

**MODULE 111 – PRINCIPLES OF PHYSICAL AND INORGANIC CHEMISTRY**  
**UNIT 1 – MOLES, FORMULAE AND REACTIONS**

**COURSE OUTLINE 2018**

Moles, formulae and measurement: What is the meaning of relative atomic mass and how is it measured? What is a chemical formula and what is the difference between relative formula mass, relative molecular mass and molar mass? What is a mole and how is it related to mass, molar mass, aqueous volume, molar concentration (molarity), mass concentration and gaseous volume?

How should we record measurements? What is the difference between accuracy and precision? How can we calculate apparatus error and why is it important?

Chemical equations: What is a chemical equation? How do we balance equations? How can we use equations to predict the amount of substance needed or produced in a chemical reaction? What is a limiting reagent? What is the difference between yield and atom economy? What are the main types of chemical reaction and what are the differences between them? What are ionic equations? What is the pH scale? What is volumetric analysis? What is the meaning of chemical equivalency?

Principles of Scientific Enquiry: What is the difference between a theory, a theorem and a hypothesis? What is the difference between a rule and a law? What is the difference between deduction and induction? What is serendipity?

**CONTENTS**

<b>1.</b>	<i>atomic mass units, use of carbon-12, relative isotopic mass, relative atomic mass, principles of mass spectroscopy</i>
<b>2.</b>	<i>review of chemical structures: giant ionic, giant covalent, simple molecular, giant metallic structures; unit formula, molecular formula</i>
<b>3.</b>	<i>measuring amount of substance; the mole; molar mass; amount of substance in solution; ideal gas equation, empirical formula</i>
<b>4.</b>	<i>principles of scientific measurement; accuracy and precision, appropriate degrees of accuracy, units and their interconversion; % error</i>
<b>5.</b>	<i>chemical equations; balancing equations; amount of substance calculations</i>
<b>6.</b>	<i>limiting reagents; simple volumetric analysis; atom economy and percentage yield</i>
<b>7.</b>	<i>acid-base reactions, the pH scale</i>
<b>8.</b>	<i>oxidation and reduction, oxidation numbers, redox reactions, disproportionation, redox reactions in acidic and basic solution</i>
<b>9.</b>	<i>Redox titrations; manganate, dichromate and iodine-thiosulphate titrations</i>
<b>10.</b>	<i>Chemical equivalency of elements; concept of normality; principles of scientific enquiry; induction, deduction, hypothesis, theory, serendipity</i>

*items in italics are covered at senior secondary level*

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## Lesson 1

### 1. Atomic Structure

- The basic properties of these three particles can be summarized in the following table:

Particle	Relative Charge	Relative Mass
<b>Proton</b>	+1 unit	Approx 1 unit
<b>Neutron</b>	No charge	Approx 1 unit
<b>Electron</b>	-1 unit	Approx 1/1840 units (very small)

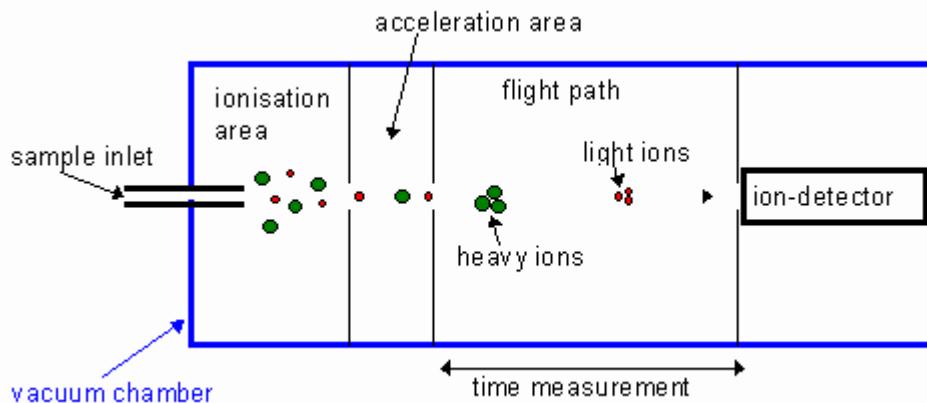
- All atoms have an **atomic number** (number of protons) and a **mass number** (number of nucleons); the chemical symbol and main identity of the atom is based on its atomic number; atoms with the same atomic number but different mass numbers are **isotopes**; they have the same chemical symbol; if a specific isotope is being referred to it can be identified by its mass number as a superscript prefix to the chemical symbol or separated by a hyphen after the name (eg  $^{210}\text{Po}$  or polonium-210)
- One unit of charge is  $1.602 \times 10^{-19}$  coulombs; protons have a charge of +1 unit and electrons have a charge of -1 unit; all charges are measured in these units; the positive charge on a proton is exactly equal to the negative charge on an electron
- Protons and neutrons (nucleons) have a similar, but not identical, mass; furthermore, the mass of nucleons in one atom can vary from one atom to another (because of different binding energies); the mass of a nucleon is therefore not constant; the masses of protons and neutrons, and hence the masses of different atoms, are measured in **atomic mass units** (amu); 1 amu is  $1.661 \times 10^{-27}$  kg; this is  $1/12^{\text{th}}$  of the mass of an atom of carbon-12, and is also therefore the average mass of a nucleon in carbon-12; a nucleus of carbon-12 therefore has a mass of 12.00000 atomic mass units by definition; the mass of any individual atom, or isotope, can be measured on this scale (ie in atomic mass units); the ratio of the mass of an atom to  $1/12^{\text{th}}$  of the mass of an atom of carbon-12 is called the **relative isotopic mass** of an atom
- carbon-12 is chosen because its mass per nucleon neither unusually high nor unusually low, which means the all relative isotopic masses are usually very close to the value of the mass number; in most chemical calculations, the relative isotopic mass is taken to be equal to the mass number:

Isotope	Mass number	Relative isotopic mass
$^1\text{H}$	1	1.007825
$^4\text{He}$	4	4.002603
$^9\text{Be}$	9	9.012182
$^{27}\text{Al}$	27	26.981538
$^{59}\text{Co}$	59	58.933200

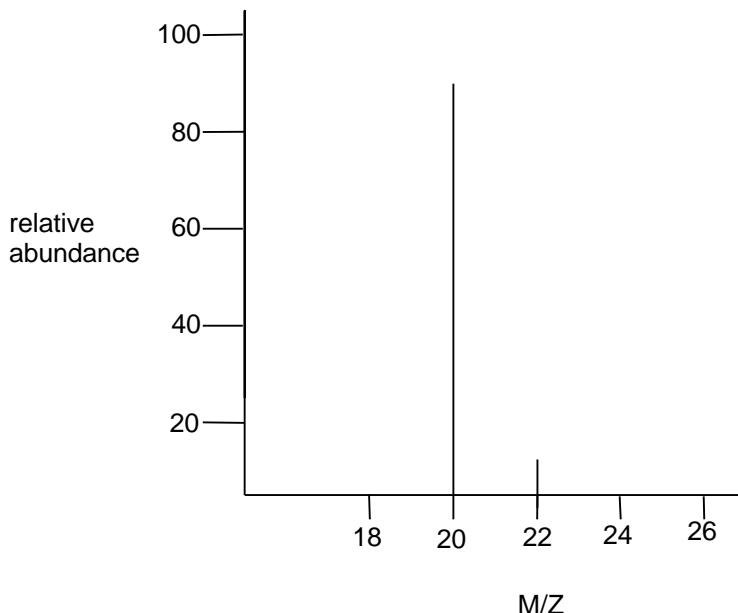
- The **relative atomic mass** of an atom is the ratio of the average mass of an atom to  $1/12^{\text{th}}$  of the mass of one atom of carbon-12; it is the weighted average mass of the different isotopes of the atom; it can be found to 1 decimal place in the Periodic Table

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- The time of flight (TOF) mass spectrometer is an instrument used for measuring accurately the masses of atoms and molecules; it can also be used to measure the relative abundance of different isotopes and to predict the structure of more complex molecules; a TOF mass spectrometer has the following structure:



- first a sample is vaporized
  - then it is ionized
  - the ions are then accelerated to a uniform KE by an electric field
  - the ions are then allowed to drift towards the detector; heavier ions move more slowly
  - an electric current is detected each time an ion hits the detector
  - the mass of the ion can be deduced from the time it takes to reach the detector, and its abundance can be detected from the size of the current produced
- The relative atomic mass of the element can be calculated from its mass spectrum. An example of the mass spectrum produced by Ne is shown below:



- The peak at 20 is  $^{20}\text{Ne}^+$ , and the peak at 22 is  $^{22}\text{Ne}^+$

## Lesson 2

### 2. Review of chemical structures and formulae

- Pure substances can be classified as either elements or compounds:
  - An element is a substance containing only one type of atom
  - A compound is a substance containing two or different atoms bonded together with a fixed composition
- Atoms can combine to form elements or compounds in a number of different ways:
  - Atoms can exchange electrons and become oppositely charged ions; the attraction between oppositely charged ions is an **ionic bond**
  - Atoms can share one or more pairs of electrons; each shared pair of electrons is attracted to the nuclei of both atoms and is a **covalent bond**
  - Atoms can give up electrons into a delocalised sea, forming cations in the process; the collective attraction between the cations and the delocalised electrons is a **metallic bond**
- These bond types can give rise to a number of different chemical structures:
  - (i) **Giant ionic lattice structures (compounds only)**
    - Oppositely charged ions tend to arrange themselves in ordered 3D lattices, each ion is surrounded by several others of opposite charge and the lattice is held together by ionic bonds; this is known as a **giant ionic lattice structure**; compounds with this structure are often called **ionic compounds**.
    - The ions in ionic compounds can be monatomic or polyatomic; the bonding within polyatomic ions is covalent but they still form ionic bonds and giant ionic lattice structures with other ions
    - Ionic compounds are represented by a **unit formula**, which is the ratio of each ion present in the lattice, with the charges omitted
    - The **relative formula mass** of an ionic compound is the sum of the relative atomic masses in the unit formula
  - (ii) **Giant metallic lattice structures (usually elements)**
    - Atoms lose electrons to form positive ions and arrange themselves in ordered 3D lattices, held together by the attraction to the delocalised electrons (metallic bonds)
    - Elements which have giant metallic lattice structures are called **metals**
    - The formula of a metal is the symbol of the atom it is made from

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**(iii) Simple molecular or simple atomic structures (elements and compounds)**

- When atoms form covalent bonds with other atoms, it usually takes only a few bonds to form before all atoms reach their bonding capacity; a small group of atoms held together by covalent bonds is called a molecule
- Different molecules are held together by intermolecular forces (Van der Waal's forces with or without hydrogen bonding), but these forces are much weaker than covalent bonds
- The resulting structure consists of discrete molecules held together by intermolecular forces; this is known as a **simple molecular structure** and is common in both elements and compounds
- Molecules are represented by a **molecular formula**, which is the total number of atoms of each element in one molecule of that substance
- The **relative molecular mass** of a molecule is the sum of the relative atomic masses in the molecular formula
- In noble gases, which form no bonds with each other at all, the individual atoms are held together by weak Van der Waal's forces; these are known as **simple atomic structures**; their formula is the symbol of the individual atom

**(iv) Giant covalent lattice structures (elements and compounds)**

- In some cases, when atoms form covalent bonds, it is not possible to satisfy the bonding requirements of each atom by the formation of a small molecule; in such cases the network of covalent bonds stretches over a large number of atoms in two or three dimensions and a giant lattice is formed; this is known as a **giant covalent lattice structure** and is found in both elements and compounds
  - The giant lattice can be a 2D lattice or a 3D lattice
  - The formula of elements with giant covalent structures is the symbol of the atom is made from; compounds with giant covalent structures are represented by a unit formula, which is the ratio of the number of each type of atom in the structure
  - The **relative formula mass** of a giant covalent structure is the sum of the relative atomic masses in the unit formula
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- The physical properties of elements and compounds depend on their chemical structure
  - Analysis of any compound can give the composition of each element by mass; this can in turn can give the **empirical formula** of a compound; this is the simplest whole number ratio of atoms of each element in a compound; this is not the same as the molecular formula (of molecules) or the unit formula (for ionic compounds) and further analysis is required to deduce the molecular formula of molecules and the unit formula of ionic compounds
  - In some cases, the pure substance is not the most convenient form in which to use it; instead, substances are dissolved in a solvent to make a **solution**; this is common with molecules and ionic compounds; in most cases the solvent is water and the solution is called an **aqueous solution**

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## Lesson 3

### 3. The mole

- The mole is the SI unit for **amount of substance (n)**; one mole is equivalent to  $6.02 \times 10^{23}$  particles (N) this is known as **Avogadro's number (L)**; amount of substance has units of **mol**
- Avogadro's number is defined as the number of atoms in 12.000 g of carbon-12; this means that one mole of any atom must have a mass equal to its relative atomic mass in grams; this is known as the **molar mass (m<sub>r</sub>)** of an atom and has units of **gmol<sup>-1</sup>**
- The molar mass of any element or compound (m<sub>r</sub>) is its relative formula mass, relative atomic mass or relative molecular mass in grams
- The amount of substance, in moles, can be determined directly by measurement in three ways:
  - of any pure substance, from a **mass (m)** measurement using a mass balance:  $n = \frac{m}{m_r}$ ; because molar mass has units of  $\text{gmol}^{-1}$ , the mass should be measured in grams (g); the SI unit of mass is the kilogram (kg)
  - of any solution of known **molarity (C)**, from a **volume (V)** measurement using a measuring cylinder, pipette, burette or volumetric flask:  $n = CV$ ; because molarity has units of  $\text{mol dm}^{-3}$ , the volume should be measured in  $\text{dm}^3$ ; the SI unit of volume is  $\text{m}^3$  and most instruments measure volume in  $\text{cm}^3$
  - of any gas, from a volume measurement using a gas syringe or inverted measuring cylinder, in addition to **temperature (T)** and **pressure (P)** measurements using the ideal gas equation:  $n = \frac{PV}{RT}$ ; in most cases the pressure can be taken to be atmospheric pressure ( $1.01 \times 10^5 \text{ Pa}$ ); SI units must be used for the ideal gas equation, so volume should be in  $\text{m}^3$ , pressure in Pa and temperature in K

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## Lesson 4

### 4. Accuracy, Precision and measurement error

- Almost all quantities used in Chemistry are measurements, and are therefore subject to measurement error:
  - Graduated instruments such as measuring cylinders, gas syringes, burettes, mass balances and thermometers typically have an error equal to half of the smallest graduation
  - If instruments are used twice to obtain a single value (mass balances, thermometers, burettes) then the error per reading is doubled
  - Fixed measurement devices such as pipettes and volumetric flasks are individually labelled with the measurement error
  - Burettes have an additional error of  $0.05 \text{ cm}^3$  per reading due to the individual drop on the tip
  - These errors can be expressed as percentage errors by dividing by the reading itself and expressing as a percentage
  - The total percentage error in an experiment can be calculated by summing the individual percentage errors from each instrument in the experiment
- Apparatus errors are a limiting factor in the **precision** and **accuracy** with which results can be obtained; measurements should reflect this limitation in the way they are recorded; the number of significant figures used in an answer is an indication of the confidence in the **accuracy** of the result; results must always be given to the number of significant figures than the apparatus error can justify; no more and no less
- The final answer resulting from calculations which use measurements should be given to the same number of significant figures as the least precise measurement
- Accuracy is a measure of the closeness of the result to the correct result; precision is a measure of the closeness of the results to each other; the number of significant figures used in an answer is an expression of the limits on precision, and hence the confidence in accuracy, with which the results have been obtained

## **Lesson 5**

### **5. Equations and Chemical Reactions**

- Chemical equations show the formulae (unit or molecular) of the reactants and products in a chemical reaction; due to the law of conservation of mass, the total numbers of each atom must be the same on both sides of the equation
- Chemical equations also show the molar ratio in which the reactants react together and produce products; the ratios are written in front of each formula and are known as stoichiometric coefficients; the stoichiometric coefficients can be deduced by balancing the equation
- The stoichiometric coefficients can be used to deduce the number of moles (and hence any mole-dependent quantities) of any reactant or product involved in the reaction once the number of moles of one reactant or product is known
- Unless the reactants are mixed together in the same mole ratio in which they react, one reactant will be used up before the others; this is known as the limiting reactant and the other reactants are said to be in excess
- The number of moles of product predicted by mole calculations should be regarded as a maximum; in practice, the amount of product obtained will be less than this; the reaction may not proceed to completion and other practical losses may occur as a result of the synthetic process; the amount of product obtained is called the yield; this can be expressed as the **percentage yield**, by expressing it as a percentage of the maximum product possible; the percentage yield will vary based on the conditions and the practical details of the synthesis
- In many reactions, only one of the products formed is useful; the other products are waste products; the sum of the molar masses of useful products in an equation can be expressed as a percentage of the sum of the molar masses of all of the products in an equation; this is known as the **percentage atom economy of the reaction**; atom economy is a property of the equation itself; it cannot be changed by changing the conditions and is unrelated to the percentage yield
- The vast majority of chemical reactions can be described as either acid-base, redox or precipitation reactions

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## Lesson 6

### 6. Acid-base reactions and pH

- Several definitions exist for acids and bases (Arrhenius, Bronsted-Lowry, Lewis)

#### (a) Arrhenius definition

- An Arrhenius acid is a species which gives  $\text{H}_3\text{O}^+$  ions (often simplified as  $\text{H}^+$  ions) in water; neither  $\text{H}^+$  nor  $\text{H}_3\text{O}^+$  ions exist in any form other than aqueous so the species must react with water to produce  $\text{H}_3\text{O}^+$  ions:  $\text{HA} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{A}^-$ ; this is often simplified to  $\text{HA} \rightarrow \text{H}^+ + \text{A}^-$ ; if the acid dissociates fully it is known as a strong acid ( $\text{H}_2\text{SO}_4$ ,  $\text{HCl}$  and  $\text{HNO}_3$ ); most other acids only dissociate slightly and are known as weak acids; some acids can produce more than one  $\text{H}^+$  ion (eg  $\text{H}_3\text{PO}_4$  and  $\text{H}_3\text{PO}_4$ )
- An Arrhenius base is a species which gives  $\text{OH}^-$  ions in water; some bases are solutions of ionic hydroxides (eg  $\text{NaOH}$ ); others dissociate in water to give  $\text{OH}^-$  ions (eg  $\text{NH}_3$ ,  $\text{CO}_3^{2-}$ ); the Arrhenius definition of a base is very narrow and Arrhenius bases are more commonly referred to as alkalis; bases which fully dissociate in water to give  $\text{OH}^-$  are called strong bases (only soluble ionic hydroxides); bases which only partially dissociate are called weak bases; some bases can react with more than one  $\text{H}^+$  ion
- a more useful definition of a base is a species which can react with  $\text{H}^+$  ions to form a salt; this includes all alkalis as well as insoluble ionic oxides, hydroxides and carbonates; a salt is a species formed by the replacement of  $\text{H}^+$  in an acid with a metal ion or ammonium ion; a neutralisation reaction is a reaction between an acid and a base to form a salt; acids and bases which can donate or accept more than one  $\text{H}^+$  ion can form more than one salt
- Neutralisation reactions are often used in quantitative analysis (especially titrations) to analyse acids and bases; the equivalence point in titrations is usually observed by using acid-base indicators although conductimetric titrations can also be used; neutralisation reactions are most conveniently written as ionic equations, in which spectator ions are omitted

#### (b) Bronsted-Lowry and Lewis definitions

- Bronsted-Lowry acid-base theory provides a general definition of acids and bases: acid = proton donor, base = proton acceptor; acid-base reaction = a reaction which involves the transfer of protons; Arrhenius acids and bases are just special cases of Bronsted-Lowry reactions in which water acts as the base and the acid respectively ( $\text{HA} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{A}^-$ ,  $\text{B} + \text{H}_2\text{O} \rightarrow \text{BH}^+ + \text{OH}^-$ )
- Lewis acid-base theory provides an even more general definition: acid = electron pair acceptor; base = electron pair donor; acid-base reaction = a reaction involving the transfer of electrons

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**(c) Acidic, alkaline and neutral solutions**

- In practice, both  $\text{H}^+$  and  $\text{OH}^-$  ions coexist in all aqueous solutions due to the auto-ionisation of water; aqueous solutions in which  $[\text{H}^+] > [\text{OH}^-]$  are called acidic solutions; aqueous solutions in which  $[\text{OH}^-] > [\text{H}^+]$  are called alkaline solutions; aqueous solutions in which  $[\text{H}^+] = [\text{OH}^-]$  are called neutral solutions
- The value of  $[\text{H}^+]$  is therefore taken as a measure of the acidity of a solution, often expressed as  $\text{pH} = -\log_{10}[\text{H}^+]$ ; in neutral solutions  $[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7}$  mol dm<sup>-3</sup> so the pH = 7; in acidic solutions  $[\text{H}^+]$  is higher than this so pH < 7; in alkaline solutions  $[\text{H}^+]$  is lower than this so pH > 7
- The expression  $\text{pH} = -\log_{10}[\text{H}^+]$  can be used to calculate the pH of any solution if its  $[\text{H}^+]$  is known and vice versa; in this way the pH of strong acids can be calculated as follows: if  $\text{H}_x\text{A} \rightarrow x\text{H}^+ + \text{A}^{x-}$  and the molarity of the acid  $\text{H}_x\text{A}$  is C, then  $[\text{H}^+] = xC$  so  $\text{pH} = -\log_{10}(xC)$ ; similarly, the molarity of a strong acid can be deduced from its pH by applying the inverse formula:  $C = \frac{10^{-\text{pH}}}{x}$

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## 7. Precipitation reactions

- A precipitation reaction is one in which two ions in aqueous solution combine to form an insoluble solid; they occur when two different soluble salts are mixed – this mixing creates two new combination of ions, and if either combination is insoluble then a precipitation reaction will take place
- Group I and ammonium cations, and nitrate anions, form no insoluble salts and are never involved in precipitation reactions; some other general rules for predicting precipitation are as follows:

Insoluble	Soluble
carbonates (other than with Group I and ammonium cations)	All Group I and ammonium salts
Hydroxides (other than with Group I, ammonium, strontium and barium ions)	All nitrates
Silver halides (except AgF)	AgF
BaSO <sub>4</sub> and SrSO <sub>4</sub> (and a few other sulphates)	Most other sulphates

- Precipitation reactions are best represented as ionic equations, with spectator ions omitted
- Precipitation reactions can be used in quantitative analysis, either in conductimetric titrations (because conductivity reaches a minimum) or in gravimetric analysis (the insoluble compound can be washed, dried and weighed)

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## Lesson 7

### 8. Redox reactions

- In inorganic chemistry, oxidation and reduction are usually defined in terms of electron transfer; oxidation is the loss of electrons - when a species loses electrons it is said to be oxidised; reduction is the gain of electrons - when a species gains electrons it is said to be reduced; processes which show the gain or loss of electrons by a species are known as half-equations or half-reactions
- The **oxidation number** of an atom is the charge that would exist on an individual atom if the bonding were completely ionic; in simple ions, the oxidation number of the atom is the charge on the ion; in molecules or compounds, the sum of the oxidation numbers on the atoms is zero; in polyatomic ions, the sum of the oxidation numbers on the atoms is equal to the overall charge on the ion; in all cases each individual atom is allocated a charge as if the bonding was completely ionic; in elements in their standard states, the oxidation number of each atom is zero.
- Many atoms can exist in a variety of oxidation states; the oxidation number of these atoms can be calculated by assuming that the oxidation number of the other atom is fixed:

Atom	Oxidation state in compounds or ions
Li, Na, K, Rb, Cs	+1
Be, Mg, Ca, Sr, Ba	+2
Al	+3
F	-1
H	+1 unless bonded to a metal, Si or B, in which case -1
O	-2 unless bonded to a Group I or Group II metal or H, in which case it can also exist as -1, or F, in which case it exists as +2

- Oxidation numbers are used when naming compounds according to the internationally agreed IUPAC rules:
  - Binary ionic compounds are named by stating the cation followed by the anion
  - Binary covalent compounds are named by stating the atom with a positive oxidation number followed by the atom with a negative oxidation number
  - Simple cations (and atoms in a positive oxidation state in binary covalent compounds) are named using the name of the atom followed by its oxidation number in brackets and Roman numerals: (+1 = I, +2 = II, +3 = III, +4 = IV, +5 = V, +6 = VI, +7 = VII)
  - Simple anions (and atoms in a negative oxidation state in binary covalent compounds) are named by changing the final one or two syllables of the atom to -ide
  - In anions containing more than one atom, one of the atoms has a positive oxidation number and the other has a negative oxidation number; these anions are named by changing the last one or two syllables of the atom with a positive oxidation number to -ate, and then adding the oxidation state of that atom in brackets and Roman numerals; the presence and number of atoms with a negative oxidation state is indicated as a prefix with the final one or two syllables of the atom changed to -o, preceded by the number of atoms if more than one (two = di, three = tri, four = tetra, five = penta, six = hexa)

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= hexa); in many cases the negative atom is oxygen and the prefix oxo- is usually omitted unless the number of oxygen atoms is unclear

- It is common to leave out the Roman numeral if the atom has only one known oxidation number (such as sodium or magnesium)
  - In binary compounds, adding prefixes such as mono, di, tri to denote the number of atoms is unnecessary when oxidation numbers are being used but is sometimes used as an alternative to using oxidation numbers
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- During oxidation and reduction, the oxidation numbers of atoms change; if an atom is oxidized, its oxidation number increases (ie it becomes more +ve or less -ve); if an atom is reduced, its oxidation number decreases (ie it becomes less +ve or more -ve)
  - Many oxidation and reduction processes involve polyatomic ions or molecules; the half-equations for these processes are more complex and are pH dependent, so can be written either using  $H^+$  (if the conditions are acidic) or  $OH^-$  (if the conditions are alkali); there are two ways to construct balanced half-equations:
    - Method 1:
      - Identify the atom being oxidised or reduced, and make sure there are the same number of that atom on both sides (by balancing)
      - Insert the number of electrons being gained or lost (on the left if reduction, on the right if oxidation) using the equation: No of electrons gained/lost = change in oxidation number x number of atoms changing oxidation number
      - balance O atoms by adding water
      - balance H atoms by adding  $H^+$
    - Method 2 (easier in more complex reactions):
      - Identify the atom being oxidised or reduced, and make sure there are the same number of that atom on both sides (by balancing)
      - balance O atoms by adding water
      - balance H atoms by adding  $H^+$
      - add the necessary number of electrons to ensure the charge on both sides is the same
- Both of the above methods give you equations for acidic conditions; to convert into alkaline conditions, add  $OH^-$  ions to both sides of the equation so that the number of  $OH^-$  and  $H^+$  ions on one side are equal, then convert each pair into a water molecule and cancel out water molecules until they only appear on one side
- 
- Half-equations consider gain and loss of electrons, but in fact electrons cannot be created or destroyed; they can only be transferred from species to species; gain of electrons by one species necessarily involves loss of electrons by another; oxidation and reduction thus always occur simultaneously; an oxidation is always accompanied by a reduction and vice versa; any reaction consisting of the oxidation of one species and the reduction of another is known as a **redox** reaction
  - A redox reaction can be derived by combining an oxidation half-equation with a reduction half-equation in such a way that the total number of electrons gained is equal to the total number of electrons lost

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- In redox reactions, the species which is reduced is accepting electrons from the other species and thus causing it to be oxidised; it is thus an **oxidising agent**; the species which is oxidised is donating electrons to another species and thus causing it to be reduced. It is thus a **reducing agent**; a redox reaction can thus be described as a transfer of electrons from a reducing agent to an oxidising agent
- There are many substances which readily undergo both oxidation and reduction, and which can therefore behave as both oxidising agents and reducing agents; species such as these are capable of undergoing oxidation and reduction simultaneously; the simultaneous oxidation and reduction of the same species is known as **disproportionation**; a disproportionation reaction is a type of redox reaction

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## Lesson 8

- Redox reactions can be used in quantitative analysis (especially titrations) to analyse oxidising and reducing agents:
  - (i) Reducing agents can be analysed by acidifying them and then titrating them against a standard solution of  $\text{KMnO}_4$ , which behaves as an oxidising agent as follows:  $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ ; the purple  $\text{MnO}_4^-$  is decolorised as it is added to the reducing agent until the reducing agent has been fully oxidised; any excess  $\text{MnO}_4^-$  will then turn the solution pink and this can be used to identify the equivalence point
  - (ii) Reducing agents can also be analysed by acidifying them and then titrating them against a standard solution of  $\text{K}_2\text{Cr}_2\text{O}_7$ , which behaves as an oxidising agent as follows:  $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$ ; the orange  $\text{Cr}_2\text{O}_7^{2-}$  is reduced to the green  $\text{Cr}^{3+}$  as it is added to the reducing agent in the presence of diphenylaminosulphonate indicator, which turns violet when excess  $\text{Cr}_2\text{O}_7^{2-}$  is present, which allows the equivalence point to be observed
  - (iii) Oxidising agents are generally analysed by reacting them with an excess of aqueous potassium iodide ( $\text{KI}$ ); the iodide ion is oxidised to iodine as follows:  $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$ ; the resulting iodine is then titrated against a standard solution of sodium thiosulphate ( $\text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{I}^-$ ) using starch indicator, which turns blue/black in the presence of excess iodine and hence disappears when the iodine has all been used up, which allows the equivalence point to be determined

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**Lesson 9**

**9. Normality and Equivalent Weights**

- It is sometimes helpful to count elements and compounds in terms of “equivalents” rather than moles, especially during acid-base or redox reactions; one equivalent weight is the mass of substance required to completely react with 1 mole of  $\text{H}^+$  or  $\text{OH}^-$  ions, or to gain or lose one mole of electrons
- For elements, the equivalent weight is the molar mass divided by the valency; for monoprotic acids this is equal to the molar mass; for diprotic acids such as  $\text{H}_2\text{SO}_4$  it is equal to 0.5 of the molar mass; for compounds and ions involved in redox reactions it is the molar mass divided by the number of electrons gained or lost
- in some cases the concentration of aqueous solutions is expressed in terms of normality (equivalent weights per  $\text{dm}^3$ ) rather than molarity ( $\text{mol dm}^{-3}$ ); for example 1 N  $\text{H}_2\text{SO}_4$  = 0.5 M  $\text{H}_2\text{SO}_4$

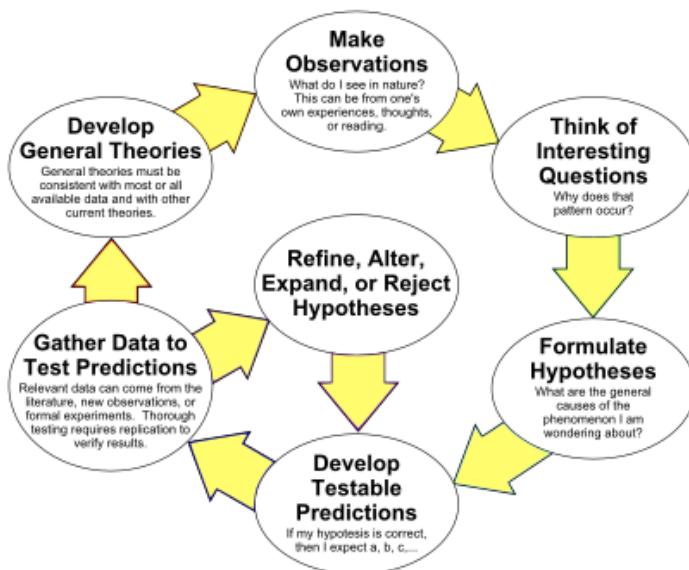
**10. Principles of Scientific Enquiry**

- Scientific enquiry is a continuously repeating (iterative) process of developing and testing theories through the collection of experimental data (observation and measurement)

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- The process of scientific enquiry can be summarised as follows:
  - observations are made about the natural world
  - attempts to explain these observations lead to the development of a hypothesis
  - the hypothesis leads to predictions which can be tested empirically (by observation)
  - based on the results of these tests, the original hypothesis may require refinement, alteration, expansion or even rejection
  - if a particular hypothesis becomes very well supported, a general theory or law may be developed, via a process of induction

### The Scientific Method as an Ongoing Process



[1]

- A Law is a descriptive principle of nature which holds in all circumstances covered by the wording of the law (it is experimental); a theory is a description of nature which encompasses more than one law
- A hypothesis is law or theory which is not sufficiently supported to be considered universally true; a hypothesis may become a theory or law if it is repeatedly and extensively tested and supported by observations
- Inductive reasoning is a method of reasoning in which the evidence provides strong evidence for the truth of a conclusion, meaning that the truth of the conclusion is probable but not certain; in science it is the process of converting a series of particular observations into a general law or theorem
- Deduction is the process of arriving at a conclusion by logic
- A theorem is a statement that has been proved on the basis of previously established statements (it is deductive)
- Most scientific breakthroughs result from serendipity: an unsought and unexpected, but fortunate, observation