Name:

Periodic Table and Atomic Structure	**Objectives**
4. The Periodic Table	-describe the periodic table as a list of elements arranged so as to demonstrate trends in their physical and chemical properties -define the term element -associate the first 36 elements with their elemental symbols -distinguish between elements and compounds -state the principle resemblances of elements within each main group, in particular alkali metals, alkaline earth metals, halogens and noble gases -describe the reaction between water and lithium, sodium and potassium having seen the reaction demonstrated -describe by means of a chemical equation the reaction between water and lithium, sodium and potassium having seen the reaction demonstrated -outline the history of the idea of elements, including the contributions of the Greeks, Boyle, Davy and Moseley -outline the contributions of Mendeleev, **Dobereiner, Newlands and Moseley** to the structure of the modern periodic table -compare Mendeleev's with the modern periodic table -arrange elements in order of relative atomic mass and note differences with modern periodic table -define atomic number (Z) and mass number (A) -define relative atomic mass (A_r) using ^{12}C scale -define isotope -describe the composition of isotopes using hydrogen and carbon as an example -**calculate the relative atomic masses from abundance of isotopes of a given mass number** -describe the organisation of particles in atoms of elements numbers 1-20 -classify the first twenty elements in the periodic table on the basis of the number of outer electrons -list the numbers of electrons in each main energy level in atoms of numbers 1-20 -**build up the electronic structure of the first 36 elements** -**derive the electronic configurations of ions of s- and p-block elements only** -**describe the arrangement of electrons in individual orbitals of p-block atoms**

Elements:

1. The Greeks: 4 elements – earth, air, fire and water.
2. Robert Boyle: *Def^n:* An **element** is a substance that cannot be split into simpler substances by chemical means.

History of the Periodic Table:

1. Dobereiner: Put elements into groups of 3, called triads.
   *Def^n:*A **triad** is a group of three elements with similar chemical properties in which the atomic weight (relative atomic mass) of the middle element is approximately equal to the average of the other two.
2. Newlands: Put elements in order of increasing weight, and found that properties repeated themselves every eighth element.

Def^{n}: **Newland's Octaves** are arrangements of elements in which the first and the eighth element, counting from a particular element, have similar properties.

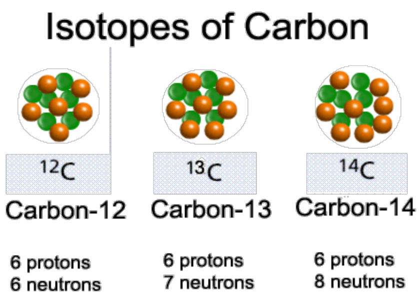
3. Mendeleev: Arranged the elements in order of increasing weight.
   Def^{n}: **Mendeleev's Periodic Law**: When elements are arranged in order of increasing atomic weight, the properties of the elements recur periodically, i.e. the properties displayed by the element are repeated at regular intervals in other elements.
   - He left gaps for undiscovered elements
   - He switched some pairs of elements in his table so they would fit in the with the properties expected in that group
   - Transition metals did not have a separate block
4. Mosely: Arranged elements in order of increasing atomic number.
   Def^{n}: The **atomic number** of an atom is the number of protons in the nucleus of the atom.
   Def^{n}: **Modern Periodic Law**: When elements are arranged in order of increasing atomic number, the properties of the elements recur periodically, i.e. the properties displayed by the element are repeated at regular intervals in other elements.
   - No gaps
   - Transition metals are in a separate block

Mass Numbers and Isotopes:

Def^{n}: The **mass number** of an element is the sum of the number of protons and neutrons in the nucleus of an atom of that element.

No. of neutrons in an atom = Mass Number (A) – Atomic Number (Z)

Def^{n}: **Isotopes** are atoms of the same element (i.e. they have the same atomic number) which have different mass numbers due to the different number of neutrons in the nucleus.



Def^{n}: **Relative atomic mass (A_r)** is the average of the mass numbers of the isotopes of an element, as they occur naturally, taking their abundances into account and expressed on a scale relative to 1/12th the mass of ^{12}C.

To calculate the relative atomic mass of an element:

Example

A sample of boron contains 18.7% ^{10}B and 81.3% ^{11}B. Calculate the relative atomic mass of boron.

Solution:

For 100 atoms,

18.7 of them have a mass of 10: 18.7 x 10 = 187

<u>81.3 of them have a mass of 11: 81.3 x 11 = 894.3</u>

Total mass of 100 atoms = 1081.3

Average mass of 1 atom = 1081.3 ÷ 100 = 10.813 ← This is our relative atomic mass.

Electron Configuration of Atoms:

Def^{n}: **Aufbau Principle:** When building up the electron configuration of an atom in its ground state, the electrons occupy the lowest available energy level.

Def^n: **Hund's Rule of Maximum Multiplicity** states that when two or more orbitals of equal energy are available, the electrons occupy them single before filling them in pairs.

Def^n: The **Pauli Exclusion Principle** states that no more than two electrons may occupy an orbital and they must have opposite spins.

The sublevels, in order of increasing energy are: 1s, 2s, 2p, 3s, 3p, 4s, 3d.

If the highest energy sublevel containing electrons is a p sublevel, split the sublevel into its p_x, p_y and p_z orbitals

Examples:

1. *Write the (s, p, etc.) electron configuration of Ar.*

Argon: 18 e^-

Solution : $1s^2$,$2s^2$,$2p^6$,$3s^2$,$3p^6$

3 Energy levels occupied (n=1, n=2, n=3: the BIG numbers)

5 Sublevels occupied (1s, 2s, 2p, 3s, 3p: the number of letters)

9 Orbitals occupied (1s, 2s, $2p_x$, $2p_y$, $2p_z$, 3s, $3p_x$, $3p_y$, $3p_z$: the number of letters, including the splitting of p and d sublevels)

2. *Write the (s, p, etc.) electron configuration of O^-*

O^- : 8+1 = 9 e^-

Solution : $1s^2$, $2s^2$, $2p_x^2$, $2p_y^2$, $2p_z^1$

Note: As the 2p sublevel was not filled, we need to split it into the p_x, p_y, and p_z orbitals. These orbitals are filled singly at first, then doubly.

2 Energy levels occupied

3 Sublevels occupied (1s, 2s, 2p: we count these "unsplit" – the $2p_x$, $2p_y$, $2p_z$ orbitals make up the 2p sublevel)

5 Orbitals (We count these with fully split up orbitals – 3 orbitals in a p sublevel, 5 orbitals in a d sublevel)

3. *Write the (s, p, etc.) electron configuration of Fe^{3+}:*

Fe^{3+}: 26-3 = 23 e^-

Solution: $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^3$

4 Energy levels occupied

7 Sublevels occupied

13 Orbitals occupied (1s, 2s, $2p_x$, $2p_y$, $2p_z$, 3s, $3p_x$, $3p_y$, $3p_z$, 4s, and 3 singly filled d orbitals)

4. *Write the (s, p, etc.) electron configuration of Cr:*

Cr: 24 e^- (this is an important number as it is an exception – so is 29 e^-)

Solution: $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $\underline{4s^1, 3d^5}$ NOT $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $\underline{4s^2, 3d^4}$

Note: 1 e^- has moved from the 4s sublevel to the 3d sublevel. This is because filled and half-filled sublevels are extra stable. By moving this 1 e^- we now have a half filled 4s sublevel and a half-filled 3d sublevel. This moving of electrons ONLY occurs for elements/ions with 24 or 29 electrons.