Stoichiometric Relationships

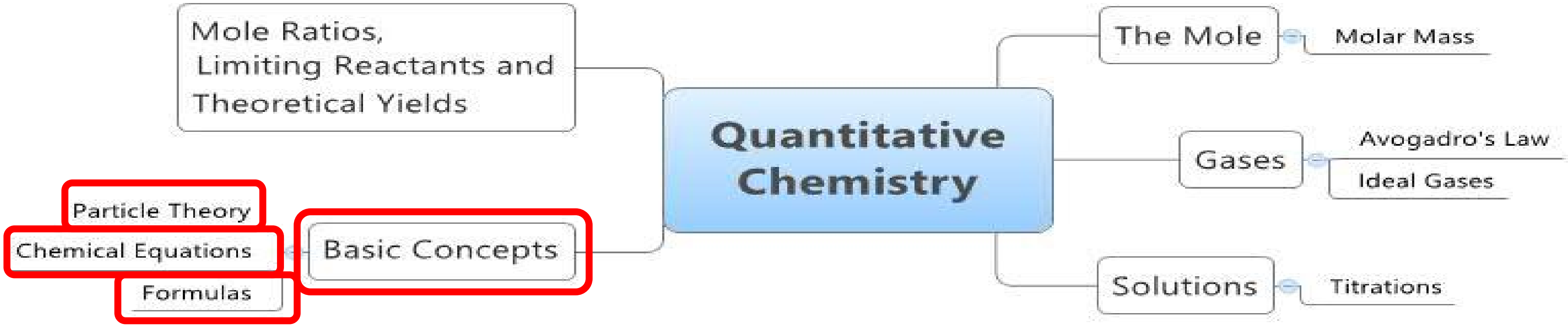
Ms. Peace

Lesson 1

Introduction to Stoichiometry

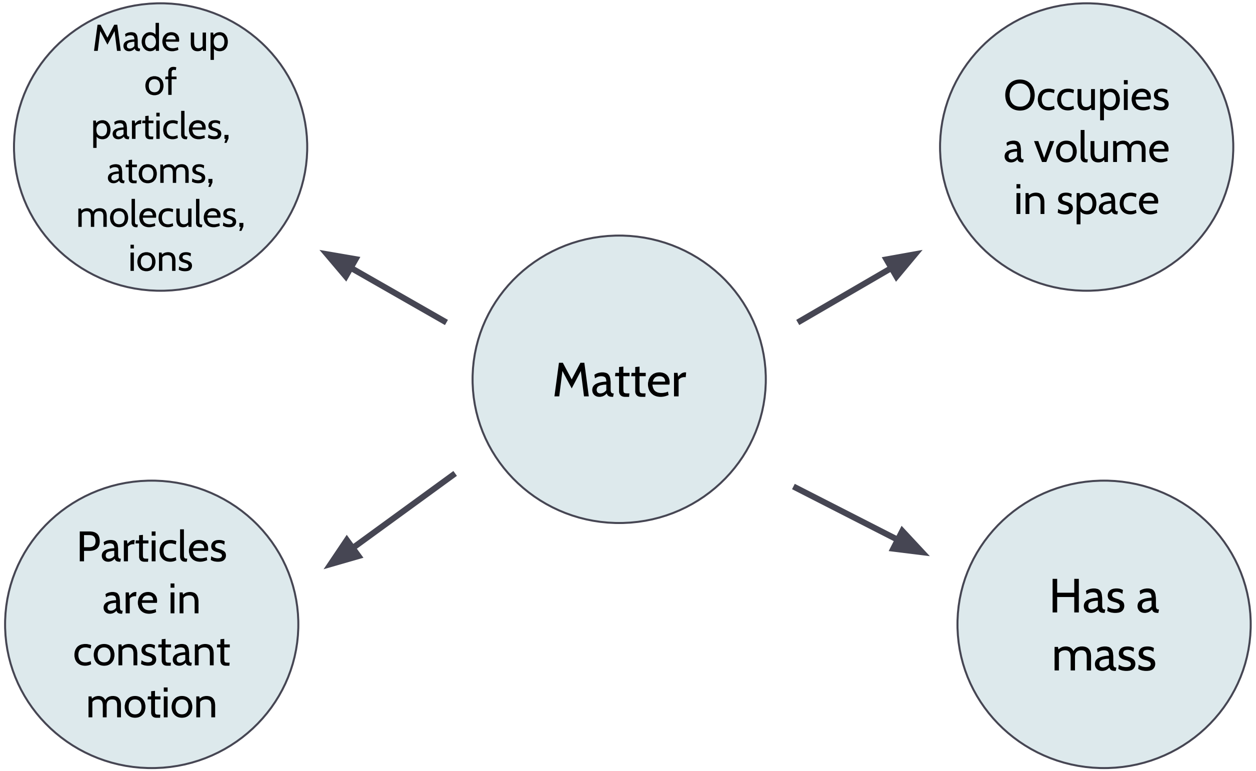
Main

We Are Here

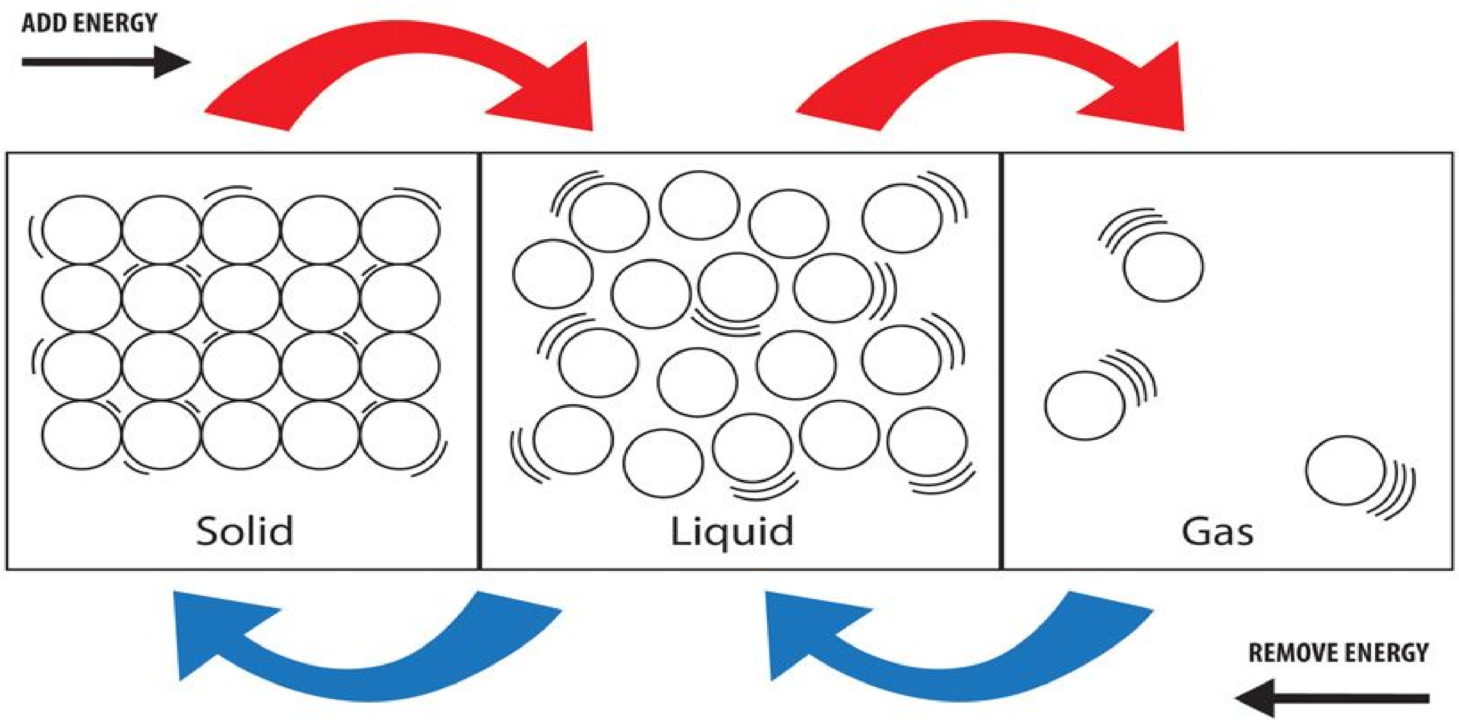


Main

States of Matter



States of Matter



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Properties of States of Matter

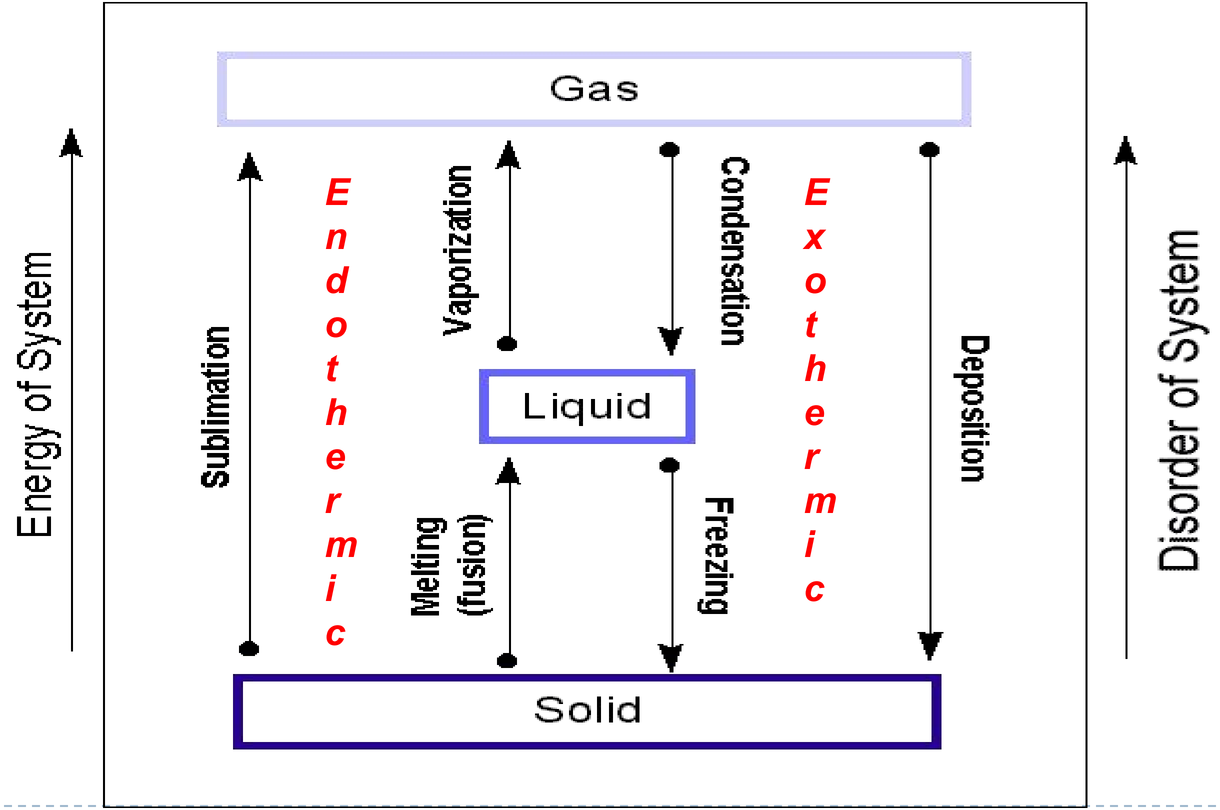
	SOLID	LIQUID	GAS
Distance	Close together	Close but further apart than in solids	Particles far apart
Arrangement	Regular	Random	Random
Shape	Fixed shape	No fixed shape-take the shape of the container	No fixed shape-fill the container
Volume	Fixed	Fixed	Not Fixed
Movement	Vibrate	Move around each other	Move around in all directions
Speed	Slowest	Faster	Fastest
Energy	Lowest	Higher	Highest
Forces of attraction	Strongest	Weaker	Weakest

Main

Temperature

- Units:
  - Fahrenheit
  - Celsius
  - Kelvin*
- SI units are a set of standard units that are used in science throughout the world
- Absolute zero: 0 K or -273° C
  - K = ° C + 273

Changes of States



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Changes of States

- Endothermic:
  - Melting and boiling
  - Energy must be transferred from the surroundings to bring about these changes
  - The potential energy of molecules increases; they vibrate more and move faster
- Exothermic:
  - Condensation and freezing
  - Energy is transferred to the surroundings
  - The potential energy of molecules decreases; the vibrate less and move slower

Main

Utilization

- Freeze-drying is a food preservation technique
- Freeze-drying uses the process of sublimation
- Foods that require dehydration are first frozen and then subjected to a reduced pressure
- The frozen water then sublimes directly to water vapor thus dehydrating the food

- How Refrigeration Works

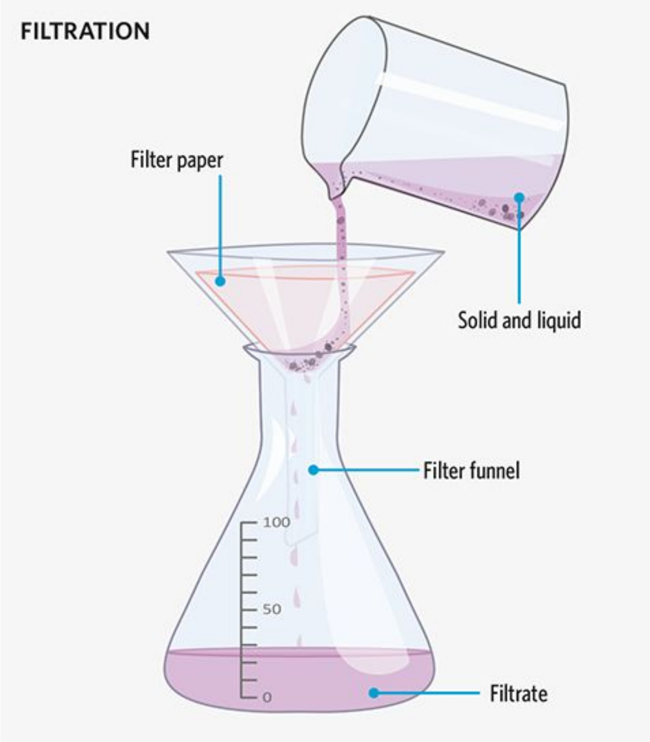
Elements and Compounds

- Elements contain atoms of only one type
- Atoms of elements combine in a fixed ratio to form compounds composed of molecules or ions
- These are the fundamental basis in formulas and chemical equations
- Properties of compounds are very different from those of its constitutional elements

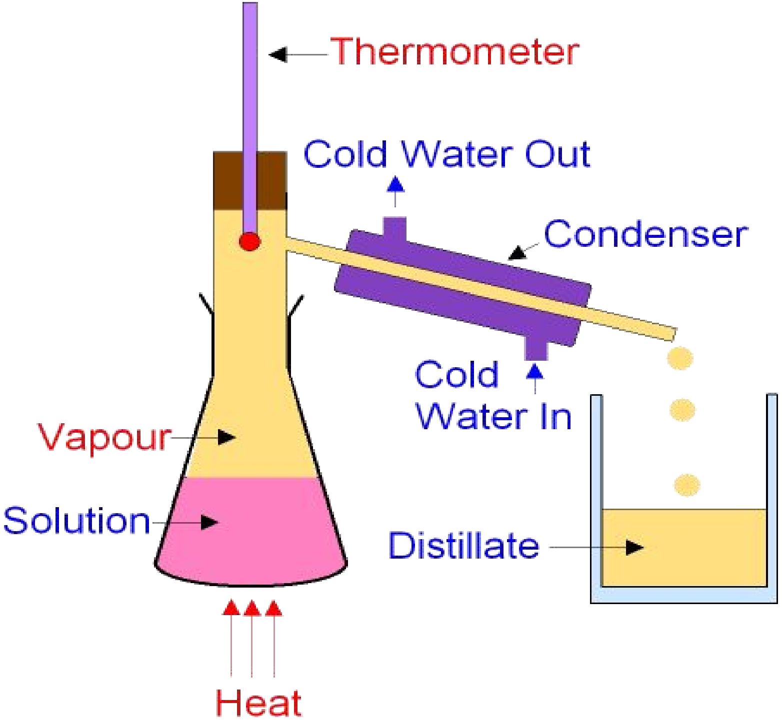
Mixtures

- A pure substance is matter that has a constant composition
  - N_2
  - H_2O
  - NaCl
  - $C_6H_{12}O_6$
- Pure substances combine physically to form a mixture
- Mixtures contain more than one element and/or compound that are not chemically bonded together and so retain their individual properties.

Mixtures

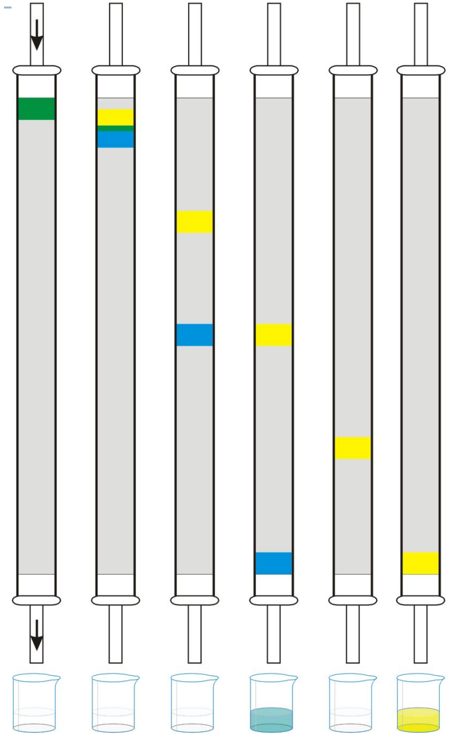


Filtration

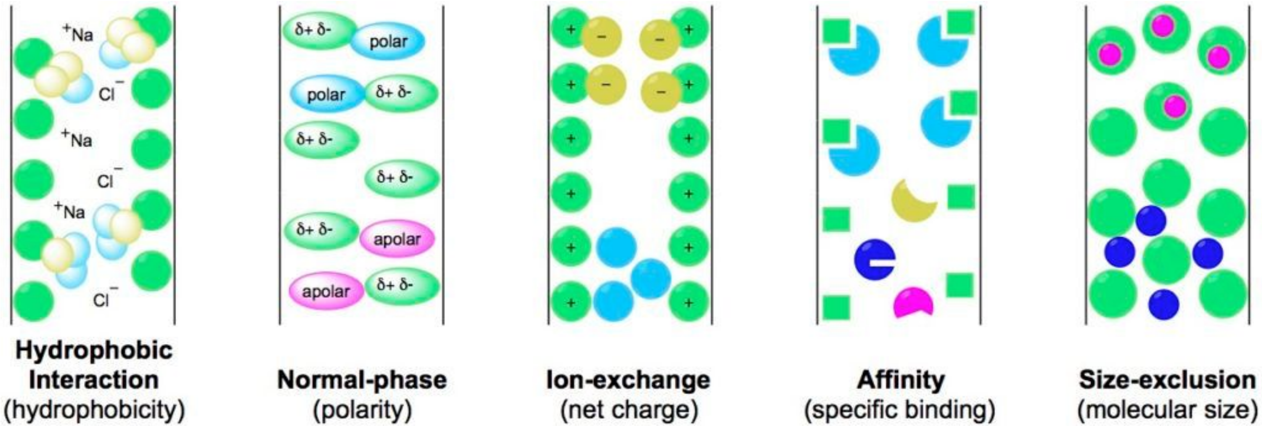


Fractional Distillation

Mixtures



Column Chromatography

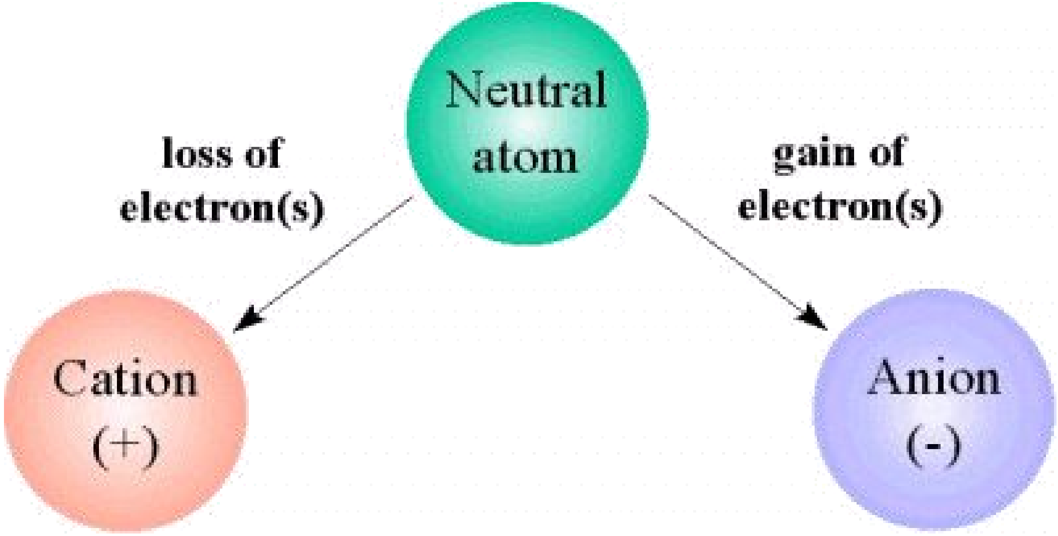


Types of Mixtures

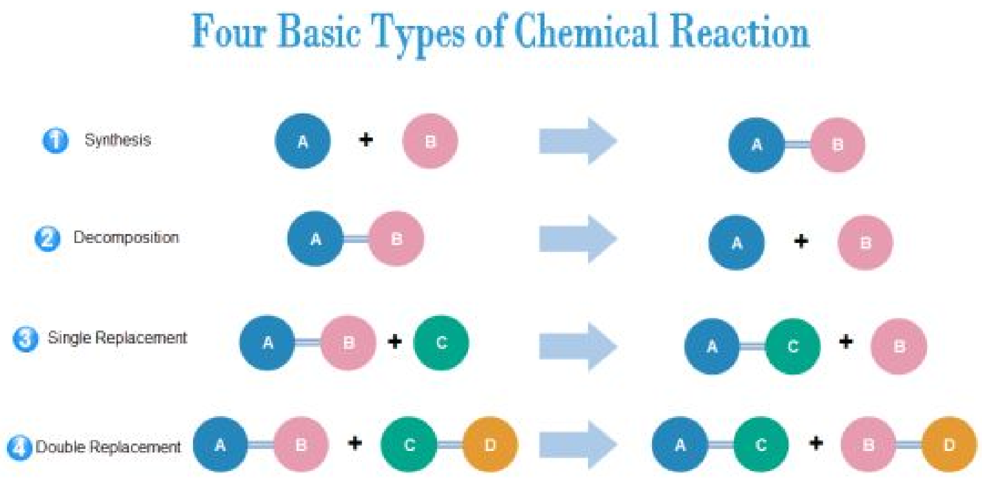
- Homogeneous
  - One phase…you can't see the separation between the parts of the mixture
  - For example: seawater is mixture of salt and water, but you can't see the salt.
- Heterogeneous
  - Multiple phases….you can see the different components of the mixture
  - For example: salad dressing is a mixture of oil and vinegar, and you can see the bits of oil and the bits of vinegar.

MYP Review

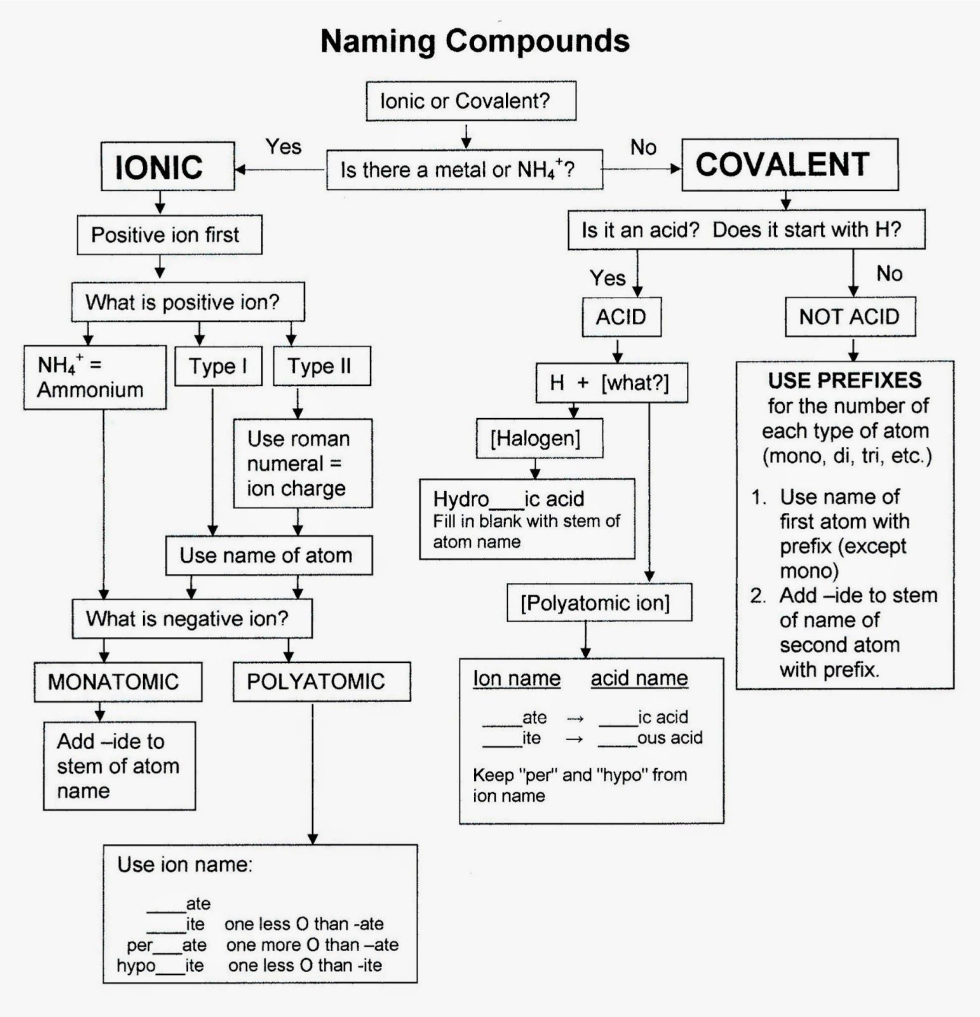
- State Symbols:
  - (s)
  - (l)
  - (g)
  - (aq)



Types of Chemical Reactions



Types	Description	Exmaple
Synthesis Reactions	Elements are joined together.	$2H_2 + O_2 \rightarrow 2H_2O$
Decomposition Reactions	A compound breaks into parts.	$2H_2O \rightarrow 2H_2 + O_2$
Single Displacement Reactions	A single element replaces an element in a compound.	$Zn + 2HCl \rightarrow H_2 + ZnCl_2$
Double Displacement Reactions	An element from each of two compounds switch places.	$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$


Naming Compounds
Ionic or Covalent?
Is there a metal or NH_4^+?
Yes
IONIC
No
COVALENT
Positive ion first
What is positive ion?
NH_4^+ = Ammonium
Type I
Type II
Use roman numeral = ion charge
Use name of atom
What is negative ion?
MONATOMIC
POLYATOMIC
Add –ide to stem of atom name
Use ion name:
____ate
____ite one less O than -ate
per___ate one more O than –ate
hypo___ite one less O than -ite
Is it an acid? Does it start with H?
Yes
No
ACID
NOT ACID
H + [what?]
[Halogen]
Hydro___ic acid
Fill in blank with stem of atom name
[Polyatomic ion]
Ion name acid name
____ate → ____ic acid
____ite → ____ous acid
Keep "per" and "hypo" from ion name
USE PREFIXES for the number of each type of atom (mono, di, tri, etc.)
1. Use name of first atom with prefix (except mono)
2. Add –ide to stem of name of second atom with prefix.

Word Equations

- Hydrogen + Oxygen → Water

REACTANTS PRODUCTS

- $H_2 + O_2 \rightarrow H_2O$

Is this balanced???

Symbol Equations

$$2\,H_2 + O_2 \rightarrow 2\,H_2O \quad \text{vs} \quad H_2 + O_2 \rightarrow H_2O$$

4 H	4 H	2 H	2 H
2 O	2 O	2 O	1 O

- The red numbers are called **coefficients** and tell you the number of each molecule involved in the reaction
  - Required to balance the equation
  - Without them the equation does not balance – each side of the reaction would have different numbers of each atom – which would break physics
  - You **can't change the subscript numbers** in the formulas as this changes the chemical
  - If there is no coefficient, it is '1'

Tips for Balancing Equations

1) Balance metals
2) Balance nonmetals
3) Balance oxygen
4) Balance hydrogen

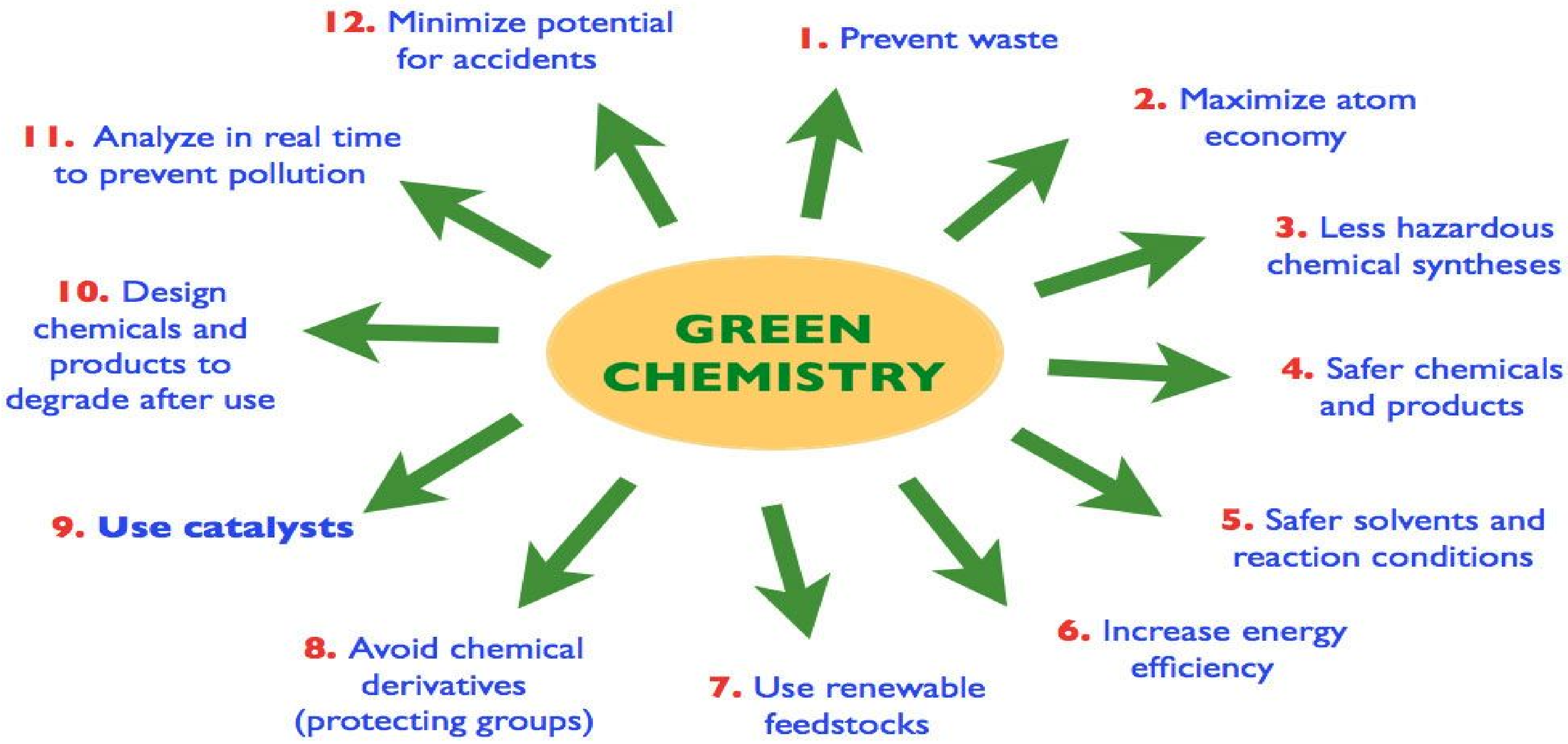
- Try keeping a tally-chart of the numbers of atoms on each side of the equation
- BE PATIENT!!!!

Construct equations and then balance them for each of the following:

- Magnesium (Mg) reacts with hydrochloric acid (HCl) to make magnesium chloride ($MgCl_2$) and hydrogen gas
- Ethane (C_2H_6) reacts with oxygen gas to make carbon dioxide and water
- Lead nitrate ($Pb(NO_3)_2$) reacts with aluminium chloride ($AlCl_3$) to make aluminium nitrate ($Al(NO_3)_3$) and lead chloride ($PbCl_2$)
- Barium nitride (Ba_3N_2) reacts with water to make barium hydroxide ($Ba(OH)_2$) and ammonia (NH_3)
- Ammonium perchlorate (NH_4ClO_4) reacting with aluminium to make aluminium oxide (Al_2O_3), aluminium chloride ($AlCl_3$), nitric oxide (NO) and water
- **Extension:** Write a flow chart that can be followed to let you to balance equations

Main

Green chemistry is the design of **chemical** products and processes that reduce or eliminate the generation of hazardous substances.



Anastas and Warner (1998)

Atom Economy

- Utilization of synthetic reactions and industrial processes that must be increasingly efficient to preserve raw materials and produce fewer and less toxic emissions
- Developed by Professor Barry Trost of Stanford University
- Atom economy looks at the level of efficiency of chemical reactions by comparing the molecular mass of atoms in the reactants with the molecular mass of useful compounds

Atom Economy

- Atom economy is the percentage of reactants changed to useful products

Calculation of Atom Economy

$$\text{atom economy} = \frac{\text{mass of atoms in desired product}}{\text{mass of atoms in reactants}} \times 100\%$$

- In an ideal chemical process the amount of reactants = amounts of products produced
- What does it mean if there is an atom economy of 100%?

Main

Key Points

- The properties of solids, liquids and gases are due to the arrangement and motion of their particles

- Mixtures can be homogeneous or heterogeneous

- Equations must be balanced to ensure that mass is conserved

- A high atom economy means fewer atoms are being wasted

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Lesson 2

Formulas and Composition by Mass

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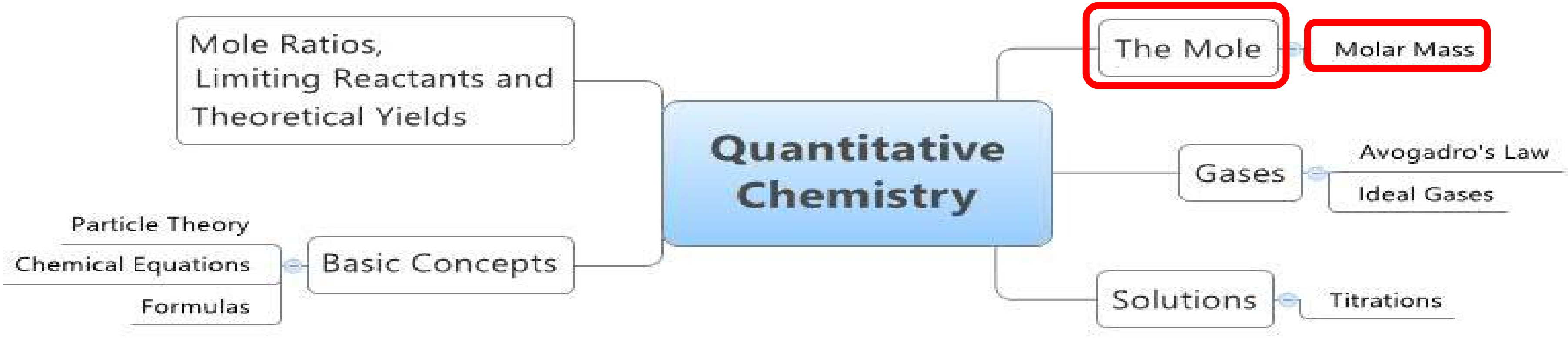

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SI Units

- System of International Units was developed to transcend all languages and cultures

- The International Bureau of Weights and Measures (BIPM) monitors the correct use of SI units in all applications of science

The SI Units

Base quantity	Name	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical Current	ampere	A
Thermodynamic temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Stoichiometry

- Stoichiometry uses the quantitative relationships between amounts of reactants and products in a chemical reaction

- Avogadro's constant, $6.02x10^{23}mol^{-1}$, enables us to make comparisons between chemical species

Multiplication Factor	Prefix	Symbol
1,000,000,000 = 10^9	giga	G
1,000,000 = 10^6	mega	M
1,000 = 10^3	kilo	k
100 = 10^2	hecto	h
1 = 1		
0.01 = 10^{-2}	centi	c
0.001 = 10^{-3}	milli	m
0.000001 = 10^{-6}	micro	µ
0.000000001 = 10^{-9}	nano	n

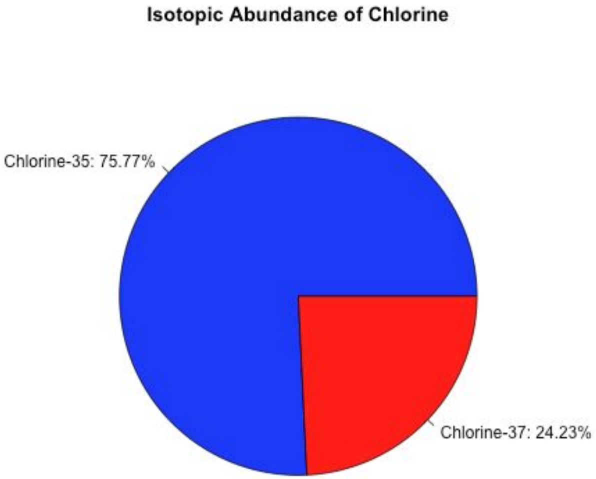
Relative Atomic Mass, A_r

- The periodic table tells you the relative atomic mass of each element
  - This is the mass of an element relative to a 12^{th} of the mass of ^{12}C.
  - It is a relative value, which means it has no units.
- Relative atomic mass has the symbol 'A_r'

- For example carbon: $A_r(C) = 12.01$
  - The reason it isn't a whole number is due to isotopes

Relative Atomic Mass, A_r

- Relative abundance of each isotope is a measure of the percentage that occurs in a sample of the element



$$\text{Relative Atomic Mass (R.A.M)} = \frac{\text{Abundance X Atomic weight}}{100} + \frac{\text{Abundance X Atomic weight}}{100}$$

Main

Relative Molecular or Formula Mass, M_r

- **Relative Molecular Mass** is calculated by adding up the A_r for each atom in a molecule.
- The related term **relative formula mass** refers to the relative mass of one unit of a formula and is used for empirical formulas
- No units as this is a ratio

Calculating M_r

- HCl
  - $A_r(H) = 1.01$
  - $A_r(Cl) = 35.45$
  - $M_r = 1.01 + 35.45 = 36.46$
- C_2H_4
  - $A_r(C) = 12.01$
  - $A_r(H) = 1.01$
  - $M_r = 2x12.01 + 4x1.01$
    $= 16.06$
- H_2SO_4
  - $A_r(H) = 1.01$
  - $A_r(S) = 32.06$
  - $A_r(O) = 16.00$
  - $M_r = 2x1.01 + 32.06 + 4x16.00$
    $= 98.08$
- $Mg(OH)_2$
  - Ar(Mg) = 24.31
  - Ar(O) = 16.00
  - Ar (H) = 1.01
  - $M_r = 24.31 + 2x16.00 + 2x1.01$
    $= 58.33$

Calculate M_r for:

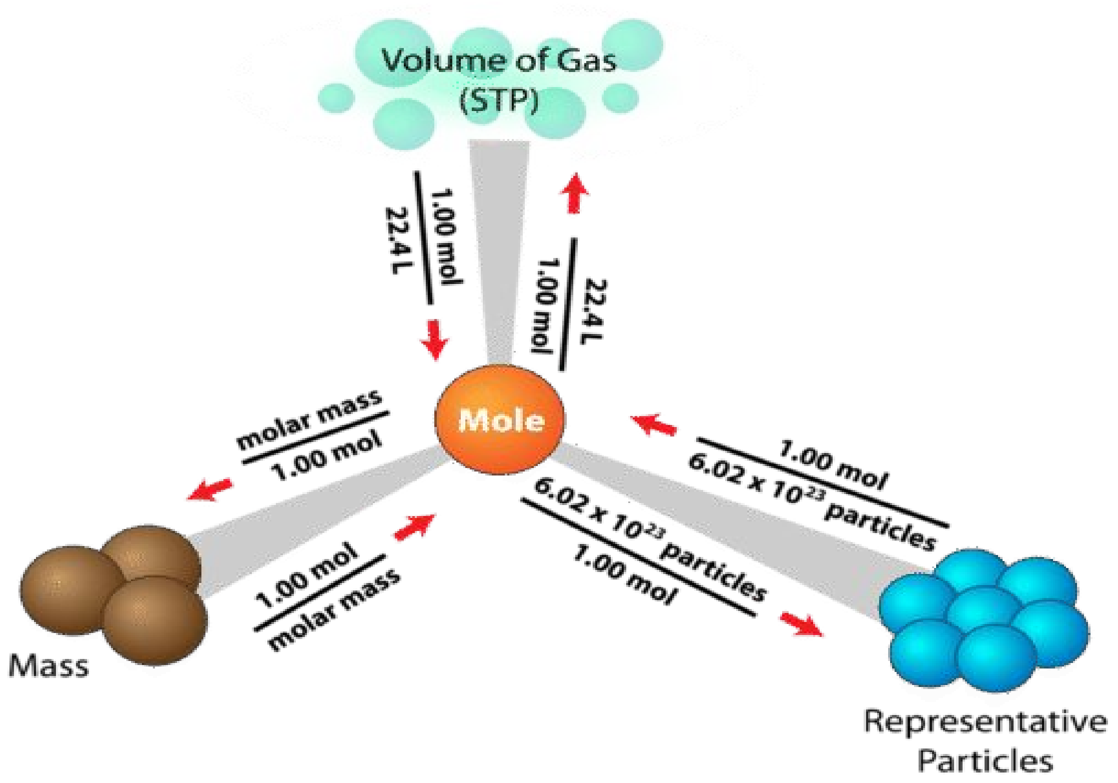
- Br_2
- $(NH_4)_2SO_4$
- C_3H_8
- $C_6H_{12}O_6$

Molar Mass

- **Molar mass** is the mass of one mole of a substance. It has the units of grams per mole, g mol^{-1}

Element	Molar Mass
H_2O	2(1.0) + 16.0 = 18.0 g/mol
$(NH_4)_2CrO_4$	2(14.0)+8(1.0)+52.0+4(16.0) = 152 g/mol
$Ba(NO_3)_2$	137.3 + 2(14.0) + 6 (16.0) = 261.3 g/mol

Mole Calculations



Types of Formulas

- Qualitative analysis: focuses on determining which elements are present
- Quantitative analysis: focuses on determining the relative masses of elements allowing to determine the exact composition
- Empirical Formula: simplest whole-number ratio of atoms or amount of each element present in a compound
- Molecular Formula: the actual number of atoms or amount of elements in one structural unit or one mole of the compound
  - Empirical formulas and molecular formulas can be the same
  - Because of their structure, ionic (and giant covalent) compounds do not form molecules so empirical formula is the only one relevant

Empirical formula: Table summary

Name of compound	Empirical formula	Molecular formula
Hydrogen peroxide	HO	H_2O_2
Water	H_2O	H_2O
Glucose	CH_2O	$C_6H_{12}O_6$
Oxalic acid	HCO_2	$H_2C_2O_4$
Ethanol	C_2H_6O	C_2H_6O
Ethane	CH_3	C_2H_6
Ethylene	CH_2	C_2H_4
Caffeine	$C_4H_5N_2O$	$C_8H_{10}N_4O_2$

Percentage Composition by Mass

- If we divide the total mass of each element in a compound by the number of atoms
- For example ethanol, C_2H_5OH, M_r = 46.08

	C	H	O
Number Present	2	6	1
Multiply by A_r	2 x 12.01 = 24.02	6 x 1.01 = 6.06	1 x 16.00 = 16.00
Divide by M_r, convert to %	24.02/46.08 x 100 = 52.1%	6.06/46.08 x 100 = 13.1%	16.00/46.08 x 100 = 34.7%

Calculate % composition by mass for:

1. H_2O
2. $Mg(OH)_2$
3. $CuSO_4.5H_2O$
4. $C_{12}H_{22}O_{11}$

Example: Calculate the percentage by mass of sulfur in H_2SO_4

H_2SO_4

$2(H) = 2(1.01) = 2.02$

$1(S) = 1\,(32.06) = 32.06$

$4(O) = 4(16.00) = 64.00$

98.08

$$\%S = \frac{32.06}{98.08} \times 100\% = 32.69\%$$

FInd the mass of each element

↓

Find the mass of the compound

↓

Divide the element by the relative atomic mass

↓

Multiply by 100

Example: A sample of a compound contains 20% hydrogen and 80% carbon by mass and M_r = 30.08

C	:	H
80%	:	20%
80/12.01 = 6.67	:	20/1.01 = 20
6.67/6.67 = 1	:	20/6.67 = 3

n/a since no awkward decimals above

Empirical formula = $\mathbf{CH_3}$: 30.08/(12.01 + 3 x 1.10) = 2

Molecular formula = CH_3 x 2 = $\mathbf{C_2H_6}$

Write symbols as a ratio

Write % composition below

Divide each % by A

Divide each by smallest

Multiply to remove fractions

Divide M_r by formula mass

Multiply empirical formula by above

Main

Example: a sample of a compound contains 8.4% hydrogen, 65.2% carbon and 29.1% nitrogen by mass, and M_r = 288.5

C : H : N

62.5% : 8.4% : 29.1%

62.5/12.01= 5.20   8.4/1.01= 8.31   29.1/14.01=2.08

5.20/2.08 = 2.5   8.31/2.08 = 4   2.08/2.08 = 1

2.5 x 2 = 5   4 x 2 = 8   1 x 2 = 2

Empirical formula = $\mathbf{C_5H_8N_2}$ so 288.5/ 96.2 = 3

Molecular formula = $C_5H_8N_2$ x 3 = $\mathbf{C_{15}H_{24}N_6}$

1. Write symbols as a ratio
2. Write % composition below
3. Divide each % by A
4. Divide each by smallest
5. Multiply to remove fractions
6. Divide M_r by formula mass
7. Multiply empirical formula by above

Main

Lesson 3

Mole Ratios and Theoretical Yields

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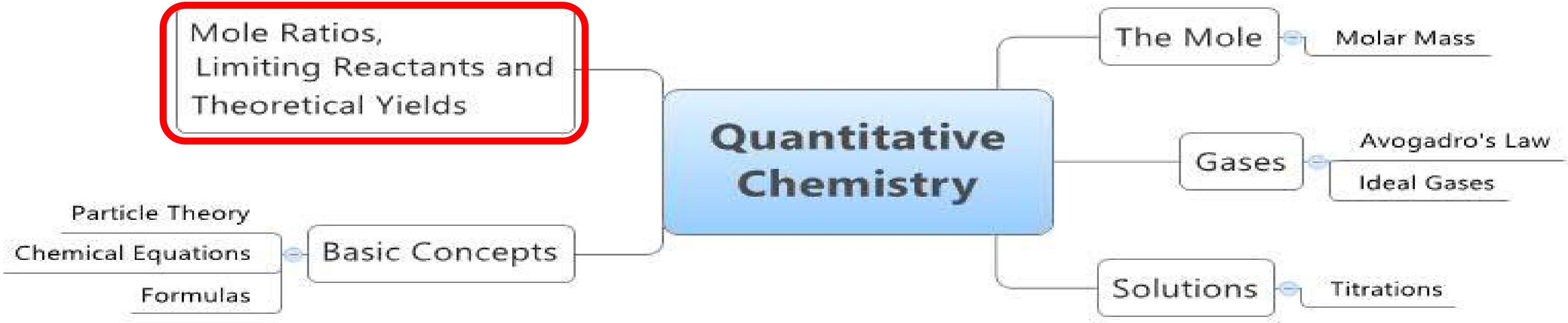
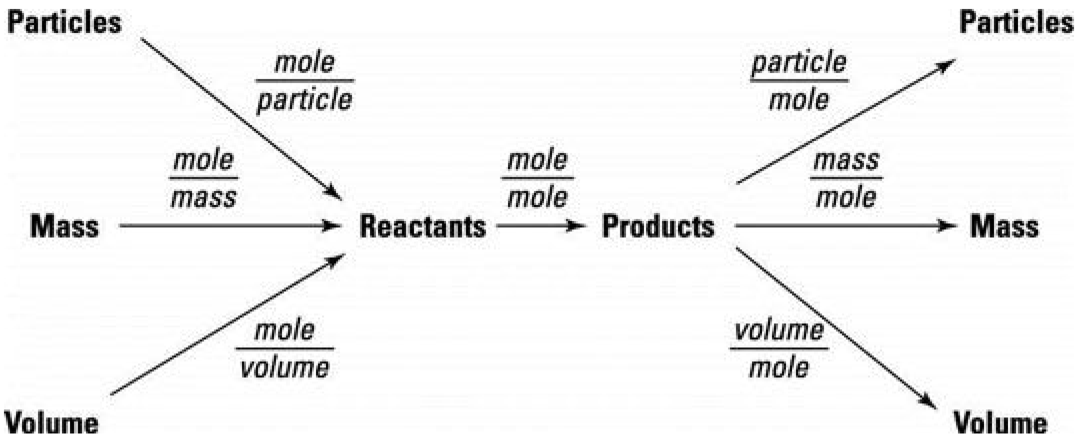

Mole Ratios,
Limiting Reactants and
Theoretical Yields
Quantitative Chemistry
The Mole
Molar Mass
Gases
Avogadro's Law
Ideal Gases
Solutions
Titrations
Basic Concepts
Particle Theory
Chemical Equations
Formulas

Stoichiometry

- Stoichiometry is the quantitative method of examining relative amounts of reactants and products
- Percentage yield is vital in monitoring the efficiency and profitability of industrial processes



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Mole Ratios

- This is the ratio of one compound to another in a balanced equation.
- For example, in the equation

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

- Hydrogen, oxygen and water are present in 2:1:2 ratio.
  - 0.2 mol of H_2 reacts with 0.1 mol of O_2 to make 0.2 mol H_2O
  - 5 mol of H_2 reacts with 2.5 mol of O_2 to make 5 mol of H_2O
  - To make 4 mol of H_2O you need 4 mol of H_2 and 2 mol of O_2

Mole Ratios in Calculations

The mole ratio!

$$n(wanted) = n(given) \times \frac{wanteds}{givens}$$

- wanted = the substance you want to find out more about
- given = the substance you are given the full info for
- n(wanted) = the number of moles you are trying to find out
- n(given) = the number of moles of you are given in the question
- wanteds = the number of wants in the balanced equation
- givens = the number of givens in the balanced equation

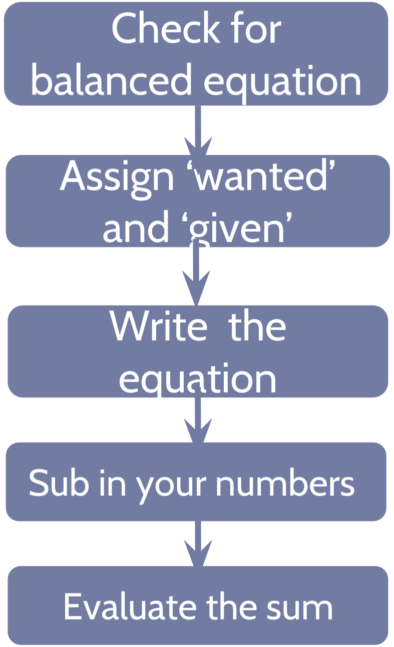
Example 1

- What quantity of $Al(OH)_3$ in moles is required to produce 5.00 mol of H_2O?

$$2\,Al(OH)_3 + 3\,H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3\,H_2O$$

$$n(wanted) = n(given) \times \frac{wanteds}{givens}$$

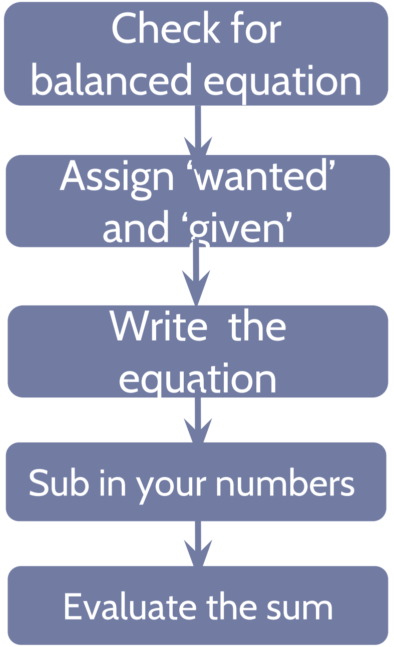
- H_2O is given, $Al(OH)_3$ is wanted.
- $n(Al(OH)_3)$ = 5.00 x (2/3)=**3.33 mol**



Example 2

- What quantity of O_2 in moles is required to fully react with 0.215 mol of butane (C_4H_{10}) to produce water and carbon dioxide?

$$2\,C_4H_{10} + 13\,O_2 \rightarrow 8\,CO_2 + 10\,H_2O$$



$$n(wanted) = n(given) \times \frac{wanteds}{givens}$$

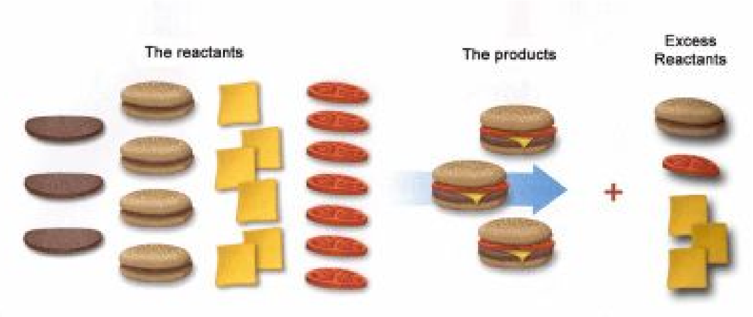
- C_4H_{10} is given, O_2 is wanted.
- $n(O_2) = 0.215 \times (13/2)$
- $n(O_2) =$ **1.40 mol**

Main

Limiting Reagant

- In a reaction, we can describe reactants as being ‘limiting’ or in ‘excess’
  - **Limiting –** this is the reactant that runs out
  - **Excess –** the reaction will not run out of this reactant

- The limiting reactant will be your ‘given’ in further calculations:
  - Determining amounts of products formed
  - Determining amounts of other reactants used



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Limiting Reagent

$$2\,H_2 + O_2 \rightarrow 2\,H_2O$$

- For example, if you have 2.0 mol H_2 and 2.0 mol O_2
  - H_2 is the limiting reactant – it will run out
  - O_2 is present in excess – there is more than enough

- To determine this, **divide the quantity of each reactant by its coefficient** in the equation. The smallest number is the limiting reactant:
  - H_2: 2.0 / 2 = 1.0 – smallest therefore limiting
  - O_2: 2.0 / 1 = 2.0

Example 1: What quantity, in moles, of $MgCl_2$ can be produced by reacting 10.5 g magnesium with 100 cm^3 of 2.50 mol dm^{-3} hydrochloric acid solution?

- $Mg + 2HCl \rightarrow MgCl_2 + H_2$
- Determine limiting reagent:
  - Mg: (10.5g / 24.31g) = 0.432 mol
  - HCl: (0.100dm^3 x 2.50mol dm^{-3})=0.25mol
  - 0.250/2(coefficient) = 0.125mol (smallest therefore is L.R.)

	Mg +	2HCl →	$MgCl_2$ +	H_2
Initial	0.432 mol	0.250 mol	0 mol	0 mol
Change	-0.125 mol	-(2)0.125 mol	+0.125 mol	+0.125mol
Final	0.307 mol	0	0.125mol	0.125mol

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Example 2 : What quantity, in moles, of carbon dioxide would be formed from the reaction of 12.0 mol oxygen with 2.00 mol propane, and how much of which reactant would remain?

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

- Determine limiting reagent:
  - C_3H_8 : 2.00 mol
  - O_2: 12.0mol/5(coefficient) =2.4 mol

	C_3H_8	+ $5O_2$	$\rightarrow$ $3CO_2$	+ $4H_2O$
Initial	2.00 mol	12.0 mol	0 mol	0 mol
Change	-2.00 mol	-5 (2.00mol)	+3 (2.00 mol)	+4 (2.00 mol)
Final	0 mol	2.00 mol	6.00 mol	8.00 mol

Main

Theoretical, actual and percentage yield

- **Theoretical yield** is the maximum amount of product you would make if the limiting reactant was fully converted to product.
  - Use the limiting reactants maths to work this out

- **Actual yield** is the actual amount of product collected in after a reaction
  - Can be different from the theoretical yield

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

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Percent Yield

Percent Yield

$$\text{Percent Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right) 100\%$$

Theoretical yield is the mass of product that you calculate using stoichiometry; it's what you are supposed to be able to get from the reaction (theoretically). Actual yield is the mass of product that you actually obtain from the reaction in the lab.

Airbag Stoichiometry

Lesson 4

Molar Volumes of Gases

Main

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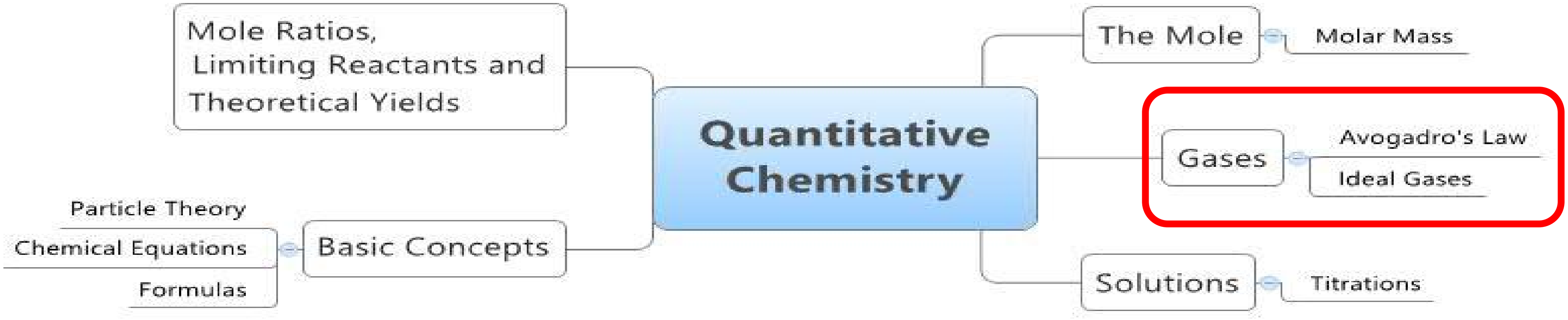


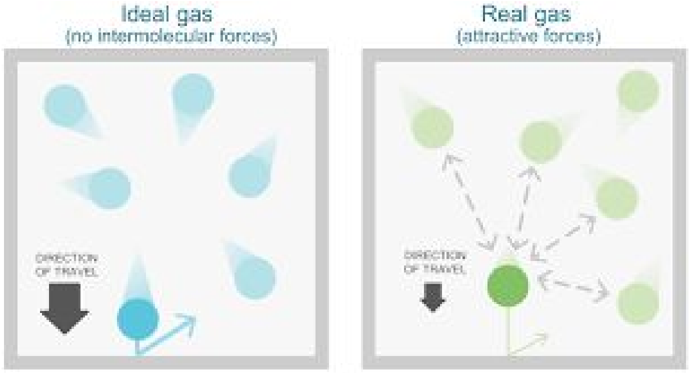
Main

Molar Volume of a Gas

- The kinetic theory of gases is a model used to explain and predict the behavior of gases at a microscopic level
- Postulates of assumptions that must be true for this theory to hold:
  1. Gases are made up of very small particles, separated by large distances. Most of the volume occupied by gas is empty space
  2. Gaseous particles are constantly moving in a straight lines, but random directions
  3. Gaseous particles undergo elastic collisions with each other and the walls of the container. No loss of kinetic energy occurs
  4. Gaseous particles exert no force of attraction on other gases

Molar Volume of a Gas

- Under conditions of STP, an ideal gas obeys these postulates
- At high temperature and low pressure, gases respond in ways that depart from the ideal gas behavior and exhibit behaviors of real gases



The Molar Volume of an Ideal Gas

- At standard temperature and pressure:
  - (STP, T = 273K, P = 1.01×10^5 Pa or 100kPa)
  - Molar Volume of Ideal Gas = 22.7 dm^3 mol^{-1}

Main

Avogadro's Law

- Calculate the moles of oxygen in a $6.73 dm^3$ sample of oxygen gas at STP

$6.73 dm^3$	1 mole
	$22.7 dm^3$

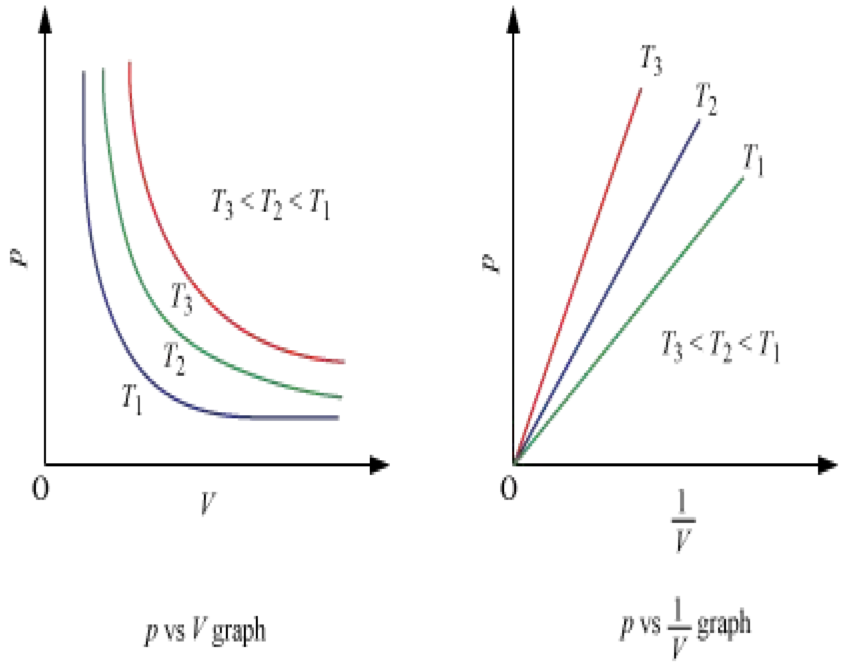
= 0.296 mol

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Boyle's Law

$P_1V_1=P_2V_2$ At a constant temperature, the volume of a fixed mass of an ideal gas is inversely proportional to its pressure

$$V \,\alpha\, \frac{1}{P}$$



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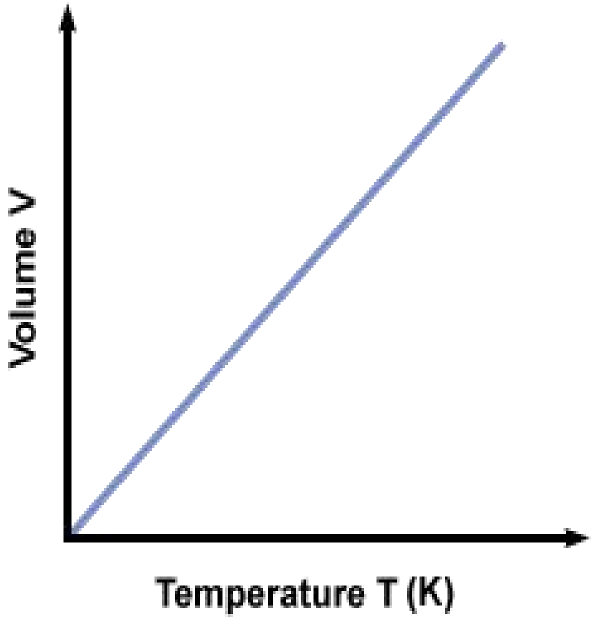
Charles' Law

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

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

$$V \propto T$$

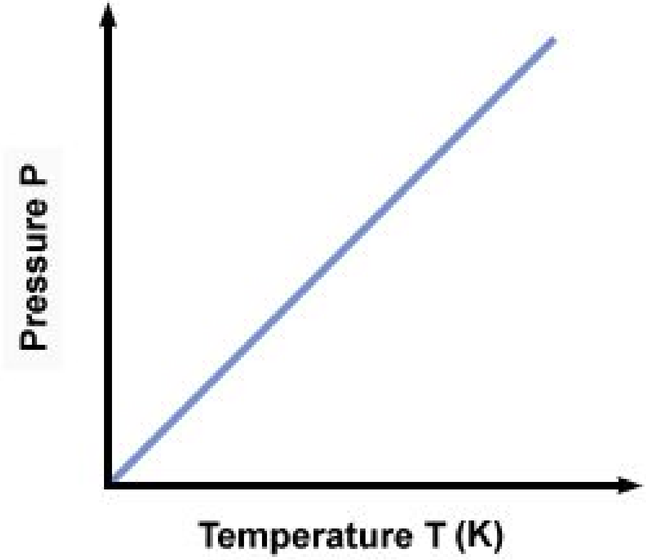


Gay-Lussac's Law

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

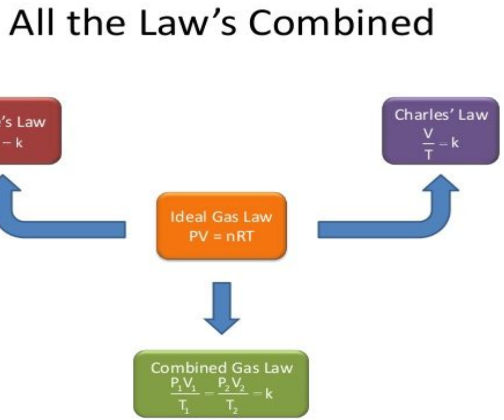
The pressure of a fixed mass of an ideal gas at a constant volume is directly proportional to its kelvin temperature

$$P \propto T$$



Combined Gas Law Equation

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$



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Ideal Gas Law

- Relationship between pressure, volume, temperature, and the amount, in mol, of gas particles

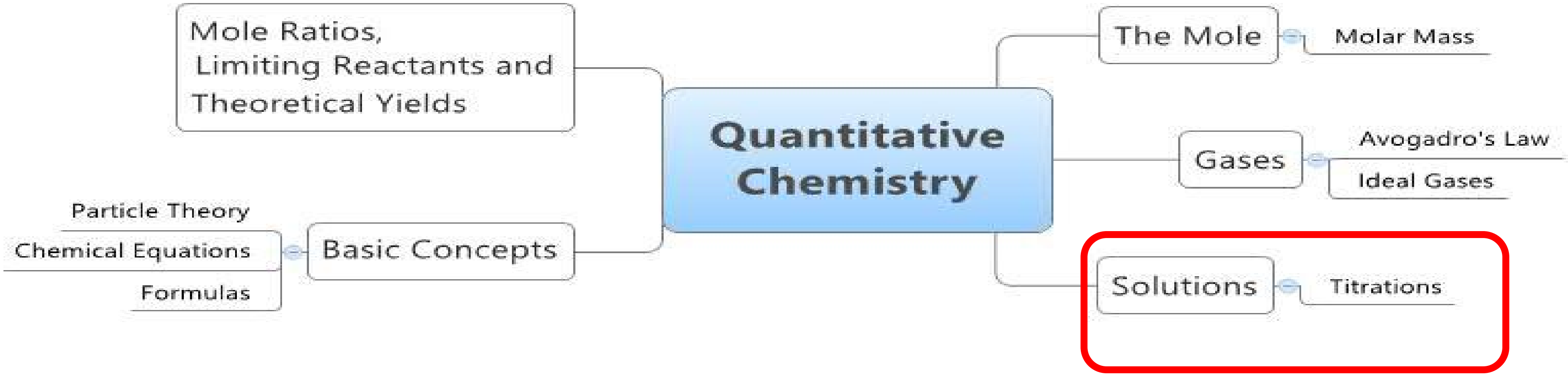
$$PV = nRT$$

$R = 0.0821\ L\ atm\ K^{-1} mol^{-1}$

Lesson 5

Solutions

We Are Here



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Solutions Basics

+

SOLUTE **SOLVENT** **SOLUTION**

A solution is a homogeneous mixture of a solute that has been dissolved in a solvent. When dissolved in water the solution is described as an aqueous solution.

Concentration

This is the strength of a solution.

Most Concentrated → Least Concentrated

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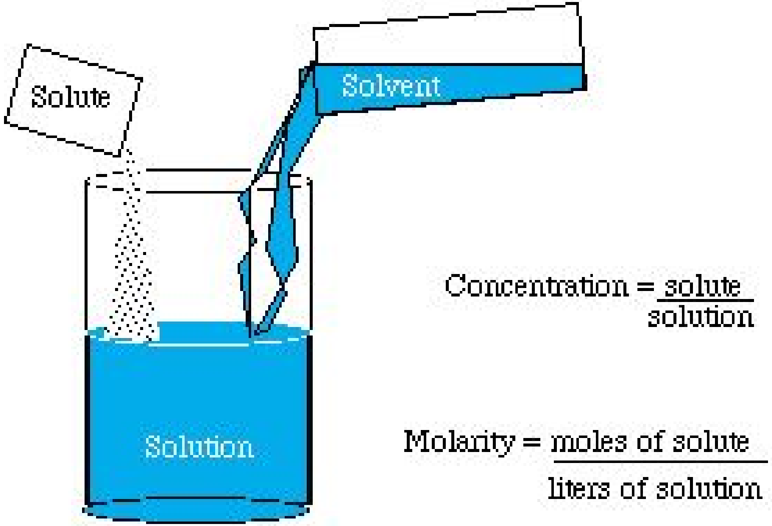
Concentration/Molarity

- The molar concentration of a solution is defined as the amount (in mol) of a substance dissolved in dm^3 of a solvent



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Concentration/Molarity

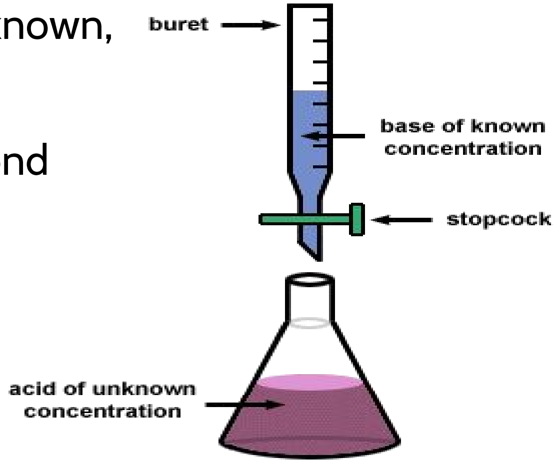
- Units:
  - mass per unit volume, g dm^{-3}
  - ***mole per unit volume, mol dm^{-3}***
  - Parts per million(ppm)
    - One part in 1×10^6 parts
    - 1 ppm=1 mg dm^{-3}

$$ppm = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

- Square brackets [ ] are used to denote molar concentration

Titrations

- Titration involves using a solution whose concentration is known, to find the concentration of another which isn't known.
- An exact volume of one solution is in a conical flask, a second solution is added to it from a burette.
- When the reaction reaches its 'endpoint', we record how much was added
  - There is always some kind of indicator which changes colour to tell us when we have reached the end.



- Determine the concentration of acids/bases
- Determine concentrations of other reactants
- Following the rate of a reaction
- Determining equilibrium constants

The mathematics of titrations

$$\frac{C_1V_1}{n_1} = \frac{C_2V_2}{n_2}$$

Where:
n = coefficient in balanced equation
C = concentration
V = volume

$$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$$

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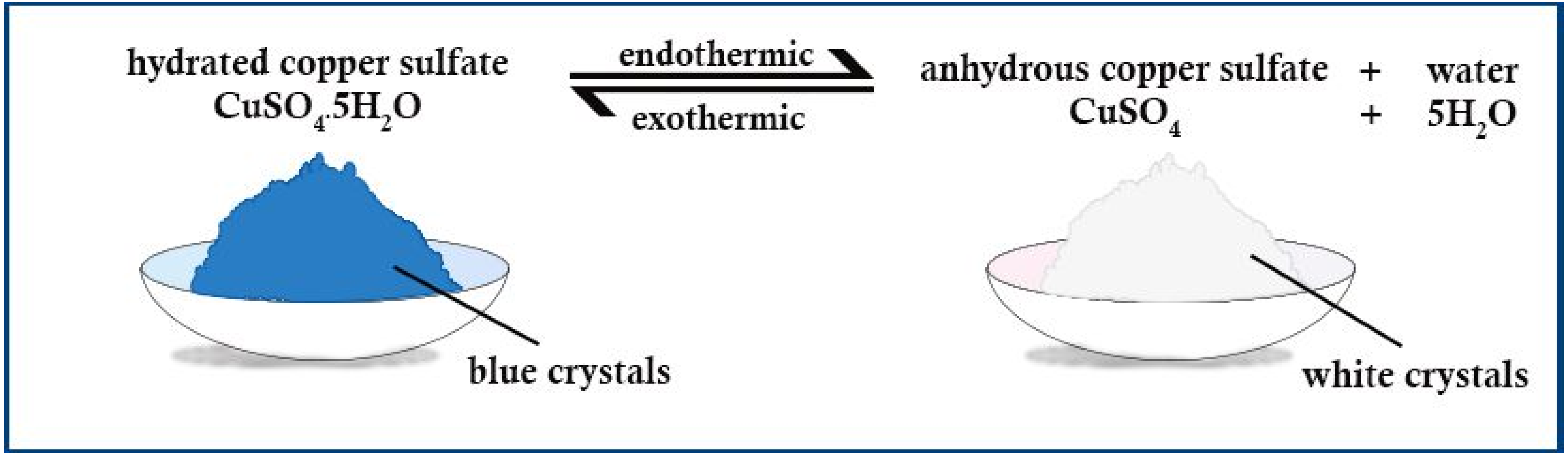
Water of Crystallisation

- Some substances crystallise with water, this is called water of crystallisation
- These substances are described as hydrated



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