- Performed in the mid 17th century by Torricelli (Italian physicist 1608-1647)
- Poured mercury into a 3ft long glass tube of 1 inch diameter, sealed at one end. When he inverted the tube into a dish filled with mercury, the level of mercury fell about 76cm. The empty space was referred to as the Torricellian vacuum.
- Pascale hypothesized that if the pressure of the air on the mercury dish was responsible for the height of the mercury in the tube, then the mercury would rise less at higher altitudes like on top of a mountain because air pressure would be less.
- In 1648, Pascale's brother-in-law, Perier, replicated Torricelli's experiment on top of Puy de Dome, a French mountain. He found that the mercury did indeed rise less in Torricelli's tube on top of the mountain than it did at the base of the mountain.

Boyle's Law

- $P_1V_1=P_2V_2$
- At constant *T*, as pressure increases, volume decreases
- Robert Boyle discovered the first of the gas laws when he studied the relationship between the volume and pressure of a gas. It may have been Torricelli's work that inspired Boyle's assistant, Robert Hooke, to build his equipment. Robert Hooke designed and built efficient air pumps needed for his experiments and Boyle used a bent tube in the shape of a J to study the elasticity of gases. He sealed the shorter arm of the tube and labelled this section in inches using a piece of paper. He then poured mercury into the tube in the unsealed longer arm and by doing this, trapped air in the shorter arm. Boyle noticed that pressure is inversely proportional to volume. He deduced this by performing many calculation where he found the pressure multiplied by the volume of any measurement was the same as another measurement for a different volume. Boyle's law is now stated **as the**

	pressure exerted by a given mass of gas at constant temperature, is inversely proportional to the volume occupied by the gas. Charles' Law ○ $\frac{V1}{T1} = \frac{V2}{T2}$ ○ At constant *P*, as volume increases, temperature increases ○ It was more than 100 years after Boyle's work that Charles was inspired to experiment with balloons after hearing of Montgolfier's efforts to send a balloon of about 900 centimetres diameter inflated by fire up in the air to a height of just over 2 kilometres. Charles observed that the volume of a gas is directly proportional to its temperature. His law was stated as **at constant pressure, the volume of a given mass of gas is directly proportional to its absolute temperature**. ○ Graphically, Charles' law can be represented to show a linear relationship between the volume of a gas and its temperate at constant pressure. Gay-Lussac's Law ○ $\frac{P1}{T1} = \frac{P2}{T2}$ ○ At constant *V*, as pressure increases, temperature increases ○ In the early 1800s, Gay-Lussac began testing Charles' law using different gases. Gay-Lussac's most significant contribution related to the volumes of gases that reacted in a chemical reaction, and he stated that **the ratio of the volumes of gases consumed or produced in a chemical reaction is equal to the ratio of simple whole numbers**. ○ For example, 2 litres of hydrogen gas will combine with 1 litre of oxygen gas to produce 2 litres of gaseous water. Gay-Lussac also stated that **if the volume and mass of an ideal gas are kept constant, then the pressure and temperature are directly proportional. That is, as temperature increases, the pressure increases.** Avogadro's Law ○ Avogadro based his hypothesis on Gay-Lussac's law and Dalton's atomic theory. Avogadro proposed that Gay-Lussac's law of combining volumes could be explained by assuming that equal volumes of different gases collected under similar conditions contain the same number of particles. Therefore, **the volume of a gas is proportional to the number of particles of the gas and mathematically, can be represented as volume = K, a constant, multiplied by n, the number of particles.** Combined Laws ○ $\frac{P1V1}{T1} = \frac{P2V2}{T2}$ Hence: Ideal Gas Law ○ PV = nRT

Gas Law Experiments

Boyle's Law Experiments (assume temperate is constant)

- Fill up a balloon just enough that it fits inside a syringe and let it sit at the bottom. Place your finger over the hole of the syringe and place the plunger at the highest mark.
- As you push the plunger down, the size of the balloon decreases, as its volume is decreasing as the pressure acting on it gets greater.
- When remove your finger momentarily and then start drawing the plunger back up, the balloon becomes smoother and larger, as the pressure on it decreases. As the pressure on the balloon decreases, the balloon increases in volume.

- We will heat a small amount of water in a can until we can see some steam coming out. The vapour from the boiling water pushes air out of the can.
- When the can is suddenly cooled by inverting it into a container of ice cold water, the water vapour condenses creating a partial vacuum. The extremely low pressure of the partial vacuum inside the can makes it possible for the can to be crushed by the air pressure outside the can.
- That is, the air pressure outside is greater than the pressure inside the can and the pressure difference is greater than the can's structure can support.

These two simple experiment demonstrate Boyle's Law which states, the pressure exerted by a given mass of gas at constant temperature is inversely proportional to the volume occupied by the gas.

Charles' Law Experiment (assume pressure is constant)

- Liquid nitrogen (-196^{o}C) can be used to see the effect of lowering the temperature of a balloon.
- When a balloon is dipped into liquid nitrogen, it begins to shrink. This is because as the temperature of the balloon decreases, as does its volume.
- When someone then blows on the shrunken balloon, it begins to warm up, and therefore, its volume increases once again.
- Charles' law states that at constant pressure, the volume of a given mass of gas is directly proportional to its absolute temperature, and this has been demonstrated by this simple experiment.

Gay-Lussac's Law Experiment

- Place a hard boiled egg at the mouth of a conical flask filled with air; the egg does not fit through the neck of the bottle as the diameter of the egg is larger than that of the conical flask's opening. The air pressure both inside and outside of the flask are the same, with gravity acting on the egg.
- Light a strip of rolled up paper and place it in the conical flask, then replace the egg at the opening. The heat from the flame is heating the gaseous particles inside, as well as producing CO2, H2O and possibly

some CO from the combustion reaction. So all these gases will now fill the container and bounce off the rigid glass walls. Once the flame has gone out, the temperature decreases, hence reducing the volume and pressure exerted by the gases. So now, with the outside pressure being greater, the egg goes through.

Avogadro's Law Experiment

- Filling syringes to the same marker with different gases, then recording the mass to find the number of moles, using n = $\frac{m}{MM}$. Then, the number of particles of gas can be determined, using $N_p = n \times N_A$ (number of particles = moles x Avogadro's number).
- 80mL of He, N and O were measured and then the masses were recorded.
  - 80mL of helium = 45.351g
  - 80mL of oxygen = 45.457g
  - 80mL of nitrogen = 45.444g
- Hence, the moles of each can be calculated, then the amount of particles. The calculations reveal that the number of particles in each volume of gas are about the same, with experimental error taken into account.

In conclusion, the ideal gas law is a combination of Boyle's, Charles', Gay-Lussac's and Avogadro's laws.

The Ideal Gas Law Calculations

Main things to consider when using Gas Law calculations

- Pressure
- Temperate
- Volume
- Amount of gas

SI units

Factor	SI unit	
Pressure	Pascal (Pa)	
Temperature	Kelvin (K)	$^{o}C + 273.15$
Volume	Cubic metre (m^3)	$1m^3 = 1000mL$

Gay-Lussac's Law

- $\frac{P1}{T1}=\frac{P2}{T2}$
- Where *P* is pressure in kPa and *T* is temperature in K
- Worked Example 1
  - The temperature of a gas is 34^{o}C and its pressure is 115kPa. If the temperature was originally 45^{o}C, what was the original pressure?
    - $\frac{P1}{T1} = \frac{P2}{T2}$

      $\frac{P1}{318.15} = \frac{115}{307.15}$

      $P_1 = 318.15 \times \frac{115}{307.15}$

      = 119Pa

Boyle's Law

- $P_1V_1 = P_2V_2$
- Where P is pressure and V is volume
- Worked Example 2
  - A gas occupies a volume of 2.12L at 200kPa. What volume will it occupy at 100kPa?
    - $P_1V_1=P_2V_2$
      $200 \times 2.12 = 100 \times V_2$
      $V_2 = \frac{200 \times 2.12}{100}$
      $= 4.24$ L
      Therefore, the gas will occupy a volume of 4.24L at 100kPa.

Charles' Law

- $\frac{V1}{T1}=\frac{V2}{T2}$
- Where *V* is volume and *T* is temperature
- Worked Example 3
  - A party balloon, that is completely elastic, is filled with helium and occupies a volume of 1.2L at 100kPa in a room with a temperature of 25°C. When the balloon is taken outside the room, where the temperature is 8°C, what will be its volume, assuming that the pressure is the same?
    - $\frac{V1}{T1}=\frac{V2}{T2}$
      $\frac{1.2}{298.15} = \frac{V2}{281.15}$
      $V_2 = 281.15 \times \frac{1.2}{298.15}$
      $= 1.1$L (2 sig fig)

Avogadro's Law

- V = kn and expressed usually as $k = \frac{V}{n}$
- Where V is volume, k is a constant for a given temperature and pressure and n is the number of moles
- Worked Example 4
  - A balloon at room temperature and pressure, is initially filled with 0.10 moles of helium gas and has a volume of 2L. Calculate the final volume of the balloon if more helium is added so that the final number moles of helium is 0.15.
    - $k = \frac{V}{n}$
      $k = \frac{2}{0.10}$
      $= 20$
      V = kn
      $V = 20 \times 0.15$
      $= 3$ L

	Ideal Gas Law ○ PV = nRT ○ Where P is pressure (kPa), V is volume (L) , n is the number of moles, R is the universal gas constant (8.13J per K per mole) and T is temperature (K) ○ Worked Example 5 ▪ Calculate the volume a 24.02g sample of methane gas will occupy at 22°C and 125kPa. • $n(CH_4) = \frac{24.02}{12.01+(4\ x\ 1.008)}$ $= 1.497319536$ PV = nRT $V = \frac{nRT}{P}$ $V = \frac{1.497319536 \text{ x } 8.13 \text{ x } 295.15}{125}$ = 28.74 L