8.3 Metals

Assumed Knowledge
It is assumed that students can:
- apply an appropriate atomic model (Rutherford model) to describe atomic structure
- determine the number of protons and neutrons in the nucleus of a neutral or charged atom
- determine the number of electrons orbiting the nucleus of a neutral or charged atom
- use the periodic table to describe some simple relationships between elements (maybe restricted to a very basic overview of the periodic table)
- identify where metals, non-metals and semi-metals are to be found in the periodic table
- identify where the most reactive metals and non-metals are to be found in the periodic table. The relatively unreactive nature of the group 18 (the noble gases) elements would also be assumed to have been brought to the attention of the students.

Students are assumed to know a series of simple reactions and to be able to qualitatively describe the reactants and products of the following reactions:
- corrosion
  (corrosion is the reaction of metals with atmospheric oxygen in the presence of water. Some metals corrode more quickly than others)
- acids on metals and carbonates
  (active metal plus acid gives solution of a salt plus hydrogen gas)
  (carbonate plus acid gives solution of a salt plus water plus carbon dioxide)
- neutralization reactions
  (metal hydroxide plus acid gives solution of a salt plus water)
  (metal oxide plus acid gives solution of a salt plus water)

1. ***Metals have been extracted and used for many thousands of years.***
Historical Perspective of Metals

"Outline and trace some uses of different metals through history, including contemporary uses, as uncombined metals or as alloys".
"Describe the use of common alloys, including steel, brass and solder and explain how these relate to their properties".
"Describe an ancient metal extraction process and recount the separation and or decomposition processes involved".

All these topics should be studied at the same time as they are so obviously related. Useful to consider the metals as they were discovered, and emphasise that as methods of extraction have improved and become more complex the range of materials that have found commercial use has increased.
Interesting story is that Napoleon the Third invested an enormous amount of State money in the development of aluminium (initially for use in the military) and after two years effort had produced a little more than 2 kg of metal. The investment was written off and the aluminium that was produced was used to produce the first set of aluminium cutlery. Visiting dignitaries used the silver service, visiting heads of states used the gold service and visiting monarchs used the aluminium service (such was the value of aluminium at that time).

Metal	**Remarks**	**Method of Extraction**
Gold, Silver and alloys	Pre-historic	Native metals
Copper, Tin and alloys	Distinctive ores, relatively common	Low temperature reduction with charcoal
Iron	Common ores	High temperature reduction with charcoal
Lead, Mercury	Distinctive ores, not very common	Heat, decompose
Steel	Requires metals such as vanadium, chromium which have relatively rare ores and difficult to process.	Requires electricity
Aluminium	Common ore, extremely difficult to process	Requires electricity

Emphasise that the **least reactive metals were the easiest to isolate and were the first metals** to be widely used, and that it is only in modern times that industry has become sufficiently skilled to isolate and refine metals such as aluminium, and the metals that are added to iron to make steel.

"Explain why energy input is required to extract a metal from its ore".
Many ores contain metal oxides. Metal oxides are more stable than metals. Energy is required to break the metal oxygen bond in these metal oxides to release the metal in its elemental form.

2. *Metals differ in their reactivity with other chemicals and this influences their uses.*

Activity Series of Metals

"Recall qualitative descriptions of reactants and products in corrosion, acids on metals and carbonates and neutralization".

"Describe observable changes when metals react with dilute acids, water and oxygen".

Both these points emphasise the need to reinforce material from previous years. Students will describe, presumably in greater detail, the reactants and products and observations of the following reactions:

(i) corrosion
(corrosion is the reaction of metals with atmospheric oxygen in the presence of water; some metals corrode more quickly than others)

(ii) acids on metals and carbonates
(active metal plus acid gives solution of a salt plus hydrogen gas)
(carbonate plus acid gives solution of a salt plus water plus carbon dioxide)

(iii) neutralization reactions
(metal hydroxide plus acid gives solution of a salt plus water)
(metal oxide plus acid gives solution of a salt plus water)

"Describe and justify the criteria used to place metals into an order of activity based on their ease of reaction with oxygen, water and dilute acids".

"Perform a first hand investigation and/or process information from secondary sources to determine the metal activity series".

I would recommend that these two points be considered together. A number of relatively simple experiments can be used to illustrate the criteria that metals can be arranged into an order of activity. Indeed a series of work-stations could be developed so that students performed a series of experimental tasks, and developed their own activity series, and they could then discuss why some elements seem so completely out of place (aluminium never appears as reactive as it is).

"Identify the dissolving of metals in acids as an ionization requiring the loss of electrons".

"Outline the relationship between the relative activities of metals and their position in the periodic table".

"Identify the importance of ionization energy in determining the relative reactivity of metals"

Interesting concept that a process only involves oxidation. The statement has omitted the accompanying reduction process. Can use the above experiment to demonstrate the oxidation of metals and to discuss the relative ease of oxidation from task three.

The experimental importance of ionization energies (as determined from the above qualitative results) can be turned around to emphasise the importance of using ionization energies to predict and to explain chemical phenomena. Graphical representations of ionization energy can be used to represent the periodic trends in ionization energy. Emphasise that ionization trends across the period can be related to the effective nuclear charge; whilst trends in ionization energy down a group of representative elements can be related to the size of the atom.

3. *As metals and other elements were discovered, scientists recognized that patterns in their physical and chemical properties could be used to organize the elements into a Periodic Table.*

Metals and the Periodic Table.

"Outline the history of the development of the Periodic Table including its origins, the original data used to construct it and the predictions made after its construction".

Antoine Lavoisier

Antoine Lavoisier was the first modern chemist to seriously challenge the notion of the Ptolemaic classification of chemistry. Ptolemy believed that there were only four elements: Earth, Wind, Fire and Water. Lavoisier believed that an element was a chemical substance that could not be reduced to simpler chemicals by either physical or chemical means. Lavoisier tabulated the chemical substances that resisted further simplification into a table that arranged these elements into four main groups.

Group 1 (gases)	Group 2 (non–metals)		Group 3 (metals)		Group 4 (earths)
Light	Sulfur	Antimony	Tin	Gold	Lime
Heat	Phosphorous	Silver	Iron	Platina	Barytes
Oxygen	Carbon	Arsenic	Manganese	Lead	Magnesia
Nitrogen	Muriatic radical (Cl^-)	Bismuth	Mercury	Tungsten	Argill
Hydrogen	Fluoric radical (F^-)	Cobalt	Molybdena	Zinc	Silex
	Boracic radical ($B_4O_7^{2-}$)	Copper	Nickel		

Lavoisier had doubts about the earths as elements, but they had not been decomposed and, as he said "......at the most this is mere conjecture [that the earths are compounds]. I hope that the reader will take care not to confound what I relate as truths of facts with what is yet only hypothetical".

Ten of the substances listed are now known not to be elements. The Earths have been shown to be oxides of alkali and alkaline earth elements. The acid radicals and the earths have also been shown to be compounds rather than elements.

Light and heat are basic forms of energy rather than elements.
Within a short period of time a great many more elements were discovered.
http://antoine.frostburg.edu/chem/senese/101/history/faq/antoines-elements.shtml
http://MAPLE.lemoyne.edu/~giunta/papers.html
http://MAPLE.lemoyne.edu/~giunta/lavtable.html

Johann Döbereiner
Another attempt to systematically organize the elements was published by Johann Döbereiner. Döbereiner discovered several groups, each containing three elements, with similar chemical properties. Each group was called a triad. The mass of the middle member of each triad is very close to the average of the remaining members of the triad. Consider the atomic weights of the following four triads given below:

Li	7	Ca	40	S	32	Cl	34.5
Na	23	Sr	88	Se	79	Br	??
K	39	Ba	137	Te	128	I	127

Döbereiner later predicted the atomic weight of bromine based on the known atomic weights of chlorine and iodine.

Example: *Estimate the atomic weight of bromine. Discuss the chemistry of bromine, paying particular attention to the similarities of the chemistry of bromine with respect to the chemistry of the chlorine and iodine (the other members of the triad).*
Answer: *The estimated atomic weight of bromine is the average of the atomic weights of iodine and chlorine.*

$$\text{Atomic weight (bromine)} = \frac{\Sigma \text{ atomic weights}}{\text{no of atoms}} = \frac{34.5 + 127}{2} = 80.8.$$

This calculated atomic weight (80.8) is not a bad estimate of the experimentally determined atomic weight (79.9).
The chemistry of bromine will be similar to the chemistry of chlorine and iodine. An enormous number of possible chemical reactions could be discussed. For example: all three elements form –1 anions. The binary hydrogen compounds of all three (i.e. HCl, HBr, HI) are acidic.
Döbereiner's system was inadequate because a great many elements did not seem to belong to any triad, and the chemistry of many elements was incompletely known, or even worse, was incorrectly known due to the limitations of chemical technique of that time.

John Newlands
John Newlands discovered a relationship that he called "the Law of Octaves". Newlands realized that when the elements were arranged in order of increasing atomic weight the eighth element had similar chemical properties to the first element, the ninth element had properties similar to the second element, and so on. Newlands also realized that Döbereiner's "triads" were largely preserved in this table, however these triads were found to form parts of larger families of elements.
When Newlands presented his ideas at the Royal Chemical Society he was openly ridiculed and Newlands' table of the elements was ignored by his contemporaries for more than twenty years.

Mendeleev and Meyer also developed a table of the elements based on increasing atomic weight. They, like Newlands, also noted periodic relationships in chemical properties.
Mendeleev
Mendeleev believed that any table of the elements should be arranged in a manner that highlighted similarities within groups of related elements. He came to call this the "Law of Periodicity". Mendeleev used the experimentally determined atomic weights of the elements to initially position the elements in his periodic table. Mendeleev realized that using the atomic weights of the elements seemed to place some elements in the wrong place in the periodic table, or place elements of unrelated chemistry in the same chemical group. Thus he

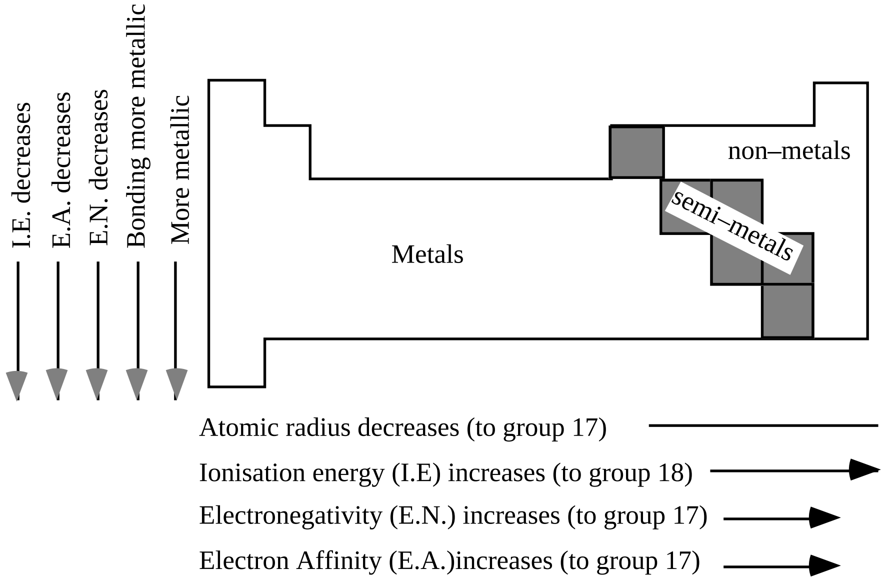
(i) left spaces in his periodic table for undiscovered elements.
(ii) only used the atomic weight of the elements as a guide to position the elements. The chemistry of the element would determine the exact position of the element in the periodic table. Thus iodine (atomic weight 126.9) was placed after tellurium (atomic weight 127.6) because tellurium had similar chemistry to selenium and other elements in that group whilst iodine had similar chemistry to bromine and other elements in group 17.
(iii) incorporated rows (periods) of varying length in the table.
Mendeleev's greatest achievement was to predict the properties of some of the undiscovered elements. Eka–silicon was the name given to germanium by Mendeleev.

	Properties of	Properties of	Predicted and Observed Properties of Germanium	
Property	Silicon	Tin	Predicted Value	Observed Value
Atomic weight	28	119	73.5	72.6
Specific gravity	2.33	7.30	4.8	5.32
Atomic radius	110	150	130	123
Specific gravity of oxide	2.6	7.0	4.80	4.70
Specific gravity of chloride	1.5	2.2	1.85	1.90
Boiling point of chloride	57°C	113°C	85°C	86°C

The value of this early periodic table was that it predicted, with reasonable accuracy, the properties of the unknown elements because their chemistry could be predicted by the position of the unfilled spaces in the periodic table.

Modern Periodic Table
There are two obvious differences between the modern periodic table and Mendeleev's periodic table:
– in the modern periodic table the elements are arranged in order of increasing atomic number (and not atomic weight as proposed by Mendeleev).
– the modern periodic table includes an extra column (the noble or inert gases) of elements that had not been discovered. Interestingly this means that Newlands octaves are more correctly nonets (since the 1st and 9th element are chemically similar).

The modern periodic table preserves the idea of columns (called groups) and rows (called periods). There is a gradual variation in the properties of the elements across a period. The groups represent families of chemically similar elements. The elements can also be classified as belonging to one of three major groups of elements in the periodic table, metals, non–metals and semi–metals (also called metalloids).



The underlying reasons for the observed trends can be related to basic atomic structure.
http://www.chemicalelements.com/
http://www.uky.edu/~holler/periodic/periodic.html

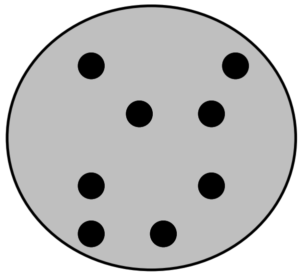
Tom Lehrer's song of the Periodic Table
http://www.sd01.k12.id.us/
select schools
select Borah High Area
select Borah High
then add to URL *teachers/purdy/links/elements/elements.htm*

"Recall descriptions of some relationships between the elements using the Periodic Table".
"Recall an appropriate model that has been developed to describe atomic structure".
"Explain the relationship between the position of elements in the Periodic Table and
- **electrical conductivity**
- **hardness**
- **ionization energy**
- **atomic radius**
- **melting point**
- **boiling point**
- **combining power**
- **electronegativity**

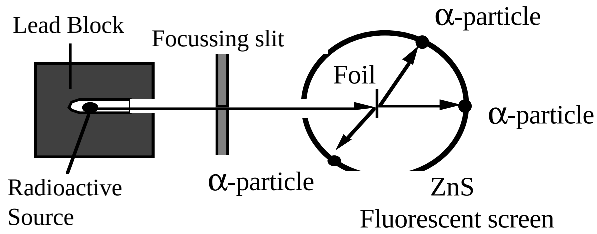
It is assumed knowledge that students have been introduced to "an appropriate atomic model" in the previous years of their higher school education, although it does not state what that appropriate model might be. Presumably the appropriate model is the Rutherford model of the atom.
It is impossible to explain the relationship between the position of the elements in terms of hardness. Note that 'combining power' and 'electronegativity' are introduced for the first time.
'Combining power' should not be confused with the charge of the mono-atomic ions. An understanding of the electronic configuration would certainly help to work out he combining capacities of the elements.

Thomson Plum–Pudding Model of the Atom

Thomson proposed that an atom was comprised of electrons embedded in positively charged material. This came to be seen as electrons, like plums, embedded in a positively charged dough. Different types of atoms had different amounts of positively charged dough, and thus different numbers of electrons.



Geiger–Marsden Experiment

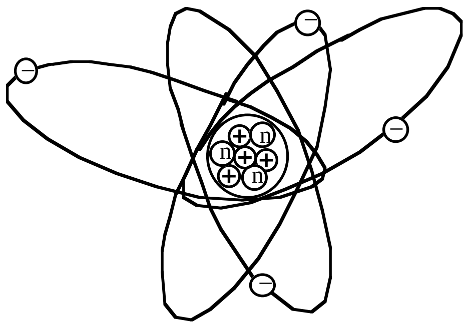


Geiger and Marsden were two young physicists who worked in the laboratory of Ernest Rutherford. They used a radioactive source to generate α-particles that were then fired at a thin sheet of metal (most notably a sheet of gold foil). The results were inconsistent with Thomson's model of the atom. If Thomson's model of the atom were correct, then all of the α–particles would pass straight through without any deflection. However, whilst most of the α-particles did pass straight through the metal foil, some were deflected through large angles. Indeed, some α-particles rebounded from the metal foil. Rutherford commented that this was as surprising as a mortar shell rebounding from tissue paper. This result could not be accounted for by the atomic model proposed by Thomson.

The Nuclear Model of the Atom

Rutherford realized that the atomic model proposed by Thomson was inconsistent with the results of the Geiger–Marsden experiment. The atomic model proposed by Rutherford included:

- The centre of the atom is the nucleus.
- The nucleus contains most of the mass and all of the positive charge of the atom.



Electrons orbit the nucleus at great distance from the nucleus (therefore an atom is mostly empty space).
Rutherford later proposed that some atomic nuclei might contain neutral particles. However the neutron was not discovered until 1932 by Chadwick.

Problems with the Rutherford "Nuclear" Model of the Atom

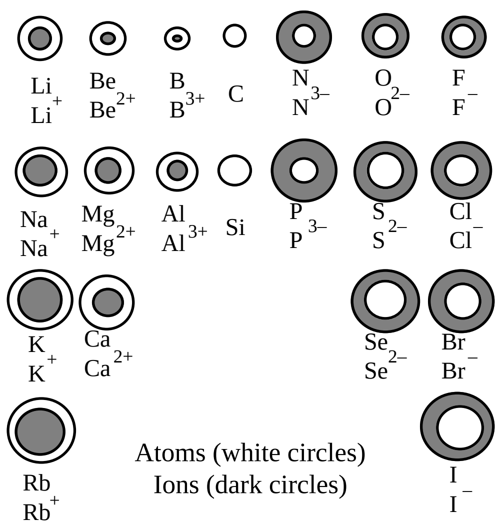
There were two main problems with the Nuclear Atomic model proposed by Rutherford:

– the electron should thus spiral in to the nucleus, and the Rutherford atomic model should decay to give the Thomson atomic model,
– atomic spectra could not be explained by the Nuclear Atomic model.

Relative Atomic Sizes

Atomic radius is a measure of the distance between an atomic nucleus and the outermost electron (the valence electron). It is extremely difficult to measure this distance, and thus a variety of measures of atomic size exist. The relative sizes of some atoms and ions are shown at the right.
The trends can be summarized:

☆ cations are smaller than their parent atoms,
☆ anions are larger than their parent atoms,
☆ within each period (horizontal row), the atomic size tends to decrease with increasing atomic number (i.e. increasing nuclear charge).
☆ Within each group (vertical column), the atomic size increases with increasing atomic number.



These general trends can be explained if we look at the two factors that primarily determine the atomic size. In going down a group the size of the atom increases as the inner shells are filled and the valence electron is forced into an atomic orbital further from the nucleus. The other is the effective nuclear charge. The electrostatic force of attraction between the valence (outermost) electron and the nucleus is lessened by any electrons in inner closed shells. This shielding by inner electrons means that the "effective nuclear charge" is always less than the (actual) nuclear charge.
The effective nuclear charge of beryllium is larger than the effective nuclear charge of lithium and thus beryllium is smaller than lithium. The valence electron of lithium is in the 2s orbital whilst the valence electron of hydrogen is much closer to the nucleus in the first orbital. Thus lithium is much larger than hydrogen.

In summary, the size of atoms decreases across the period to a minimum in group 17 because of increasing effective nuclear charge. The atomic size increases down a group as increasing numbers of electron shells are filled.

Electronegativity
Electronegativity is a measure of the ability of an atom in a molecule to draw bonding electrons to itself. In general, electronegativity increases from left to right in the periodic table (with the maximum for each period occurring in group 17). In general the electronegativity decreases from the top to the bottom in each vertical group.
Again these trends in electronegativity can be related to atomic structure. The larger the effective nuclear charge the greater the electronegativity. The group 18 elements are excluded, and have very low electronegativities reflecting their stable, filled electron shells and their lack of chemical reactivity.
The electronegativity decreases down a group because the size of the atoms increases down a group. The large distance between the atomic nucleus and bonding electrons reduces the electrostatic force of attraction between the nucleus and the bonding electrons, (i.e. reduces the electronegativity).

First Ionization energy
The first ionization energy is the energy required to remove an electron from a neutral atom in the gas phase.
$E(g) \rightarrow E^+(g) + e^-$
i.e. $Na(g) \rightarrow Na^+(g) + e^-$ ionization energy = 502 kJ
i.e. $Cl(g) \rightarrow Cl^+(g) + e^-$ ionization energy = 1257 kJ
Note: Ionization energies always refer to the formation of cations.

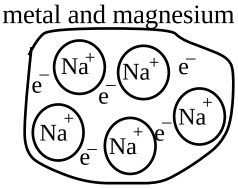
Consider the ionization energies listed below:
Successive Ionization Energies for the elements sodium to silicon

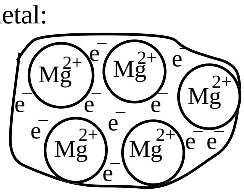
Element	Configuration for neutral atom	Ionization Energy ($kJ\ mol^{-1}$) 1st	2nd	3rd	4th	5th
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$	502	4569	6919	9550	13356
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$	744	1457	7739	10547	13636
Al	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 2p^1$	584	1823	2751	11584	14837
Si	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 2p^2$	793	1583	3238	4362	16098

It is evident from the table that, for each element, steep rises in ionization energies occur after the removal of the last valence electron. Thus the steep rise in ionization energy for sodium occurs after the first electron is removed (i.e. the second electron is nine times more difficult to remove than the first electron). The steep rise in ionization energy for magnesium occurs after the second electron is removed (i.e. the first electron requires 744 kJ of energy to remove. The second energy requires about twice as much energy to remove, the third electron requires about five times the energy that the second electron required. Note that the fourth electron required less than twice the energy that was required to remove the third electron). Thus magnesium has two electrons in the valence shell. These two electrons are relatively easily removed. Inner shell (or core) electrons are very difficult to ionize and thus remove from an atom.

In summary, the ionization energies depend on:
– the distance between the nucleus and valence electrons (the number of complete shells);
– the effective nuclear charge experienced by an electron in the valence shell;
– the details of the electronic structure.

Melting Points of Metals
Metallic bonding is the result of the sharing of all the valence (outer) shell electrons of the metal atoms and the mutual attraction of the metal cations to the shared valence shell electrons. This is sometimes described as "positive ions in a sea of electrons". The strength of a metal bond is the result of the electrostatic force of attraction between the metal cation and the shared valence electrons. Consider sodium metal and magnesium metal:
The melting point of sodium is less than the melting point of magnesium because the electrostatic force of attraction of the magnesium cation is stronger than the electrostatic force of attraction of a sodium ion to its sea of electrons.





In summary the melting points of metals generally increase as the effective nuclear charge of the metal increases and as the number of valence electrons increases.

4. *Energy is required to extract metals from their ores.*

Energy.

"Define the terms mineral and ore with reference to economic and non-economic deposits of natural resources".

"Describe the separation processes, chemical reactions and energy considerations involved in the extraction of copper from one of its ores".

"Describe the historical use of two common metals and their ease of refining and activity".

"Process information from secondary sources and use the available evidence to analyse

- **the chemistry of an extraction process in an Australian metal refinery**
- **the environmental impact of the extraction process"**

The last syllabus entry suggests students undertake an assignment to investigate an extraction technique in Australia. Earlier syllabus points have suggested that students focus on copper. The student's report should compare the extraction of copper to that of some other common metal. Perhaps the comparison of copper to either iron or aluminium is appropriate.

Assignment:

Describe the ancient process for the extraction of copper metal. The report should describe an extraction process, and compare it to the manufacture of a modern alloy. A flow chart should accompany each process; a one page summary of the chemistry involved and a detailed comparison of the extraction of copper in antiquity with modern processes.

"**Choose equipment and perform a first hand investigation to extract copper from copper carbonate (malachite or azurite).**

There are several methods of isolating copper from its carbonates. The easiest method is the acid leachate process. In industrial usage this means that dilute sulfuric acid is allowed to percolate through the ore body and the leachate is collected as it drains from the ore body. Scrap iron is then dissolved into the leachate displacing copper.

5. *For efficient resource use, industrial chemical reactions must use measured amounts of each reactant.*

Metals and the Mole Concept.

"Define the mole as the number of atoms in exactly 12 g of carbon 12 (Avogadro's Number)"

"Describe the contribution of Gay-Lussac to the understanding of gaseous reactions and relate this to an understanding of the mole concept."

"Recount Avogadro's Law and describe its importance in developing the mole concept".

"Define the mole as the number of atoms in exactly 12 g of the isotope of carbon with mass number 12 (or ^{12}C).'

Gay–Lussac's Theory and Avogadro's Law

Gay–Lussac studied the behaviour of gases, in particular the behaviour of gases involved in a chemical reaction at the same temperature and pressure. He first showed that the volume of hydrogen gas that reacts with oxygen gas to form water is in the ratio of 2 is to 1 (i.e. two litres of hydrogen gas react with one litre of oxygen gas or 10 litres of hydrogen react with 5 litres of oxygen). He then studied several other reactions and demonstrated that:

"The ratios of the volumes of gases involved in a chemical reaction are expressed as small, whole numbers."

It must be remembered that the formulae of elements in their normal state and of compounds were not known, hence any chemical reaction could only be described in words. Consider the following two examples that illustrate the simplicity of the law:

one litre of hydrogen gas reacts with one litre of chlorine gas to give two litres of hydrogen chloride gas.

two litres of hydrogen gas react with one litre of oxygen gas to give two litres of steam.

Dalton could not explain the results that were summarized in Gay–Lussac's Law, principally because Dalton believed that elements were mono–atomic.

Avogadro proposed a revolutionary idea that showed how Gay–Lussac's Law could be explained in terms of the Atomic Theory. Avogadro's Law states":

"Equal volumes of gases, at the same temperature and pressure, contain equal numbers of gas *molecules*."

Avogadro used this principle to demonstrate that Gay–Lussac's Law was in keeping with the Atomic theory proposed by Dalton. Consider the two word equations that were used to illustrate Gay–Lussac's law and the interpretation that Avogadro's Law yields:

Gay–Lussac's Law:

One litre of hydrogen gas reacts with one litre of chlorine gas to give two litres of hydrogen chloride gas.

Avogadro's Interpretation:

One molecule of hydrogen gas reacts with one molecule of chlorine gas to give two molecules of hydrogen chloride gas

Gay–Lussac's Law:
Two litres of hydrogen gas react with one litre of oxygen gas to give two litres of steam.
Avogadro's Interpretation:
Two molecules of hydrogen gas react with one molecule of oxygen gas to give two molecules of steam.

Avogadro introduced the concept of molecules and of poly-atomic elements. He believed that some elements such as hydrogen, oxygen and chlorine were diatomic.

The ideas introduced by Avogadro allowed the correct atomic weights to be determined.
The definition of the mole is:

A mole is the number of ^{12}C atoms in exactly 12 g of ^{12}C.

This definition can now be used with the ideas of Gay Lussac and Avogadro to develop modern equations and to determine relative atomic weights through experiment. Consider the following reactions of carbon with oxygen:

Carbon	plus	oxygen	gives	carbon monoxide
12.0 g		16.0 g		28.0 g
Gay Lussac		12.25 L		24.5 L
Avogadro		x mole		2x mole

One mole of carbon was used since 12.0 g of carbon was used. It follows logically, that if carbon monoxide contains only one atom of carbon, than one mole of carbon monoxide has been used. Thus, one mole of carbon monoxide must weigh 28 g. Since it is known that one mole of carbon weighs exactly 12 g, it follows that one mole of oxygen must weigh 16 g.
Since the volume of carbon monoxide is twice the volume of oxygen Avogadro's Law predicts that there must be twice as many molecules of carbon monoxide as there are molecules of oxygen. Oxygen must be diatomic since there are two molecules of carbon monoxide formed for every molecule of oxygen, and the molar mass of oxygen is therefore $2 \times 16 = 32$.
In a similar fashion, it is possible to determine the relative atomic weights of all the other elements by comparing the amount of each element that reacts with exactly 12 g of carbon 12. The values of atomic weights that are listed in the periodic table are thus derived from experiments that ultimately refer to carbon 12 as the standard.

"Perform a first hand investigation to measure and identify the mass ratios of a metal and a non-metal in a common compound and calculate its empirical formula"

Since copper is the show stopper in this part of the syllabus, it is most probably easiest to determine the empirical formula of copper oxide.

"Recall distinctions between elements, using information about the numbers of protons, neutrons and electrons".

Chemical reactions can be represented by chemical equations. Chemical equations show the formula and physical state of each reactant and product. The chemical equation shows the ratio of the number of moles of products and reactants.

For the general chemical equation

$$CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(\ell) + CO_2(g)$$

Stoichiometric Factor:

$$\text{mol } CaCO_3 = \frac{1}{2}\text{ mol } H^+$$

6. The relative abundance and ease of extraction of metals influences their value and breadth of use in the community.

Recycling

"Describe the relationship between the commercial prices of common metals, their actual abundances and relative costs of production."
"Explain why ores are non-renewable resources."
"Recount the steps taken to recycle a named metal such as aluminium."

www.comalco.com.au
www.yahoo.com.au
select Australia, type recycle