Chemistry Transition Work 2025

Bridging the Gap to A-Level Chemistry

Introduction

The A-level chemistry course builds upon the foundations of the chemistry that you have studied at GCSE. In the following booklet we have identified the key topics that are especially important for A-level students to master.

These topics are also of vital importance to GCSE students who wish to maximise their grades ... these are the key topics that underpin the entire GCSE course.

These topics are:

Please use your existing GCSE resources, the Internet, the resources in Google Drive and consult with fellow students as you go along. You are expected to master all of these exercises and complete them to an excellent standard over the summer so you will be ready to make a confident and positive start to your A-level studies.

Best wishes and good luck from the Chemistry Department

Relative Atomic Mass (A_r)
Atoms of different elements have different masses because they contain different numbers of protons, neutrons and elecrons. Different isotopes of the same element contain different numbers of neutrons. The relative atomic mass of an element is the weighted mean mass of an atom of this element, including all its isotopes, compared to 1/12 th of the mass of an atom of ^{12}C.

e.g. A_r of Mg = 24 (mass of Mg atom = 2 x mass of ^{12}C atoms)

Relative Atomic Masses of Ions
Ions are atoms that have gained or lost electrons in order to achieve a full outer shell (noble gas configuration). Because electrons have very little mass compared to the rest of the atom (protons and neutrons), the mass of an ion is the same as the mass of the atom from which it was made.

e.g. A_r Cl = 35.5, A_r Cl^- = 35.5

Relative Molecular Mass (M_r)
The relative molecular mass, M_r, of a compound is equal to the sum of the relative atomic masses of the atoms in the numbers shown in the formula.

e.g. butane, C_4H_{10} $M_r = 4 \times 12 + 10 \times 1 = 58$

Calculating Masses of Atoms and Ions -- Exercise

1. Using a periodic table find the following atomic masses:

Hydrogen, H		Iodine, I		Oxygen, O	
Carbon, C		Nitrogen, N		Chlorine, Cl	

2. Calculate the Relative Molecular Mass (M_r) of the following:

Oxygen, O_2 Iodine, I_2 Water, H_2O

Methane, CH_4 Ethane, C_2H_6 Ethanol, C_2H_5OH

Calculating the Percentage Mass of an Element in a Compound -- Exercise

$$\% \text{ Mass of Element in Compound} = \frac{(\text{No. Atoms of Element in Compound}) \text{ x } (A_r \text{ of Element})}{M_r \text{ of Compound}} \text{ x } 100$$

e.g. SO_2 contains 1 atom of sulfur and 2 atoms of oxygen in each molecule (A_r: S = 32, O=16).

Total mass of O in $SO_2 = 2 \times 16 = 32$
M_r $SO_2 = 32 + 2 \times 16 = 64$

$$\therefore \% \text{ mass of O in } SO_2 = \frac{2 \times 16}{32 + 2 \times 16} \times 100 = 50\%$$

Total mass of S in $SO_2 = 32$

$$\therefore \% \text{ mass of S in } SO_2 = \frac{1 \times 32}{32 + 2 \times 16} \times 100 = 50\ \%$$

Method:

1. Write down the formula of the compound.
2. Using a periodic table, or a list of relative atomic masses (A_r), calculate the formula mass (M_r) of the compound.
3. Write down the total mass of the element you want, as a fraction of the M_r.
4. Multiply the fraction by 100, to give a percentage.

Student Tasks (Remember to Show Your Working out):

1. Find the percentage mass of carbon, and oxygen in carbon monoxide, CO.

___

___

2. Calculate the percentage mass of carbon, and oxygen, in carbon dioxide, CO_2.

___

___

3. Calculate the percentage mass of hydrogen, and carbon in methane, CH_4.

___

___

4. Find the percentage mass of hydrogen, and oxygen, in water, H_2O.

___

___

The Mole -- Exercise

One mole of a substance contains 6.022×10^{23} particles of the substance. One mole of a substance is obtained by weighing out the relative atomic mass, or formula mass, in grams.

$$\text{number of moles (mol)} = \frac{\text{actual mass (g)}}{\text{mass of 1 mole (g mol}^{-1}\text{)}}$$

actual mass (g) = number of moles (mol) $\times$ mass of 1 mole (g mol^{-1})

substance	mass of 1 mole (g)	mass of 3 moles (g)	mass of 0.5 moles (g)
O atoms	16 (RAM O)	48	8
O_2 molecules	32 (RMM O_2 = 16$\times$ 2)	96	16
H_2O molecules	18 (RMM H_2O = 16 + 2$\times$1)	54	9

e.g. How many moles of water molecules are there in 36g of water?
mass of water = 36 g M_r of water (H_2O) = 2 $\times$ 1 + 16 = 18
number of moles = 36/18 = 2 moles

Student Tasks:

1. Find the mass of 1 mole of (Hint -- A_r, M_r values in grams):

Sodium atoms, Na | Sodium ions, Na^+ | Chlorine atoms, Cl

---

2. Find the mass of 2 moles of (Hint -- use the correct equation above):

Oxygen atoms, O | Oxygen molecules, O_2 | Water molecules, H_2O

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3. Find the mass of 3 moles of (Hint -- use the correct equation above):

Carbon atoms, C | Methane molecules, CH_4 | Ethanol molecules, C_2H_5OH

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4. How many moles of atoms are there in (Hint -- use the correct equation above):

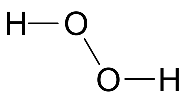
12 g of carbon | 16 g of oxygen

---

5. How many moles of molecules are there in (Hint -- use the correct equation above):

16 g of methane (CH_4) | 12 g of hydrogen (H_2)

Use of Reacting Masses to Find the Empirical Formulae of Compounds -- Exercise



Displayed Formula

H_2O_2

Molecular Formula

HO

Empirical Formula

The molecular formula is the total number of each type of atom present in a molecule. The **empirical formula** is the simplest whole number ratio of the elements present in a compound. It is possible to calculate the empirical formula if you know the mass of each element that has reacted to form the compound.

- Find the number of grams of each element that combine to form the compound.
- Calculate the number of moles of each element.
- Work out the ratio of the number of moles of each element, to the lowest whole numbers.
- Deduce the formula, from this ratio.

A compound of nitrogen and hydrogen was broken down into its elements. It was found that 1.4 g of nitrogen had combined with 0.3 g of hydrogen in the compound. What was the formula of the compound?

Actual mass N = 1.4 g
Actual mass C = 0.3 g

No. moles N = 1.4/14 = 0.1 mol
No. moles C = 0.3/1 = 0.3 mol

Molar ratio N:H = 0.1:0.3 = 1:3

Formula = NH_3

Note -- the formulae obtained from this method are empirical formulae (not molecular formulae):

- Molecular formula -- the total number of each type of atom present in a molecule
- Empirical formula -- the simplest whole number ratio of the elements present in a compound

Student Tasks:

1. Try to work out the empirical formulae of the compounds made from:

a. 12 g of carbon and 4 g of hydrogen.

____

____

b. 414 g of lead and 32 g of oxygen.

____

____

c. 3.2 g of copper, 0.6 g of carbon, and 2.4 g of oxygen.

____

____

____

Writing Symbol Equations

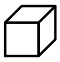
<u>Synthesis of hydrogen bromide</u>
hydrogen + bromine → hydrogen bromide
______H_2 (g) + ______Br_2 (g) → ______HBr (g)

<u>Combustion of carbon monoxide</u>
carbon monoxide + oxygen → carbon dioxide
______CO (g) + ______O_2 (g) → ______CO_2 (g)

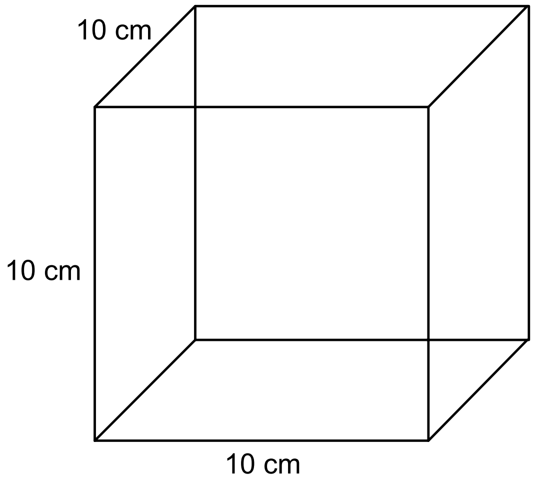
<u>Combustion of methane (natural gas)</u>
methane + oxygen → carbon dioxide + water
______CH_4 (g) + ______O_2 (g) → ______CO_2 (g) + ______H_2O (g)

<u>Decomposition of lead nitrate</u>
lead nitrate → lead oxide + nitrogen dioxide + oxygen
______$Pb(NO_3)_2$ (s) → ______PbO (s) + ______NO_2 (g) + ______O_2 (g)

Understanding Volumes



$1 \times 1 \times 1 = 1\ cm^3$



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

$1 \times 1 \times 1 = 1\ dm^3$

$10 \times 10 \times 10 = 1000\ cm^3$

Student Tasks:

1. Convert these volumes into cm^3: **(a)** 5.5 dm^3; **(b)** 2.5 L; **(c)** 0.75 dm^3. Hint -- 1 dm^3 = 1 L = 1000 cm^3.

2. Convert these volumes into dm^3: **(a)** 250 cm^3; **(b)** 20.4 cm^3; **(c)** 5.5 L. Hint -- 1 dm^3 = 1 L = 1000 cm^3.

Reacting Masses

1) What mass of carbon would be needed to make 8.8g of carbon dioxide?

$$C + O_2 \rightarrow CO_2$$

2) What mass of iron(II)sulfide is produced when 5.6g of iron completely reacts with excess sulfur?

$$Fe + S \rightarrow FeS$$

3) What mass of calcium carbonate would have to be completely decomposed by heat to give 280g of calcium oxide?

$$CaCO_3 \rightarrow CaO + CO_2$$

4) What mass of iron would be obtained from 160 tonnes of iron(III)oxide?

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

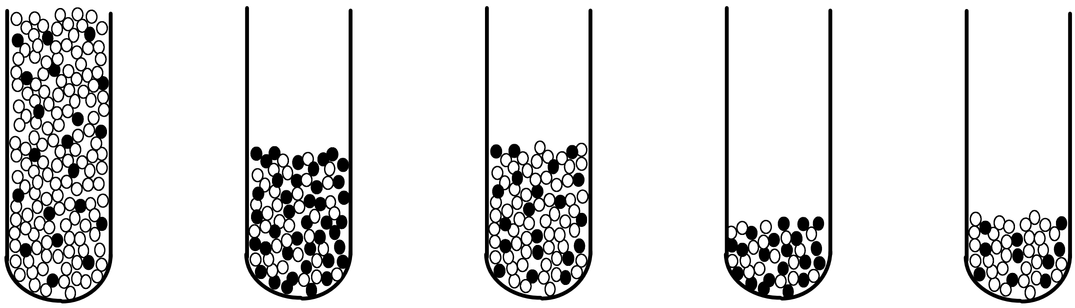
5) $CuCO_3 (s) \rightarrow CuO (s) + CO_2 (g)$ A_r: Cu = 64, C = 12, O = 16

If 31 g of $CuCO_3$ are used, how many grams of CO_2 will form? At room temperature and pressure = 1 atmosphere, 1 mole of any gas occupies 24 dm^3. Calculate the volume of CO_2 formed in this reaction.

6) $CaCO_3 (s) \rightarrow CaO (s) + CO_2 (g)$ A_r: Ca = 40, C = 12, O = 16

What mass of calcium oxide is produced from the thermal decomposition of 25 g calcium carbonate? At room temperature and pressure = 1 atmosphere, 1 mole of any gas occupies 24 dm^3. Calculate the volume of CO_2 formed in this reaction.

Understanding Molar Concentrations -- Exercise



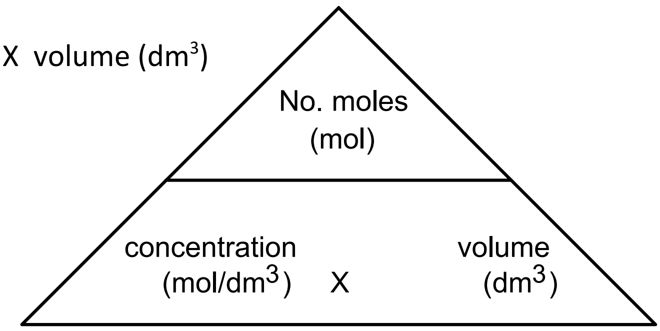
Amount Solute (mol)	1	2		1	0.5
Volume Solution (dm^3)	1	0.5	0.5		0.25
Concentration (mol/dm^3)			2	4	

Calculations Involving Molar Concentrations -- Exercise

number of moles (mol) = concentration (mol/dm^3) X volume (dm^3)

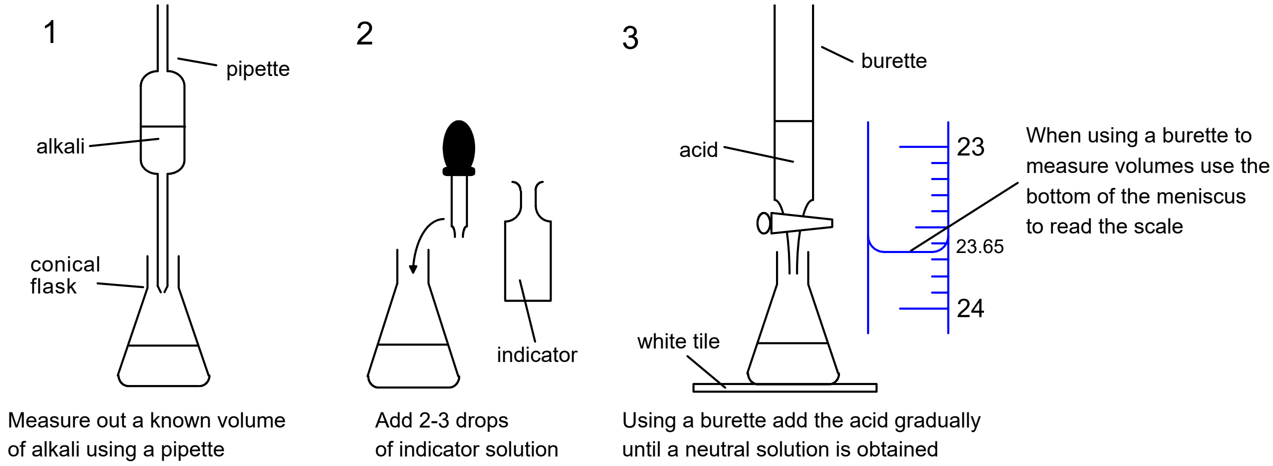
$$\text{concentration } (mol/dm^3) = \frac{\text{number of moles (mol)}}{\text{volume } (dm^3)}$$

$$\text{volume } (dm^3) = \frac{\text{number of moles (mol)}}{\text{concentration } (mol/dm^3)}$$



concentration (mol/dm^3)	number of moles (mol)	volume (dm^3)
2		1
	15	3
	4	4
0.5		2
1	0.5	
2	12	
1		7
8		0.25
	10	2
	0.5	1
	1	5
1		5

Performing Neutralisation Experiments Using Titration



Titration can be used to measure accurately the volumes of acid and alkali that react together to form a neutral solution. In a titration experiment a known volume of the alkali is measured into a conical flask using a pipette. 2-3 drops of a suitable indicator are then added to the alkali. The acid is then added from the burette until the indicator changes colour, indicating that the endpoint has been reached. The volumes of the acid and the alkali that have reacted together to form a neutral solution are now known. If the concentration of either the acid or the alkali is known, it is now possible to calculate the concentration of the other reactant.

Sodium Hydroxide and Hydrochloric Acid (Strong Alkali + Strong Acid)

25.0 cm^3 of 0.100M aqueous HCl was transferred into a conical flask using a pipette. Using litmus solution as the indicator, the HCl was titrated against 0.050M NaOH solution. Calculate the volume of NaOH used in the titration.

1. Calculate the number of moles of hydrochloric acid used in the titration.

2. Write a balanced equation for the reaction of HCl with NaOH.

3. Using the balanced equation, determine the molar ratio of HCl:NaOH.

4. Calculate the number of moles of sodium hydroxide used in the titration.

5. Calculate the volume of sodium hydroxide used in the titration.

Sodium Hydroxide and Sulfuric Acid (Strong Alkali + Strong Acid)

25.0 cm^3 of 0.100M aqueous H_2SO_4 was transferred into a conical flask using a pipette. Using litmus solution as the indicator, the H_2SO_4 was titrated against 50.0 cm^3 NaOH solution. Calculate the concentration of NaOH used in the titration.

1. Calculate the number of moles of sulfuric acid used in the titration.

2. Write a balanced equation for the reaction of H_2SO_4 with NaOH.

3. Using the balanced equation, determine the molar ratio of H_2SO_4:NaOH.

4. Calculate the number of moles of sodium hydroxide used in the titration.

5. Calculate the concentration of sodium hydroxide used in the titration.

Sodium Hydroxide and Nitric Acid (Strong Alkali + Weak Acid)

25.0 cm^3 of 0.100M aqueous HNO_3 was transferred into a conical flask using a pipette. Using phenolphthalein as the indicator, the HNO_3 was titrated against 20.0 cm^3 NaOH solution. Calculate the concentration of NaOH used in the titration.

1. Calculate the number of moles of nitric acid used in the titration.

2. Write a balanced equation for the reaction of HNO_3 with NaOH.

3. Using the balanced equation, determine the molar ratio of HNO_3:NaOH.

4. Calculate the number of moles of sodium hydroxide used in the titration.

5. Calculate the concentration of sodium hydroxide used in the titration.

Ammonia and Sulfuric Acid (Weak Alkali + Strong Acid)

20.0 cm^3 of 0.100M aqueous H_2SO_4 was transferred into a conical flask using a pipette. Using methyl orange as the indicator, the H_2SO_4 was titrated against 0.150M ammonia solution. Calculate the volume of ammonia used in the titration.

1. Calculate the number of moles of sulfuric acid used in the titration.

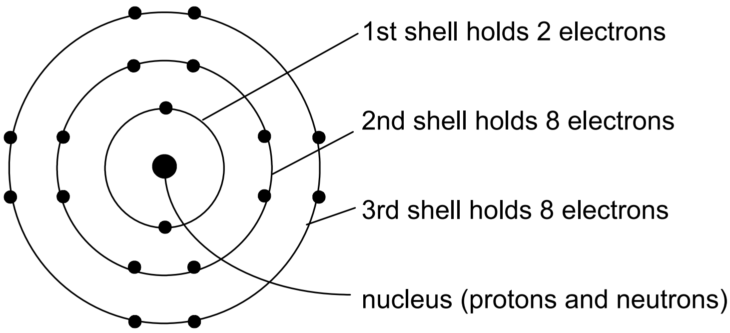
2. Write a balanced equation for the reaction of H_2SO_4 with NH_3.

3. Using the balanced equation, determine the molar ratio of H_2SO_4:NH_3.

4. Calculate the number of moles of ammonia used in the titration.

5. Calculate the volume of ammonia used in the titration.

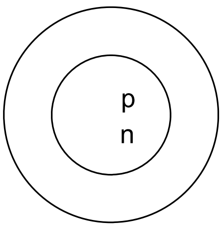
Drawing the Different Atoms -- Exercise



Oxygen

$^{16}_{8}\mathrm{O}$

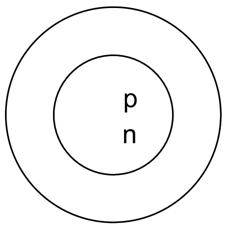
mass number =
atomic number =
protons =
neutrons =
electrons =



Nitrogen

$^{14}_{7}\mathrm{N}$

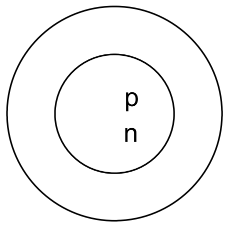
mass number =
atomic number =
protons =
neutrons =
electrons =



Fluorine

$^{19}_{9}\mathrm{F}$

mass number =
atomic number =
protons =
neutrons =
electrons =



Atomic Structure of the Elements and Some of their Isotopes -- Exercise

Element	Symbol	Number Protons	Number Neutrons	Number Electrons	Atomic Number	Atomic Mass
Hydrogen	H				1	1
Lithium	Li				3	7
Carbon	C	6	6	6		
Carbon-13	C				6	13
	C				6	14
Oxygen	O	8				16

Ionic Bonding

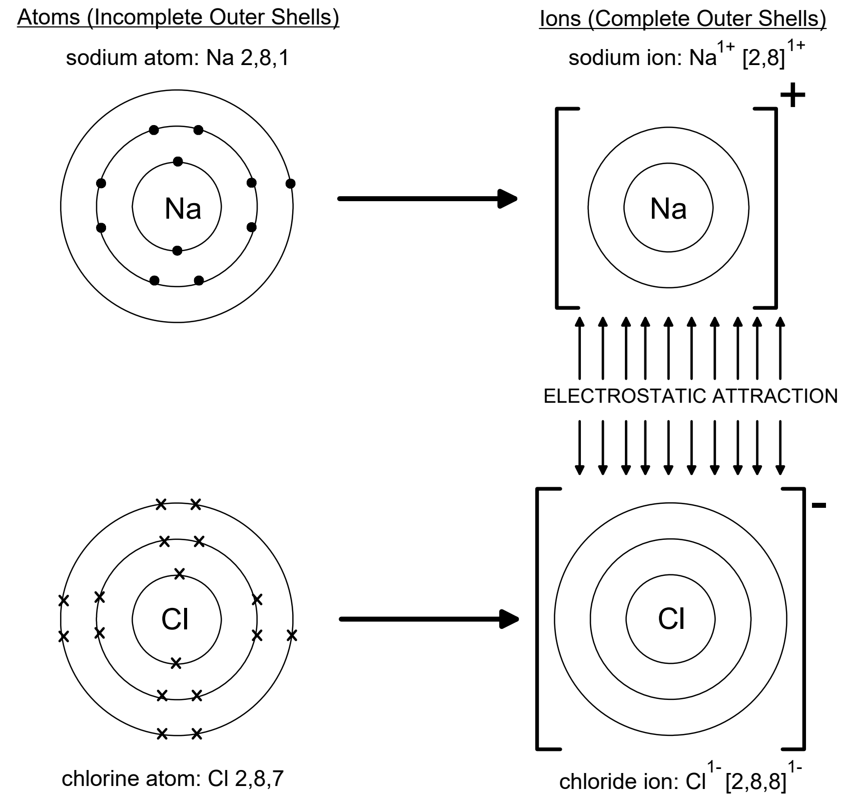
In the following section you are going to be asked to draw diagrams showing the electron arrangement of ionic compounds. Your diagrams should include the following features:

1. The electron arrangements of the 2 elements (use dots and crosses to distinguish between the electrons on each element).
2. Curly arrows to show the transfer of electrons from the metal to the nonmetal.
3. The electron arrangements of the 2 ions that are formed (use dots and crosses to clearly show the origin of the electrons).

4. The charge on the 2 ions.
5. The electrostatic attraction between the oppositely charged ions.
6. The formula of the ionic compound.

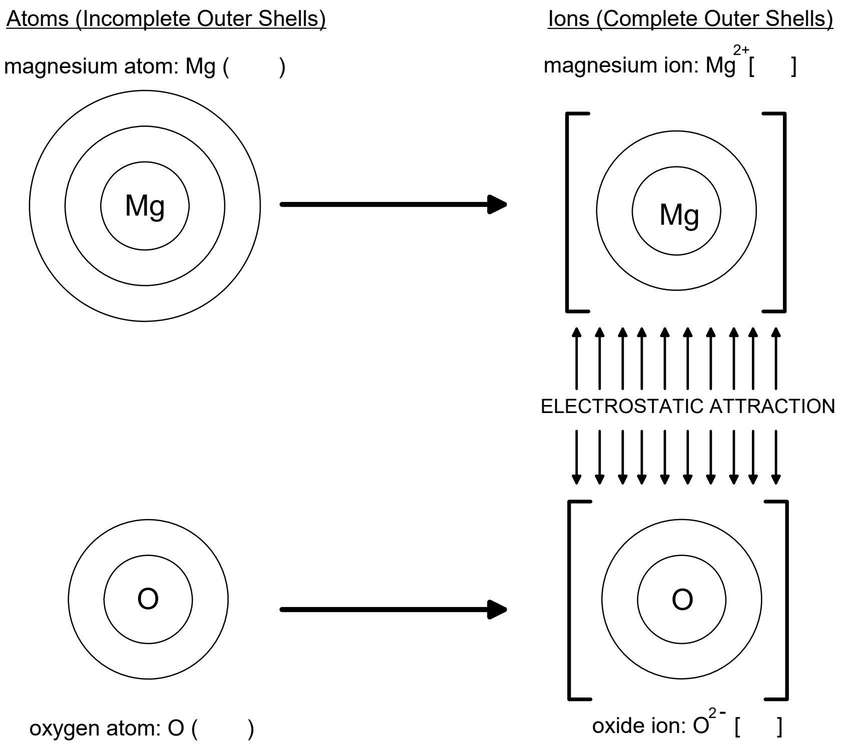
Student Tasks:

1. Complete the diagrams of the ionic bonding in sodium chloride ($^{23}_{11}Na$, $^{35}_{17}Cl$), magnesium oxide ($^{24}_{12}Mg$, $^{32}_{16}O$) and magnesium chloride ($^{24}_{12}Mg$, $^{35}_{17}Cl$).



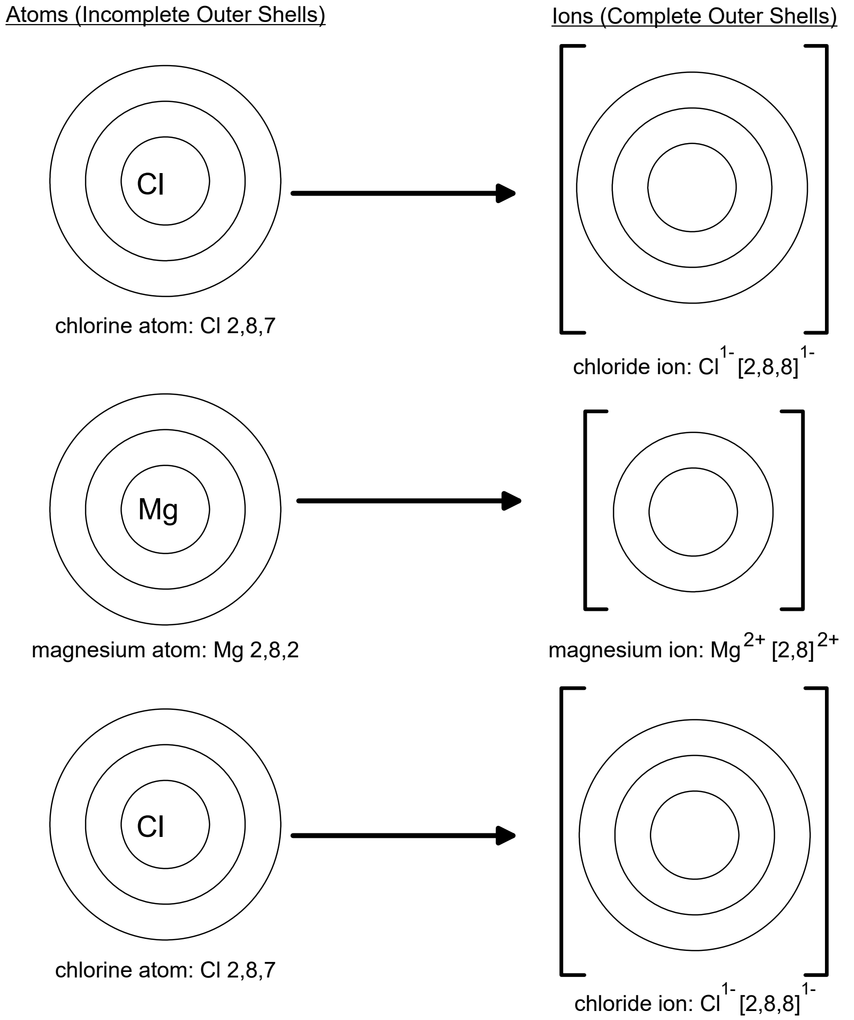
Checklist:
Correct electron arrangement of atoms
Correct use of dots and crosses
Arrow drawn from electron of metal to space on outer shell of non-metal
Correct electron arrangement of ions
Charge on ions
2,8,8 notation

Sodium Chloride
Formula --



Checklist:
Correct electron arrangement of atoms
Correct use of dots and crosses
Arrow drawn from electron of metal to space on outer shell of non-metal
Correct electron arrangement of ions
Charge on ions
2,8,8 notation

Magnesium Oxide
Formula --



Checklist:
Correct electron arrangement of atoms
Correct use of dots and crosses
Arrow drawn from electron of metal to space on outer shell of non-metal
Correct electron arrangement of ions
Charge on ions
2,8,8 notation

Magnesium Chloride
Formula --

Calculating the Formulae of Ionic Compounds

In ionic compounds positively charged metal ions combine with negatively charged nonmetal ions to form the ionic compound.

The ionic compound has no charge.

Hence, we need to have the correct number of positive and negative ions so that their charges cancel out.

E.g. Al^{3+} and F^{1-} $\qquad Al^{3+} + F^{1-} + F^{1-} + F^{1-} = AlF_3$
+++ - - - 0

E.g. Na^{1+} and O^{2-} $\qquad Na^{1+} + Na^{1+} + O^{2-} = Na_2O$
+ + - - 0

E.g. Mg^{2+} and NO_3^{1-} $\qquad Mg^{2+} + NO_3^{1-} + NO_3^{1-} = Mg(NO_3)_2$
++ - - 0

Exercise -- draw in the missing formulae of the compounds made from the ions in the box

IONS		COMPOUND
K^{1+}	Cl^{1-}	
Na^{1+}	O^{2-}	
Ca^{2+}	F^{1-}	
NH_4^{1+}	Cl^{1-}	
Al^{3+}	O^{2-}	
H^{1+}	SO_4^{2-}	
Mg^{2+}	CO_3^{2-}	
Al^{3+}	OH^{1-}	

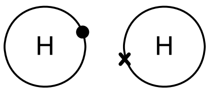
Covalent Bonding

Student Tasks:

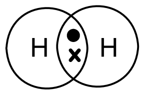
Complete the following exercise as follows:

- Indicate how many of electrons each atom needs to gain in order to get a full outer shell of electrons.
- Indicate how many covalent bonds each atom will form in order to get a full outer shell of electrons.
- Draw a diagram in which the outer shells of the component atoms overlap.
- Draw in the covalent bonds -- remember each covalent bond consists of a shared pair of electrons.
- Remember to use dots and crosses to show where the electrons came from.
- Check that each atom now has a full outer shell.
- Check that you have not created or destroyed any electrons.
- Draw the displayed formula of the compound.
- Write the molecular formula of the compound.

Atoms (Incomplete Outer Shells) | Molecules (Complete Outer Shells)

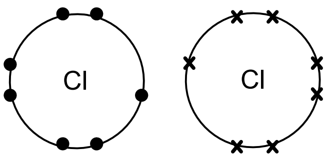


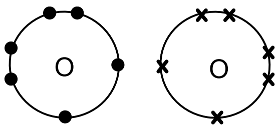
gain 1 electron
make 1 bond

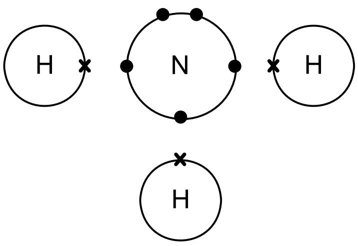


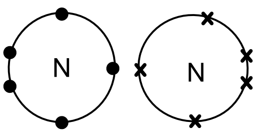
H—H

H_2

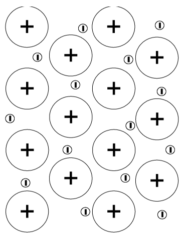
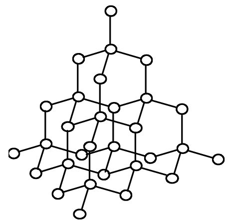
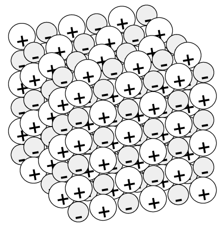
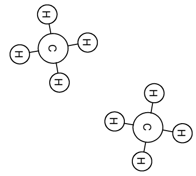








Properties of Materials -- Exercise

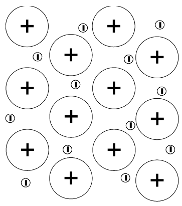
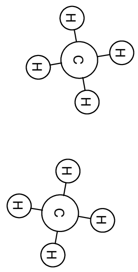
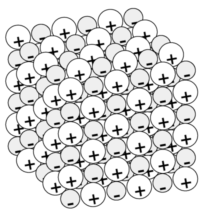
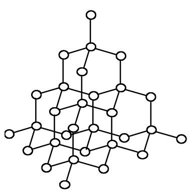
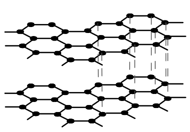
1 Structure				
(i) Examples (ii) Constituent particles (iii) Type of substance				
2 Bonding in the solid	Attraction of outer mobile electrons for positive ions -- strong metallic bonding	Atoms are linked through the whole structure by strong covalent bonds	Attraction of positive ions for negative ions -- strong ionic bonding	Strong covalent bonds hold atoms together within the molecules; weak intermolecular forces
3 Properties				
(i) Volatility	Non-volatile High melting & boiling point	Non-volatile Very high melting & boiling point	Non-volatile High melting & boiling point	Volatile Low melting & boiling point
State at room temperature	Usually solid	Solid	Solid	Usually gases or volatile liquids
(ii) Conductivity	Good conductors when solid or liquid	Non-conductors (graphite is an exception)	Non-conductors when solid. Good conductors when molten or in aqueous solution	Non-conductors when solid, liquid and in aqueous solution.
(iii) Hardness/ malleability	Hard, yet malleable	Very hard and brittle	Hard and brittle	Soft
(iv) Solubility	Insoluble in polar & non-polar solvents, but soluble in liquid metals -- making alloys	Insoluble in all solvents	Soluble in polar solvents (e.g. water), insoluble in non-polar solvents (e.g. hexane)	Usually insoluble in polar solvents (e.g. water), soluble in non-polar solvents (e.g. hexane)

Student Tasks:

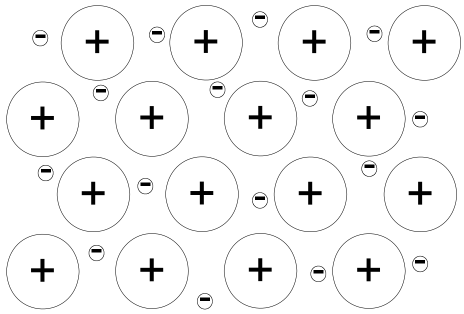
1. Complete the table by writing in the correct term (use the following hints):

- **Structure: (a)** giant ionic, **(b)** simple molecular, **(c)** giant molecular, **(d)** giant metallic.
- **Examples: (a)** I_2, S_8, O_2, CH_4, **(b)** NaCl, CaO, MgO, **(c)** Na, Fe, Cu, **(d)** diamond, graphite.
- **Constituent particles: (a)** atoms, **(b)** molecules, **(c)** ions, **(d)** atoms**.**
- **Type of substance: (a)** metal/non-metal compound, **(b)** non-metal element or non-metal/non-metal compound, **(c)** metal element, **(d)** non-metal element in group 4 or its compound.

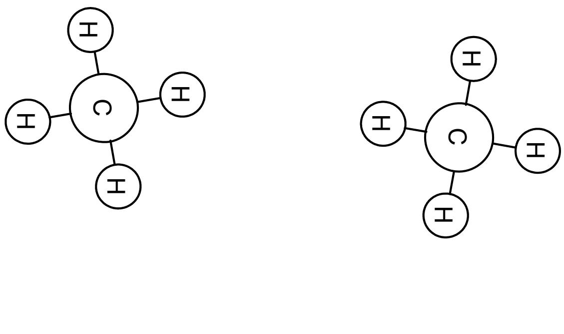
Properties of Materials -- Exercise -- Complete the table by inserting the appropriate words.

Type of Material	Solubility in Water (soluble or insoluble)	Solubility in Nonpolar Organic Solvent e.g. Oil (soluble or insoluble)	Conductivity (conductor or insulator)	Melting Point (high, medium or low)	Hardness (hard or soft)
metal (e.g. iron) 					
simple molecular substances (e.g. wax) 					
ionic (e.g. sodium chloride) 			solid: liquid: aqueous:		
giant molecular substances (e.g. diamond) 					
giant molecular substances (e.g. graphite) 					

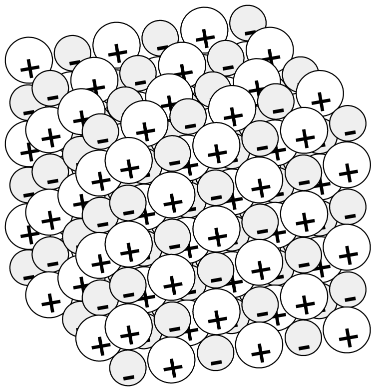
Structure and Bonding of Metals



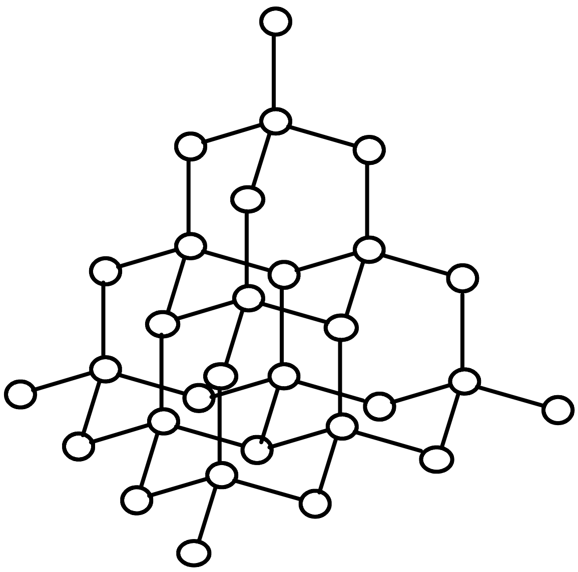
Structure and Bonding of Simple Molecular Substances



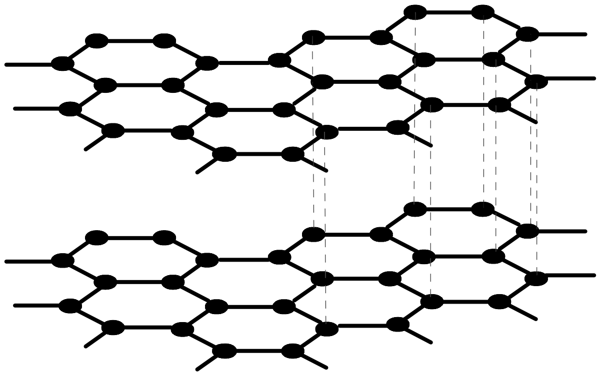
Structure and Bonding of Ionic Substances



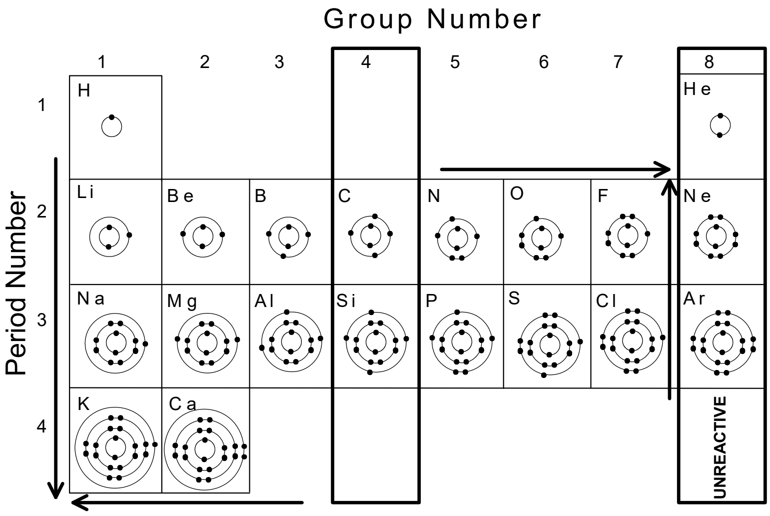
Structure and Bonding of Giant Molecular Substances



Structure and Bonding of Giant Molecular Substances



Reactivity of the Elements



As you go across a period, the number of electrons in the outer shell increases by 1 each time. As you go down a group, the number of electron shells increases by 1.

- Group number = number of outer electrons
- Period number = number of shells

Student Tasks:

1. Describe what happens to the number of protons as you go from 1 element to the next, across a period.

2. Describe how the electron configuration changes as you go: **(a)** across a period **(b)** down a group.

3. Explain why elements in the same group have similar chemical properties.

4. When elements react, they gain or lose electrons to achieve a full outer shell:

i. Describe & explain whether you would expect the metals to gain or lose electrons when they react.

ii. Describe & explain whether you would expect the nonmetals to gain or lose electrons when they react.

5. Describe and explain the chemical properties of the group 8 elements (Hints: noble/inert gases, electron arrangement).

6. Describe and explain the trends in chemical reactivity: **(a)** in group 1, **(b)** and in group 7 (Hints: metal or non-metal?, gain or lose electrons to get full outer shells?, electron arrangement, shielding electrons, how easy is it to gain or lose electrons?).

Reactivity Series	Water	Acid	Heating in Oxygen
potassium	melts + flame	dangerous	burns with lilac flame
sodium	melts	dangerous	burns with yellow flame
calcium	vigorous reaction	dangerous	burns with red flame
magnesium	slow reaction (vigorous with steam)	vigorous reaction	catches fire easily; burns with blinding white flame
aluminium	forms stable oxide layer	quite fast	forms stable oxide layer
carbon			
zinc	very slow reaction	quite slow	reacts slowly with oxygen
iron	rusts very slowly	slow reaction	does not burn; metal glows brightly
hydrogen			
copper	no reaction	no reaction	does not burn; metal becomes coated with black oxide layer
gold	no reaction	no reaction	no reaction

IMPORTANT HINT -- they do not all react!

metal + acid → metal salt + hydrogen

sodium + hydrochloric acid →

magnesium + nitric acid →

copper + sulfuric acid →

metal + water → metal hydroxide + hydrogen

lithium + water →

calcium + water →

copper + water →

metal + oxygen → metal oxide

magnesium + oxygen →

copper + oxygen →

gold + oxygen →

displacement reactions (more reactive metal wins)

copper oxide + potassium → potassium oxide + copper

iron oxide + magnesium →

copper oxide + aluminium →

magnesium oxide + copper →

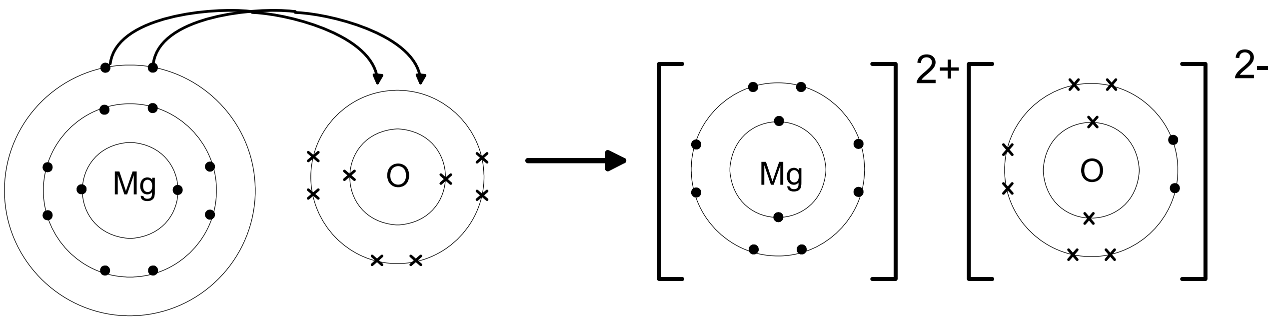
copper sulfate + sodium → sodium sulfate + copper

iron sulfate + magnesium →

copper sulfate + potassium →

magnesium sulfate + copper →

Redox Reactions



magnesium + oxygen → magnesium oxide
$2Mg\ (s) + O_2\ (g) \rightarrow 2MgO\ (s)$

In this reaction, the magnesium atom loses 2 electrons and the oxygen atom gains 2 electrons. The magnesium has been oxidised -- gained oxygen to form MgO.

copper oxide + hydrogen → copper + water
$CuO\ (s) + H_2\ (g) \rightarrow Cu\ (s) + H_2O\ (g)$

In this reaction, the copper ions in copper oxide have gained electrons to form copper atoms:
$Cu^{2+} + 2e^- \rightarrow Cu$

The copper oxide has been reduced -- lost oxygen to form copper.

Using these observations, we can now produce definitions of oxidation and reduction that can be applied to reactions that do not involve oxygen:

1. **Oxidation** -- loss of electrons;
2. **Reduction** -- gain of electrons.

In a reaction, if one species is _____________ (electron loss), another species must be reduced (_____________ gain). It is not possible to have oxidation without _____________. Reactions that involve oxidation and reduction are called _____________ reactions.
OILRIG -- _____________ is _____________ , _____________ is _____________.

A species has been oxidised if it has:	A species has been reduced if it has:
gained/lost oxygen	gained/lost oxygen
gained/lost hydrogen	gained/lost hydrogen
gained/lost electrons	gained/lost electrons

Hints: redox reaction; reducing agent; oxidation; oxidising agent; reduction

	the loss of electrons from a species
	the gain of electrons by a species
	a substance that removes electrons from a species
	a substance that donates electrons to a species
	a reaction that involves oxidation and reduction

Reactions of Strong and Weak Acids

Hints: all, strong, some, weak, hydrochloric, ethanoic.

In a _______ acid (e.g. _______ acid, sulphuric acid and nitric acid) nearly _______ of the acid molecules split up to produce hydrogen ions. In a _______ acid (e.g. _______ acid, citric acid and carbonic acid), only _______ of the acid molecules split up to produce hydrogen ions -- carboxylic acids do not dissociate fully in water.

HCl (aq) ⇌ H^+ (aq) + _______
0.01% 99.99%

CH_3COOH (aq) ⇌ H^+ (aq) + _______
99.95% 0.05%

Student Tasks:
1. Make a list of the strong acids and the weak acids.

___

___

2. What is the difference between a strong acid and a weak acid.

___

___

___

___

How Are pH Values Related to the Concentration of Hydrogen Ions?

The pH of a 1.0 mol/dm^3 solution of hydrochloric acid is equal to 0.
The pH of a 0.1 mol/dm^3 solution of hydrochloric acid is equal to 1.
Hydrochloric acid is a strong acid which fully dissociates in water.
Hence, a 1.0 mol/dm^3 solution of hydrochloric acid contains _________ mol/dm^3 H^+ (aq) ions.

The pH of the solution depends on the concentration of the H^+ (aq) ions.
As you decrease the concentration of H^+ (aq) ions by a factor of 10, the pH _________ by 1.
As you increase the concentration of H^+ (aq) ions by a factor of 10, the pH _________ by 1.

HCl concentration in mol/dm^3	H^+ (aq) concentration in mol/dm^3	Number of decimal places	H^+ (aq) concentration in mol/dm^3 in standard form	pH value
1	1	0	10^0	0
0.1	0.1			
0.01		2	10^{-2}	2
0.001	0.001			
0.0001		4	10^{-4}	4
0.00001	0.00001			
0.000001		6	10^{-6}	6
0.0000001	0.0000001			

Producing Salts 1

$Mg + H_2SO_4 \rightarrow$ _____________ + _____________

$MgO + H_2SO_4 \rightarrow$ _____________ + _____________

$NaOH + HCl \rightarrow$ _____________ + _____________

$CaCO_3 + H_2SO_4 \rightarrow$ _____________ + _____________ + _____________

$Mg + 2H^+ \rightarrow Mg^{2+} +$ _____________

$O^{2-} + 2H^+ \rightarrow$ _____________

$OH^- + H^+ \rightarrow$ _____________

$CO_3^{2-} + 2H^+ \rightarrow$ _____________ + _____________

General Equations

acid + metal → _____________ _____________ + _____________

acid + metal oxide → _____________ _____________ + _____________

acid + alkali → _____________ _____________ + _____________

acid + metal carbonate → _____________ + _____________ + _____________

Producing Salts 2

Acid	ion 1	ion 2	Formula	Salt
hydrochloric acid	H^+		HCl	
sulfuric acid	H^+		H_2SO_4	
nitric acid	H^+		HNO_3	

hydrochloric acid + magnesium → _____________ _____________ + hydrogen

sulfuric acid + lithium → _____________ _____________ + _____________

nitric acid + _____________ → iron nitrate + hydrogen

nitric acid + sodium oxide → _____________ _____________ + water

nitric acid + calcium oxide → _____________ _____________ + _____________

hydrochloric acid + sodium hydroxide → ____________ chloride + water

sulfuric acid + sodium hydroxide → ____________ ____________ + ____________

____________ ____________ + calcium hydroxide → calcium nitrate + water

hydrochloric acid + lithium carbonate → ____________ ____________ + carbon dioxide + water

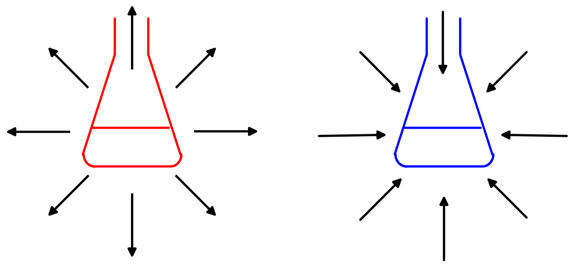
sulfuric acid + sodium carbonate → _________ _________ + _________ _________ + _________

Exothermic and Endothermic Reactions

Hints: produce, endothermic, EXITS, from, ENTERS, heat, warm, cold, surroundings, to or from.

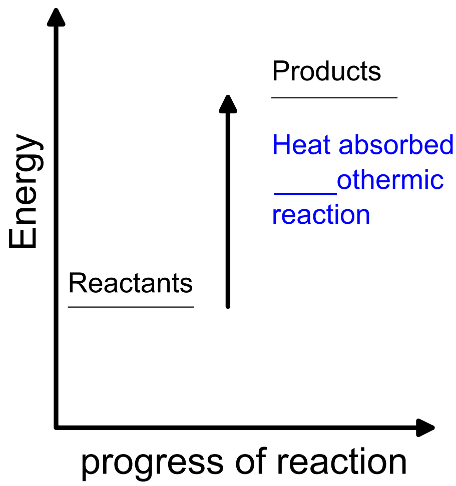
When a chemical reaction occurs heat is transferred __________ the surroundings. Exothermic reactions __________ heat which they give out to the __________. They feel __________. Some reactions feel __________, because they take __________ from their surroundings. These are called __________ reactions.

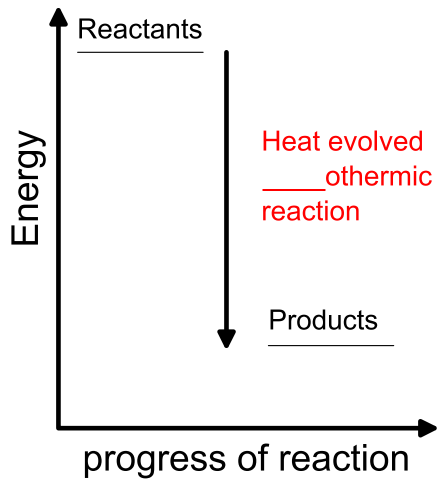
- EXothermic reaction -- heat __________ from the reaction and is given off to the surroundings.
- ENdothermic reaction -- heat is taken in (__________) from the surroundings.



EXothermic ENdothermic

Heat Changes in Chemical Reactions





<u>Exothermic and Endothermic Reactions</u>

- If the products have more/less energy than the reactants, the difference in energy is given out as heat. The reaction is endothermic/exothermic.
- If the products have more/less energy than the reactants. The extra energy needed to form the products is taken from the surroundings. The reaction is exothermic/endothermic.

Making and Breaking Bonds

Hints: made, endothermic, bond energy, attraction, exothermic, more, 1 mole, energy, magnets, swap, broken, less, stronger.

In chemical reactions atoms ___________ partners. This means that the bonds which join atoms to each other must be ___________. New bonds must be ___________ as the products form.

Bond Breaking and Bond Making -- breaking a bond is like pulling apart 2 strong ___________. Bond breaking requires ___________. It is ___________. Bond making is like 2 strong magnets coming together because of their ___________ to each other. Making new bonds gives out energy. It is ___________.

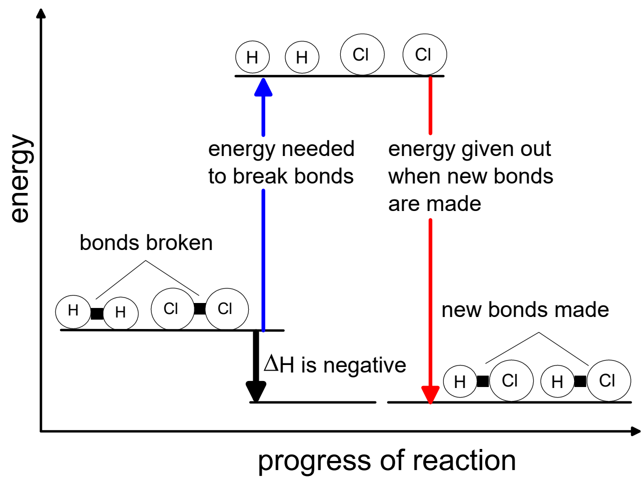
Bond Energy -- the energy required to break different bonds is called the ___________. Some bonds are ___________ than others. The units are kJ/mol (kilojoules per mole) -- this is the energy required to break ___________ of bonds.

Energy Change(ΔH) = Energy Required to Break Bonds - Energy Released when New Bonds are Formed

- **Exothermic reactions (negative value)**: ___________ energy is given out making new bonds than was required to break the original bonds.
- **Endothermic reactions (positive value)**: ___________ energy is given out making new bonds than was required to break the original bonds.

Energy Level Diagrams

The energy changes that occur during a reaction can be shown on an energy level diagram.



Student Tasks:

1. In this reaction, what involves more energy -- breaking the original bonds (endothermic) or making new bonds (exothermic)?

____

____

2. Is the overall reaction exothermic or endothermic?

____

____

3. Some energy will need to put in to get the reaction started -- why?

____

____

How Catalysts Work

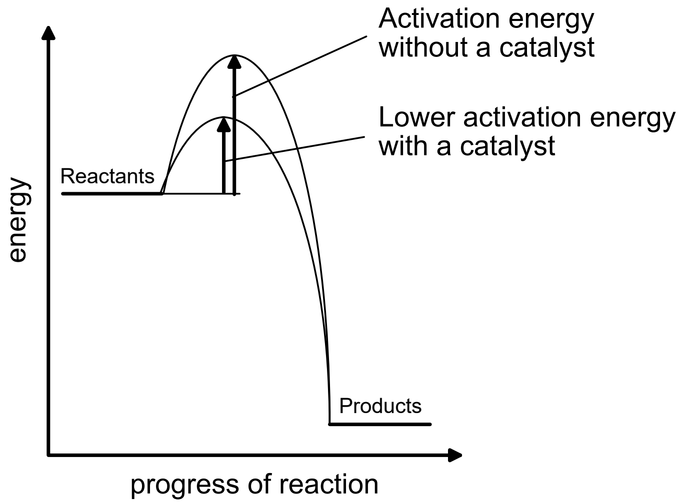
Hints: activation energy, increase, lowering, break, successful, minimum.

The energy needed to start a reaction is called the __________.

The activation energy is the __________ energy required by the reactant molecules before a reaction can occur.

The activation energy is the energy needed to __________ some of the original bonds, and get the reaction started.

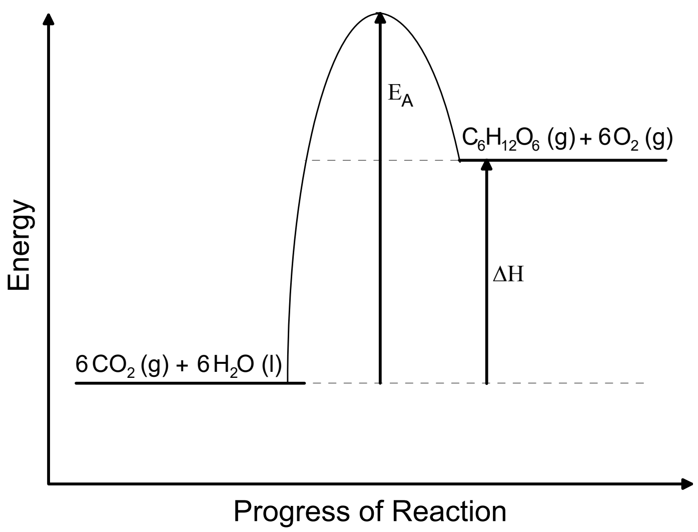
Catalysts __________ the rate of a reaction by __________ the activation energy, and increasing the proportion of collisions that are __________.



Heat Changes in Chemical Reactions

1. Label the diagram to show whether it represents: **(a)** photosynthesis, **(b)** combustion of hexane, **(c)** respiration, **(d)** thermal decomposition of calcium carbonate.
2. Indicate whether the reaction is exothermic or endothermic.
3. Draw in arrows to show the activation energy (E_a) and the energy change (ΔH) on the graphs.
4. Write a balanced equation for the reaction.

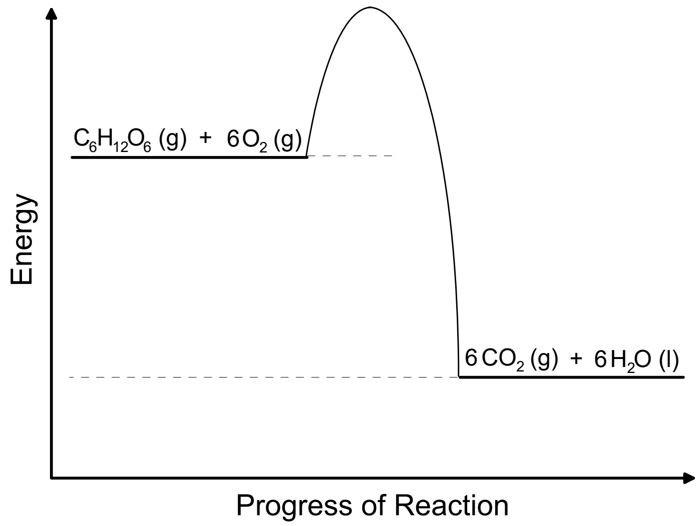
   Note – the arrow showing the activation energy must be an upwards arrow drawn from the energy of the reactants to the highest point of the curve. The arrow showing the energy change must be drawn from the energy of the reactants to the energy of the products. For endothermic reactions, where the energy has increased, an upwards arrow should be used. For exothermic reactions, where the energy has decreased, a downwards arrow should be used.



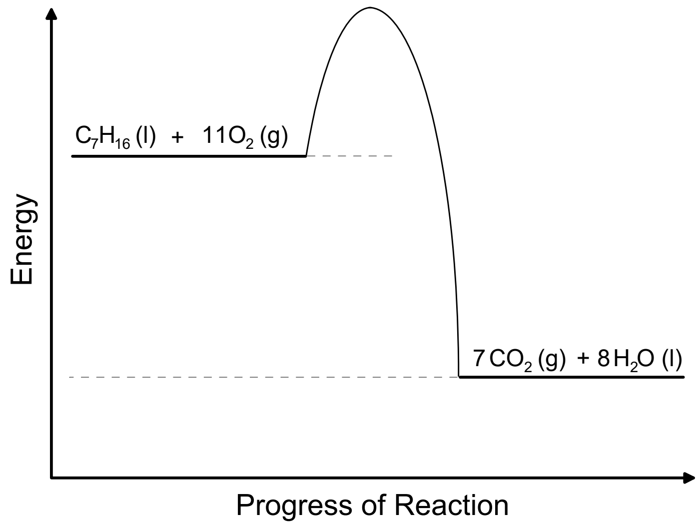
Photosynthesis

Exothermic -- more energy released making new bonds than was used up breaking the original bonds

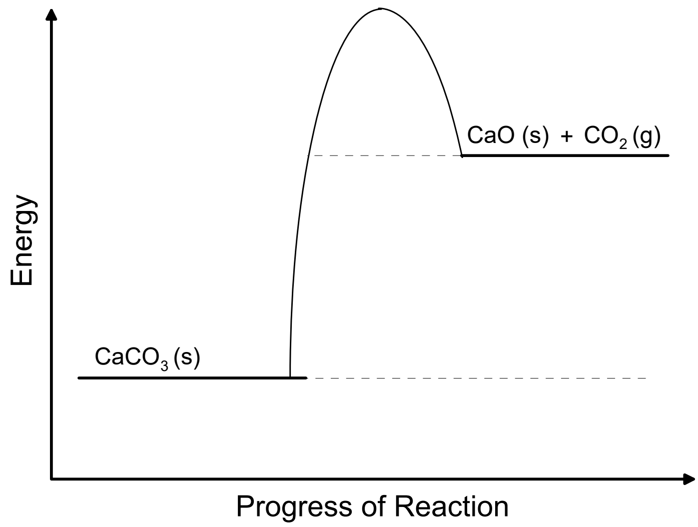
$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$



Energy
$C_6H_{12}O_6$ (g) + 6 O_2 (g)
6 CO_2 (g) + 6 H_2O (l)
Progress of Reaction



Energy
C_7H_{16} (l) + 11 O_2 (g)
7 CO_2 (g) + 8 H_2O (l)
Progress of Reaction



Energy
CaO (s) + CO_2 (g)
$CaCO_3$ (s)
Progress of Reaction

Calculating Enthalpy Changes from Bond Enthalpy Data

Student Tasks:

1. Rewrite the equation to show the bonds present in each molecule
2. List the bonds broken during the reaction
3. List the bonds made during the reaction
4. Calculate the energy required to break all the bonds in the reactants
5. Calculate the energy produced in forming all the bonds in the products
6. Calculate the energy change of the reaction using the following equation:

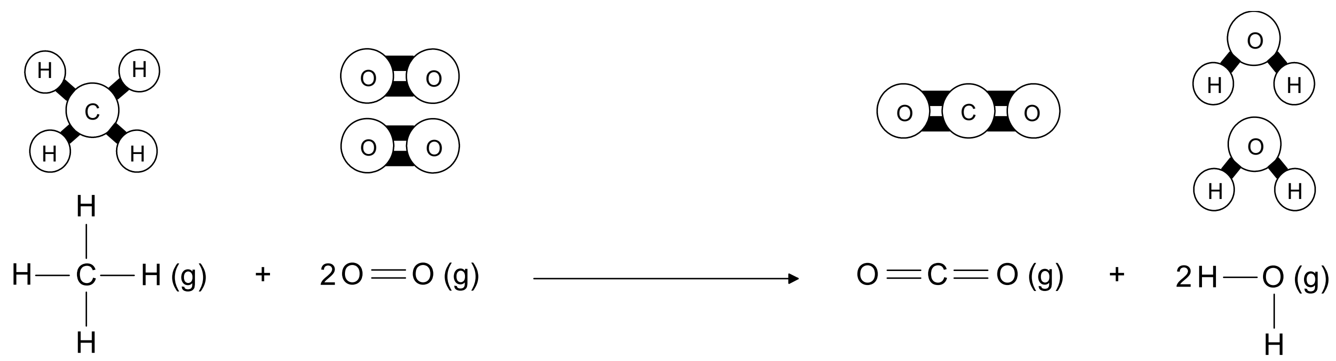
ΔH = Energy Used to Break Bonds - Energy Released by Bond Making

7. Explain whether the reaction is exothermic ($\Delta H<0$) or endothermic ($\Delta H>0$)

Calculating Energy Changes from Bond Energy Data 1

$CH_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O (l)$

Bond	C-H	O=O	O-H	C=O
Bond Energy (kJ/mol)	416	497	485	804



Bonds Broken

=

Bonds Made

=

Energy Used to Break Bonds

=

=

Energy Released by Bond Making

=

=

ΔH = Energy Used to Break Bonds - Energy Released by Bond Making

=

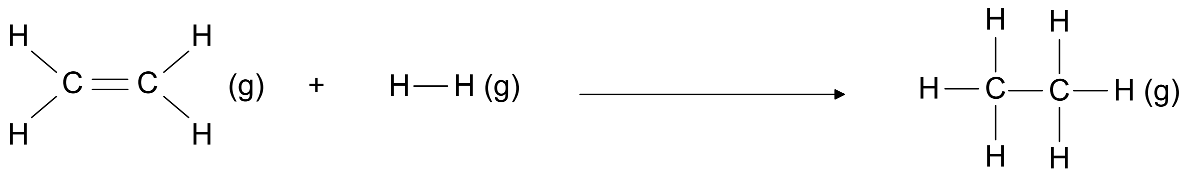
=

The reaction is exothermic. More energy is given out making new bonds, than is taken in breaking the original bonds. Energy is given out to the surroundings.

Calculating Energy Changes from Bond Energy Data 2

C_2H_4 (g) + H_2 (g) → C_2H_6 (g)

Bond	C-H	C=C	H-H	C-C
Bond Energy (kJ/mol)	416	611	436	330



Bonds Broken

=

Bonds Made

=

Energy Used to Break Bonds

=

=

Energy Released by Bond Making

=

=

ΔH = Energy Used to Break Bonds - Energy Released by Bond Making

=

=

The reaction is __________. __________energy is given out making new bonds, than is taken in breaking the original bonds. Energy is ____________________ the surroundings.

Calculating Energy Changes from Bond Energy Data 3

H_2 (g) + Cl_2 (g) → 2HCl (g)

Bond	H-H	Cl-Cl	H-Cl
Bond Energy (kJ/mol)	436	242	431

H—H (g) + Cl—Cl (g) → 2 H—Cl (g)

Bonds Broken

=

Bonds Made

=

Energy Used to Break Bonds

=

=

Energy Released by Bond Making

=

=

ΔH = Energy Used to Break Bonds - Energy Released by Bond Making

=

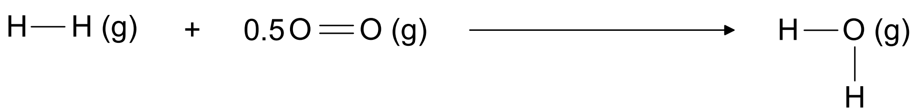
=

The reaction is ___________. ___________energy is given out making new bonds, than is taken in breaking the original bonds. Energy is ______________________ the surroundings.

Calculating Energy Changes from Bond Energy Data 4

H_2 (g) + $1/2O_2$ (g) $\rightarrow$ H_2O (g)

Bond	H-H	O=O	O-H
Bond Energy (kJ/mol)	436	497	485



Bonds Broken

=

Bonds Made

=

Energy Used to Break Bonds

=

=

Energy Released by Bond Making

=

=

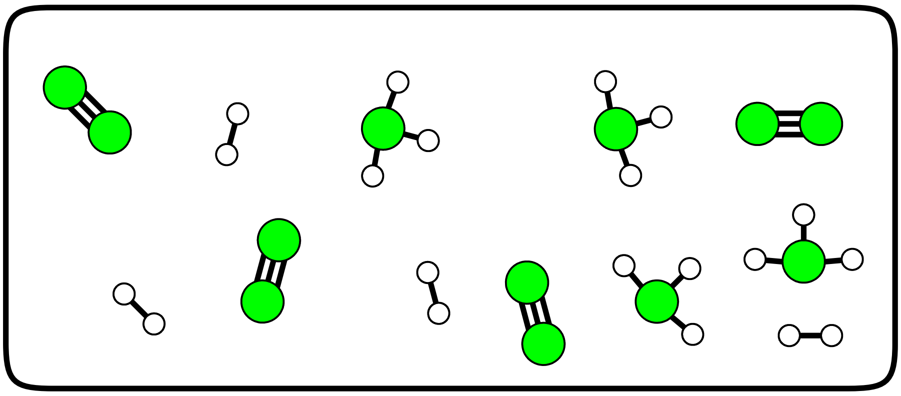
ΔH = Energy Used to Break Bonds - Energy Released by Bond Making

=

=

The reaction is ___________. ___________energy is given out making new bonds, than is taken in breaking the original bonds. Energy is ______________________ the surroundings.

Manufacturing Ammonia using the Haber Process



$$N_2\,(g) + 3H_2\,(g) \overset{Fe}{\rightleftharpoons} 2NH_3\,(g)$$

This reaction is reversible. It does not go to completion. Some ammonia molecules break down into nitrogen and hydrogen, whilst other ammonia molecules form. This results in an equilibrium mixture of nitrogen, hydrogen and ammonia. At equilibrium, the rates of the forward and reverse reactions are exactly the same AND the amounts of nitrogen, hydrogen and ammonia remain constant.

Exam note – at equilibrium the rate of the forward and reverse reactions are equal AND the amounts of reactants and products are constant BUT the amounts of reactants and products are NOT equal.

Student Tasks:

1. Describe what is happening in an equilibrium mixture of ammonia, nitrogen and hydrogen (Hints: What is happening to the ammonia?; What is happening to the nitrogen and hydrogen?).

2. What can you say about the rates of the forward and reverse reaction for a system at equilibrium (Hint -- at equilibrium the reaction is in balance)?

3. What can you say about the amounts of the reactants and products in an equilibrium mixture (Hint -- NOT the same)?

Le Chatelier's Principle

Le Chatelier's principle states:

If a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium shifts so as to cancel out the change.

- If you increase the concentration of the reactants…the equilibrium will shift in the forward direction…lowering the concentration of the reactants
- If you increase the concentration of the products…the equilibrium will shift in the reverse direction…lowering the concentration of the products

$$N_2\,(g) + 3H_2\,(g) \overset{Fe}{\rightleftharpoons} 2NH_3\,(g)$$

Student Tasks:

1. An equilibrium mixture is set up in a closed system with nitrogen, hydrogen and ammonia. In order to make more ammonia **which substances**: would you continually add to the mixture; would you continually remove from the mixture? Explain your answer using Le Chatelier's principle.

____

____

____

____

____

2. What would happen to the position of the equilibrium if you added ammonia to the reaction mixture? Explain your answer using Le Chatelier's principle.

____

____

____

____

____

Effect of Temperature on Equilibria

Nitrogen dioxide molecules can combine to form dinitrogen tetroxide, N_2O_4. In a closed system, an equilibrium is established between the pale yellow dimer and the dark brown monomer. The colour of the resultant mixture can be used to monitor the equilibrium position.

N_2O_4	$\rightleftharpoons$	$2NO_2$ (g)	ΔH^{θ} = +57 kJ mol^{-1}
pale yellow		dark brown	

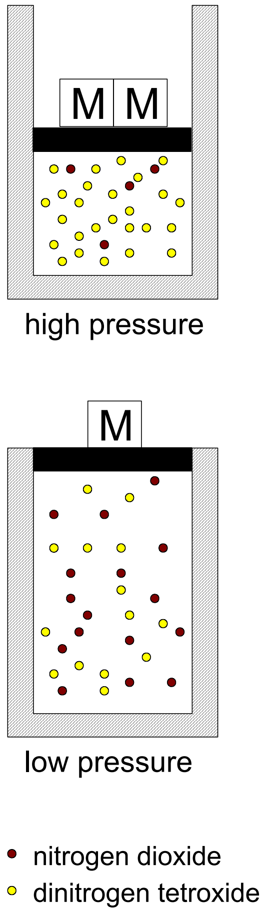
Le Chatelier's principle – when a change in conditions is introduced to a system at equilibrium, the position of the equilibrium shifts so as to cancel out the change. If you increase the temperature the equilibrium shifts in the direction of the endothermic reaction and more nitrogen dioxide is formed (dark brown).

If you decrease the temperature, the equilibrium shifts in the direction of the exothermic reaction and more dinitrogen tetroxide is formed (pale yellow).

Student Tasks:

1. Describe what you would expect to happen to the colour as the mixture is: **(a)** cooled, **(b)** heated. Explain your answers in terms of: bond making/breaking; exothermic/endothermic processes; equilibrium position; Le Chatelier's principle.

Effect of Pressure on Equilibria



In an equilibrium mixture of nitrogen dioxide and dinitrogen tetroxide, nitrogen dioxide molecules continually react to form dinitrogen tetroxide, whilst dinitrogen tetroxide continually breaks down to reform nitrogen dioxide.

Increasing the pressure shifts the equilibrium in the direction that has the fewest gas particles (to counteract the pressure increase). More dinitrogen tetroxide is formed (pale yellow).

Decreasing the pressure shifts the equilibrium in the direction that has the most gas particles (to counteract the decrease in pressure). More nitrogen dioxide is formed (dark brown).

N_2O_4	$\rightleftharpoons$	$2NO_2$ (g)	$\Delta H^{\theta} = +57$ kJ mol^{-1}
pale yellow		dark brown	

Le Chatelier's principle – when a change in conditions is introduced to a system at equilibrium, the position of the equilibrium shifts so as to cancel out the change.

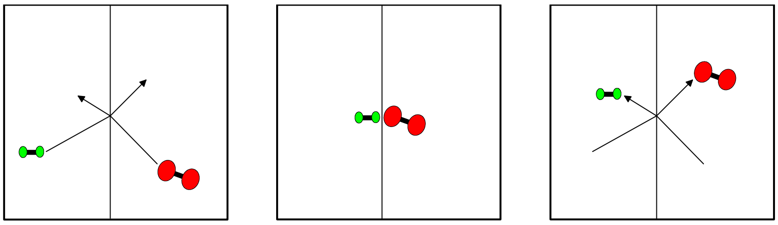
Student Tasks:

1. Describe what you would expect to happen to the colour as the pressure is: **(a)** increased, **(b)** decreased. Explain your answers in terms of: counteracting changes in gas pressure; number of particles in gas phase; equilibrium position; Le Chatelier's principle.

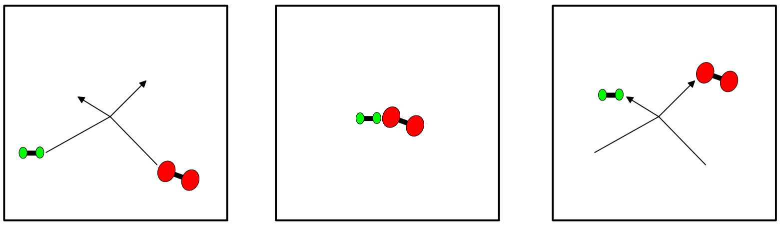
Understanding Rates of Reactions

In order for substances to react with each other two things need to happen:

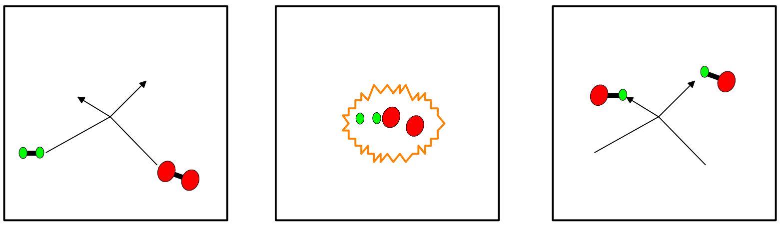
1. the particles of the substances must collide with each other.
2. the collision must have enough energy to overcome the activation energy of the reaction.



No reaction can occur between these substances because their particles do not come into contact.



In this example, the particles of the substance do collide. However, there is still no reaction because the collision did not have sufficient energy to overcome the activation energy of the reaction -- the particles bounced off each other.



The particles have collided with each other with enough energy to break their original bonds (activation energy). The particles then form new bonds, and combine together to make new compounds.

Student Tasks:

1. Identify two things that need to happen for a reaction to occur.

2. What is: **(a)** a successful collision, **(b)** an unsuccessful collision?

3. Explain how you could change the conditions of a chemical reaction in order to make the particles collide with each other more frequently.

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4. Explain how you could change the conditions of a chemical reaction in order to make the particles collide with each other with more energy.

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Altering the Rate of a Chemical Reaction 1

The rate of a reaction depends on the number of successful collisions that occur within a given time period. Hence, there are two main ways of increasing the rate of a reaction:

- Increase the collision frequency (make the particles collide with each other more often).
- Increase the proportion of the collisions that have enough energy to overcome the activation energy of the reaction. This can either be done by making the particles collide with each other with more energy, or by weakening the bonds in the molecules.

Rate = collision frequency x fraction of collisions that are successful

In practice, scientists can alter the rate of a chemical reaction by changing one or more of the following reaction conditions: temperature, pressure, concentration, surface area, mixing rate, use of a catalyst.

Student Tasks:

1. Complete the following table by placing ticks/crosses in the appropriate boxes.

Factor	Increases Frequency of Collisions	Increases Energy of the Collisions	Reduces Activation Energy	Increases Fraction of Collisions that are Successful
Increasing the Concentration (Liquids)				
Increasing the Surface Area (Solids)				
Increasing the Pressure (Gases)				
Adding a Catalyst or Enzyme				
Increasing the Temperature				

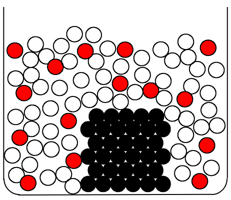
Rates Summary

Factor	Increases Collision Frequency	Increases Collision Energy	Reduces Activation Energy	Increases Fraction of Successful Collisions
Increasing Concentration (Liquids)	✓			
Increasing Surface Area (Solids)	✓			
Increasing Pressure (Gases)	✓			
Adding a Catalyst/Enzyme	?		✓	✓
Increasing the Temperature	✓	✓		✓

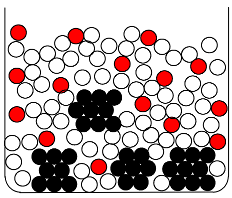
Rate = collision frequency x fraction of collisions that are successful

Student Tasks:

1. In the following examples, the reaction conditions have been changed e.g. the surface area of a solid reactant has been increased. You have to describe how this change in the reaction conditions will alter the reaction rate. In your answer, you should describe and explain the effect of the change on: the collision frequency; the collision energy and the proportion of collisions that will be successful.

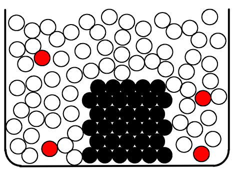


Large Particle Size -- Low Surface Area

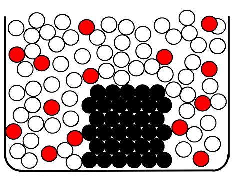


Small Particle Size -- High Surface Area

Particle Size

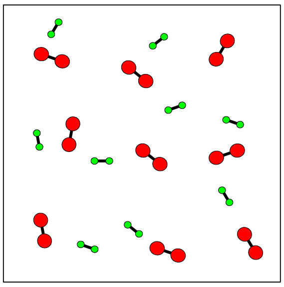


dilute acid

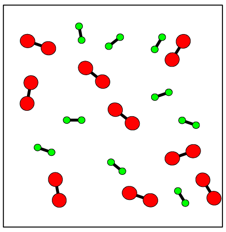


concentrated acid

Concentration



low pressure



high pressure

Pressure

Catalyst