Elements, Mixtures and Compounds

Definitions

Element:
A substance that is made up of only one type of atom. It cannot be chemically split into simpler substances. (on the atomic level)

Mixture:
A substance that contains two or more substances physically together but not reacted with one another chemically.

Compound:
A substance that contains two or more elements chemically combined together

Separation Techniques

Filtration
Purpose: To separate an insoluble solid from a liquid or a solution.

Crystallisation
Purpose: To obtain a pure solid from a saturated solution containing the solid.

Simple Distillation
Purpose: To separate solvent from a solution/miscible liquids that have significantly different boiling points.

(Note the water flow in the condenser: It goes against gravity)
Fractional Distillation
Purpose: To separate miscible liquids having different boiling points

Sublimation
Purpose: To separate a mixture of solids, of which one can sublime.

Separating Funnel
Purpose: To separate a mixture of immiscible liquids

Atomic Structure

Mass

Relative Atomic Mass is denoted by A_r.

Relative Molecular Mass, which is defined as the mass of 1 mole of molecules, is denoted by M_r

Relative Formula Mass, which is defined as the mass of 1 mole of ions, is denoted by M_r .

Isotopes

Definition: Atoms of the same element with different masses/same number of protons, but different number of neutrons.

Radioactive isotopes
- Gives out radiation that is invisible but harmful to life.
- Made artificially by Man.

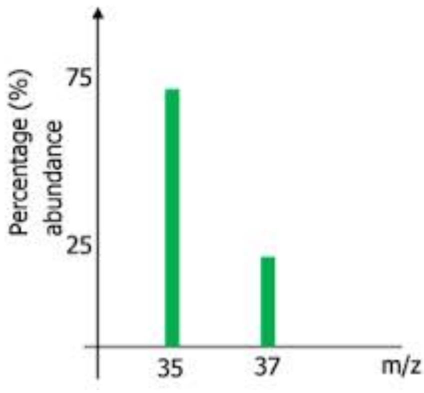
Mass Spectrometry

Uses
- To accurately determine the isotopes in an element sample
- To determine the relative atomic mass of an element and relative molecular masses of molecules.
- Molecular structure determination through studying of fragmentation patterns.

Procedure
- A sample is introduced and vapourised.
- By means of an electron gun, the sample atoms or molecules are converted into positive ions when electrons bombard them to knock of electrons. ($X_{(g)}$ + e^-(from gun) => $X^+_{(g)}$ +$2e^-$)

- The newly formed positive ions then accelerate through an electric field between negatively charged plates, into a chamber with a magnetic field.
- The magnetic field deflects the positive ions based on a particular mass/charge or m/e ratio.
- The deflected ions arrive at a detector.

Calculation of Relative Atomic Mass



- The percentage abundance of chlorine-35 is 3 times that of chlorine-37
- The **relative atomic mass of chlorine** can be calculated as 353 +3713+1=35.5

Ionisation energy

- Defined as the amount of energy required to remove one electron from a gaseous atom
- Ionisation energy increases with each successive removal of an electron from the valence shell
- There will be a modest increase in ionisation energy between electrons in different subshells, and a significant increase between electrons in different quantum shells.

Subshells and Electronic Configuration

Aufbau Principle

The electrons fill the subshell with the lowest energy level first.

Pauli Exclusion Principle
Each orbital can be filled by only two electrons, each with an opposite spin.

Hund's Rule of Multiplicity
If given a choice, electrons would prefer to be fill empty orbitals over partially-filled orbitals.

Chemical Bonding

Ionic Bonding

- Occurs between metallic elements and non-metallic elements
- Metal atoms (with 1,2,3 valence electrons) tend to lose electrons to form positively-charged ions.
- Non-metal atoms (with 5,6,7 valence electrons) tend to gain electrons to form negatively-charged ions. (thanks so much :) lol)

Covalent Bonding

- Occurs when two atoms share electrons from their outermost shell. Usually, each atom atom contributes one electron, so one covalent bond constitutes two electrons
- Shorter bond length usually means a stronger bond
- When orbitals overlap head-on, a sigma bond is formed
- When orbitals overlap side-on, a pi bond is formed

Coordinate Bond

- A covalent bond in which the shared pair of electrons is provided by only one of the bonded atoms.
- The symbol -> is used for a coordinate bond, an arrow pointing pointing from the donor towards the acceptor

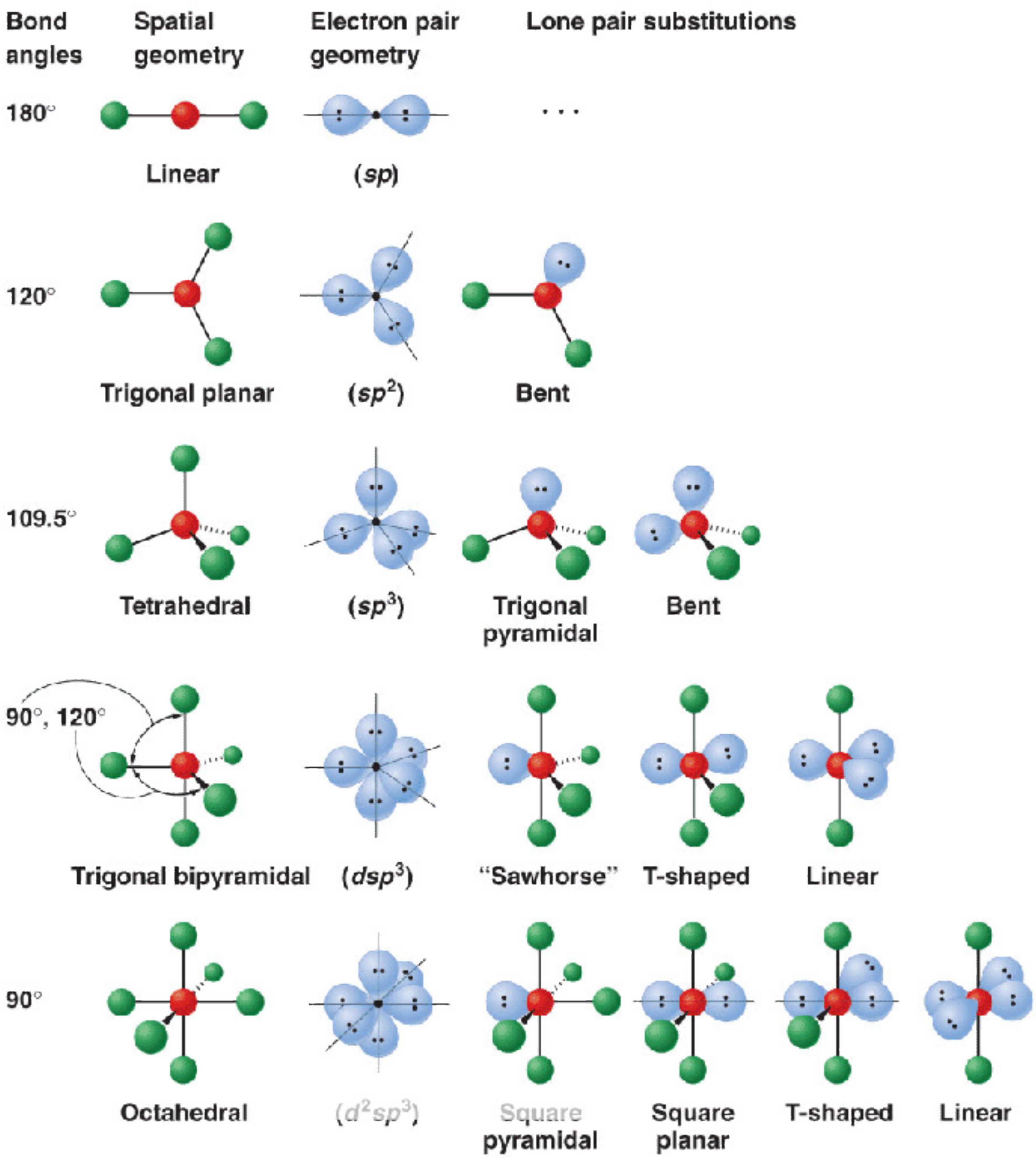
Metallic Bond

- Outermost electrons in each metal atom shared equally by all the metal atoms.
- Each atom's valence electron (one only) breaks free from the atom, making it a positively charged atom.
- This produces an array of of metal cations embedded in a "sea" of freely moving delocalised electrons.
- Metal atoms are held together by the electrical interaction between the positively charged ions and the negatively charged delocalised electrons.

- Metal atoms are arranged in layers that allow the atoms to slide over each other easily.

Valence Shell Electron Pair Repulsion Theory (VESPR)

- Pairs of electrons in the valence shell of a central atom in a molecule, that is bonding electron pairs and lone electron pairs, would mutually repel and thereby spatially arrange about the central atom so as to minimise repulsions
- Repulsion between lone electron pairs is stronger than bonding electron pairs



Intermolecular Forces

Van der Waals Forces

- Instantaneous dipole- Induced dipole
- Permanent dipole- Permanent dipole

Instantaneous dipole- Induced dipole

- Electrons in a molecule are in continual motion
- At any particular moment, the electron cloud around the molecule will not be perfectly symmetrical
- There will be more negative charge on one side of the molecule than on the other side
- Thus, weak forces of attraction exists between molecules (between the slightly negative charge area of one molecule and the slightly positive charge of one molecule)

Permanent dipole- Permanent dipole

- Interaction due to polar molecules
- Polar molecules have a slightly positive end and a slightly negative end
- This results in attraction between the positive and negative ends of molecules

Hydrogen Bonding

- Occurs when hydrogen is covalently bonded to oxygen or nitrogen
- The hydrogen atom becomes attracted to the oxygen or nitrogen atom of the neighbouring molecule.
- Plays an important role in the behaviour of water and DNA

Mole Concept

1 mole= 6 x 10^{23}

Empirical Formula

- Can be determined starting from percentage of mass present

Example: X comprises 80% Carbon and 20% Hydrogen, and has an M_r of 30

	C	H
Mass of Element in 100g of X	80g	20g
Number of Moles	80/12.0=6.67	20/1.0=20
Divide by smallest number of moles	6.67/6.67=1	20/6.67=3
Simplest Ratio	1	3

Therefore X has an empirical formula of CH_3

Redox Reactions

OIL RIG

Oxidation is loss of electrons, Reduction is gain of electrons

- Oxidation – increase of Oxidation States of ONE element
- Reduction – decrease of Oxidation States of ONE element

Oxidising and reducing agents

An oxidising agent is the agent that is being reduced in the reaction. It oxidises others (makes others lose electrons).

A reducing agent is the agent that is being oxidised in the reaction. It reduces others (makes others gain electrons).

An oxidising agent increases the oxidation state of another substance, but its own oxidation state is decreased.

A reducing agent decreases the oxidation state of another substance, but its own oxidation state is increased.

Oxidation state rules (four basic and seven specific)

- **Four Basic Rules**

1. Any atom in its elemental form is assigned an oxidation state of **zero**.
2. For monoatomic ions, the oxidation state of the element/atom equals the ion's charge.
3. For polyatomic ions, the sum of the constituent atoms' oxidation states is equal to the charge of the ion.
4. For neutral compounds, the sum of the constituent atoms' oxidation numbers is equal to zero.

4-part answering technique guide (RP)

1. Oxidation state of which element (for Oxidation state, always write sign before number)
2. ↑ or ↓ (increase or decrease)
3. from ? in where
4. to ? in where

<u>Common oxidising agents</u>

- Potassium Dichromate (VI)
- Potassium Permaganate(VII)

<u>Common reducing agents</u>

- Potassium Iodide
- Sulfur Dioxide

Qualitative Analysis

Solubility Rules

Nitrates
All nitrates are **soluble**

Chlorides/Halides
All halides are **soluble** with the exception of Lead and Silver.

Hydroxides/Oxides
All hydroxides/oxides are **not soluble** with the exception of group I hydroxides/oxides.

Sulfates
All sulfates are soluble with the exception of Barium, Lead and Calcium.