UNIT 02: Atoms and Atomic Theory

Brief history of atomic theory

Circa. 400-5 BC. Greek philosopher Democritus proposes the idea of matter being made up of small, indivisible particles (*atomos*).
Late 18th Century. Lavoisier proposes the Law of conservation of mass and Proust proposes the Law of constant composition.
Early 19th Century. Using the previously unconnected ideas above, John Dalton formulates his Atomic Theory.

Dalton's atomic theory

LO 1.13

1. Elements are made from tiny particles called atoms.
2. All atoms of a given element are identical (N.B., see isotopes).
3. The atoms of a given element are different to those of any other element.
4. Atoms of different elements combine to form compounds. A given compound always has the same relative numbers and types of atoms. (Law of constant composition).
5. Atoms cannot be created or destroyed in a chemical reaction they are simply rearranged to form new compounds. (Law of conservation of mass).

Structure of the atom and the periodic table

Several experiments were being carried out in the 19th and 20th centuries that began to identify the sub-atomic particles that make up the atom. A summary of those experiments is given below.

Scientist	Experiment	Knowledge gained	Relating to
Crookes	Cathode Ray Tube	Negative particles of some kind exist	Electron
J. J. Thomson	Cathode Ray Deflection	Mass/charge ratio of the electron determined	Electron
Millikan	Oil Drop Experiment	Charge on the electron	Electron
Rutherford, Marsden and Geiger	Gold Foil Experiment	Nucleus present in atom	The nucleus of an atom and the proton

In the first part of the 20th Century, Bohr built upon Rutherford's idea by introducing quantum theory to the *Solar System Model*, and proposed the idea that the atom was made up of a nucleus containing protons, that was being orbited by electrons, *but only in specific, allowed orbits*. Schrödinger subsequently expanded upon Bohr's model, in order to incorporate the wave nature of the electrons. Once Chadwick's discovered the neutron in 1932, the modern picture of the atom *in its simplest form* was complete.

Particle	**Charge**	**Mass in atomic mass units (amu)**	**Position in atom**
PROTON	+1	1	Nucleus
NEUTRON	0	1	Nucleus
ELECTRON	-1	$\frac{1}{1836}$	Outside of the nucleus

The atomic numbers (in the periodic table below shown above the element symbol and sometimes referred to as Z) and mass numbers (in the periodic table below shown below the symbol and sometimes referred to as A) have specific meanings.

Atomic number = the number of protons in the nucleus of one atom of the element

Since all atoms are neutral it also tells us the number of electrons surrounding the nucleus.

N.B., when atoms lose or gain electrons the proton and electron numbers become unbalanced and the atoms become charged particles, i.e., they are no longer neutral. These charged particles are called *ions*. A negative ion is formed when an atom gains electrons to possess a greater number of electrons than protons, and is called an *anion*. A positive ion is formed when an atom loses electrons to possess a fewer number of electrons than protons, and is called a c*ation*.

Mass number = the number of protons + the number of neutrons in one atom of the element

Period	GROUP																	
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1																	2 He 4
2	3 Li 7	4 Be 9	T R A N S I T I O N M E T A L S										5 B 11	6 C 12	7 N 14	8 O 16	9 F 19	10 Ne 20
3	11 Na 23	12 Mg 24											13 Al 27	14 Si 28	15 P 31	16 S 32	17 Cl 35.5	18 Ar 36
4	19 K 39	20 Ca 40	21 Sc 45	22 Ti 48	23 V 51	24 Cr 52	25 Mn 55	26 Fe 56	27 Co 59	28 Ni 59	29 Cu 64	30 Zn 65	31 Ga 70	32 Ge 73	33 As 75	34 Se 79	35 Br 80	36 Kr 84
5	37 Rb 86	38 Sr 88	39 Y 89	40 Zr 91	41 Nb 93	42 Mo 96	43 Tc 99	44 Ru 101	45 Rh 103	46 Pd 106	47 Ag 108	48 Cd 112	49 In 115	50 Sn 119	51 Sb 122	52 Te 128	53 I 127	54 Xe 131
6	55 Cs 133	56 Ba 137	57 La* 139	72 Hf 178	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 Tl 204	82 Pb 207	83 Bi 209	84 Po 210	85 At 210	86 Rn 222
7	87 Fr 223	88 Ra 226	89 Ac† 226	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

*Lanthanides	58 Ce 140	59 Pr 141	60 Nd 144	61 Pm 147	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 163	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
†Actinides	90 Th 232	91 Pa 231	92 U 238	93 Np 237	94 Pu 242	95 Am 243	96 Cm 247	97 Bk 251	98 Cf 251	99 Es 254	100 Fm 253	101 Md 256	102 No 254	103 Lr 257

KEY:

Metal	Semi Metal	Non-metal

13 Al 27	14 Si 28	15 P 31

In this example Al is a metal, Si is a semi-metal (metalloid) and P is a non-metal.

Task 02a

1. ***Determine the number of protons, electrons and neutrons in,***

   (a) $^{210}Pb_{82}$

   (b) $^{34}S_{16}$

2. ***Using only the periodic table above, determine how many elements within the first 20, have atoms with;***

   (a) The same numbers of protons and electrons

   (b) The same numbers of protons and neutrons

Isotopes

Elements occur in nature as a number of different *isotopes*. Atoms with the same number of protons and electrons, *but different numbers of neutrons* are called isotopes. This leads to the modification of the postulate in Dalton's atomic theory that claimed all atoms of a given element were identical, to more accurately state;

All atoms of the same element contain the same number of protons and electrons but may have different numbers of neutrons.

Since it is the electrons in atoms that affect the chemical properties of a substance, isotopes of the same element have the same chemical properties.

Task 02b

1. Consider the following pairs. Does either pair represent a pair of isotopes?

(a) $^{40}K_{19}$ ***and*** $^{40}Ar_{18}$

(b) $^{90}Sr_{38}$ ***and*** $^{94}Sr_{38}$

All periodic tables have atomic mass numbers that are *not* integers. What does this mean? A good starting point is to analyze what it does *not* mean. For example, the atomic mass of Cl is often quoted on periodic tables as 35.5, and may be represented thus; $^{35.5}Cl_{17}$. This does *not* mean that there are 17 protons, 17 electrons and 18.5 neutrons in an atom of chlorine. It is not possible to have a fraction of a neutron in an atom. So what does it mean, and where does the '0.5' come from?

The non-integer values mean that there is more than one isotope of chlorine that exists in nature, in the case of chlorine, ^{35}Cl and ^{37}Cl. A quick calculation shows that these two species have the same number of protons and electrons, but different, whole numbers of neutrons (18 and 20 respectively). That is, they are isotopes of one another. These isotopes happen to exist naturally in the approx. abundance, ^{35}Cl, 75 % and ^{37}Cl, 25 %.

A simple calculation can be applied to calculate the average atomic mass when considering all the isotopes present in a natural sample.

$$\text{Average atomic mass} = \frac{\Sigma(\text{\% of each isotope})(\text{atomic mass of each isotope})}{100}$$

In this case,

$$\text{Average atomic mass} = \frac{((35)\,(75)) + ((37)\,(25))}{100} = 35.5$$

Another example is provided by boron. Boron has two isotopes, ^{10}B and ^{11}B. They have the abundance 18.7 % and 81.3 % respectively.

$$\text{Average atomic mass} = \frac{((10)\,(18.7)) + ((11)\,(81.3))}{100} = 10.8$$

Task 02c

1. ***Neon has three isotopes of masses 22, 21 and 20 amu. If the isotopes have the abundance 8.01 %, 1.99 % and 90.00 % respectively, what is the average atomic mass of neon atoms?***

2. ***A naturally occurring sample of an element consists of two isotopes, one of mass 85 amu and one of mass 87 amu. The abundance of these isotopes is 71 % and 29 % respectively. Calculate the average atomic mass of an atom of this element.***

3. ***If the two isotopes of gallium, ^{69}Ga and ^{71}Ga occur in the respective percentages of 62.1 and 37.9, calculate the average atomic mass of gallium atoms.***