Basic Skills for Chemistry
CHEM-1020
Chapter 4 Lecture Notes
Kroschwitz, 3rd edition

The Atom

What are Atoms?
The atom is the smallest representative particle of an element.
“Atom” means indivisible.
Democritus’ Though Experiment about the existence of atoms.

(Historically not always easy to know if a substance was an *element* or a *compound*.
Water, etc.)

John Dalton ca 1803 updated the ancient Greek notion of the atom and assigned atomic weights to atoms, based on H = 1.

Atoms are the fundamental particles of elements.
(An element is defined as a substance made of one kind of atom.)
(Review some examples of elements.)

Dalton’s Postulates (& modern view of each):

1. All matter consists of *atoms*, small hard eternal spheres.
2. Atoms of a particular element are identical in properties, mass, etc.
3. Atoms of different elements have different properties and masses.
4. Atoms combine is small whole number ratios to form compounds. (Dalton’s “Hooks”)
5. Atoms are never divided, destroyed or changed into other atoms during chemical change

Dalton was the first to propose chemical formulas such as CO_2, H_2O, NaCl, $CaCl_2$ etc.

Atoms are so small they cannot be directly observed by light.
(Many 19th century scientists resisted the idea of something that could not be observed)
e.g. 1 g of H contains 6.022×10^{23} H atoms. One H atom weighs 1.6×10^{-24}g
Now atoms can be easily studied by many modern means.

Atoms cannot be “cut” by any nonchemical means.
However, atoms can be subdivided by high-energy (nonchemical) means into *subatomic particles*.

Three subatomic particles

Particle Name	Electric Charge	Mass (grams)	Mass (amu)	Location	Symbols
Proton	+1	1.67262×10^{-24}	1.0072764 ≈ 1	nucleus	p $^{1}_{1}p$ $^{1}_{+1}p$ $^{1}_{1}H$
Neutron	0	1.67493×10^{-24}	1.0086649 ≈ 1	nucleus	n $^{1}_{0}n$
Electron	-1	9.10939×10^{-28}	0.0005485 ≈ 0	electron cloud	e $^{0}_{-1}e$ $^{0}_{-1}\beta$ β

Many historical views of the internal structure of atoms.
Plum pudding model
Bohr Planetary Model
Modern Quantum Mechanical Atomic Model

Contemporary view of the atom is a dynamic high-energy nucleus surrounded by a dynamic electron cloud.
Electrons in the cloud have various specific energies.
The gain, loss or sharing of electrons allows atoms to combine in small whole number ratios.

Properties of Atoms

Atomic Number:
The atomic number is the number of protons in the nucleus.
Atomic number, Z, determines atomic identity.
Henry Moseley was the first to determine Z values for each element.
Atomic number is a fundamental invariant quantity of each element, unlike atomic mass

Atomic numbers shown in the periodic table.
Atomic numbers are the reason for the sequence of the elements.
(*Layout* of the periodic table to be discussed later.)
Atomic symbols may be shown with or without the (redundant) Z value at the lower left.

Isotopes
Atoms of the same atomic number (same chemical identity) having different numbers of neutrons.
Therefore different atomic masses.
(Dalton did not anticipate atoms of the same element with different atomic masses.)
Describe isotopes of H, F, Cl, Sn and U.
Relation of elements (eg Cl) with several isotopes to empirical atomic masses.

Mass Number:
Mass number, "A", not shown in the periodic table.
A value is shown at upper left of atomic symbol only when discussing specific isotopes.
Again, H, F, Cl.

Ions

Ion: electrically charged atom or group of atoms.
Definition of electrically neutral atom
Tendency of some (noble gas) atoms to retain all their electrons and remain electrically neutral.
Tendency of some (metal) atoms to lose electrons and become *cations*.
Tendency of some (nonmetal) atoms to lose electrons and become *anions*.

Examples of common ions:
Na^+ (cf Na or Na^0) analogy to K and Li
Ca^{2+} (cf Ca or Ca^0) analogy to Mg and Baa
Al^{3+} (cf Al or Al^0)
Cl^- and F^-
O^{2-} and S^{2-}

Ionic charges of *some* atoms can be predicted
Charge of *transition metal* ions cannot be easily predicted.
Group 1 or IA
Group 2 or IIA
Group 3 or IIIA
Group 18 or 8A
Group 17 or 7A
Group 16 or 6A

Counting protons, neutrons and electrons in any atomic or ionic *species*.

H	H^+,	H^-
F	F^-	
O^{2-}	O	
Na	Na^+	
Fe	Fe^{2+}	Fe^{3+}

Relative Atomic Mass

Many atomic masses listed in the periodic table are far from whole numbers.
Two reasons:

1) Most elements consist of mixtures of isotopes.
   Examples: chlorine, hydrogen, carbon.

2) Atomic Mass Deficit (more subtle and harder to understand)
   Protons and neutrons are held together to form nuclei by the strong nuclear force.
   Energy must be expended to separate a nucleus into individual *nucleons*.
   At the subatomic level, energy and mass are equivalent.
   When a system gains energy, it gains mass. When a system loses energy, its mass decreases.

   The carbon-12 nucleus contains 6 protons of mass 6 x 1.007276 amu = 6.043756 amu
   plus 6 neutrons of mass 6 x 1.008664 amu = 6.051984 amu
   Separately, the nucleons would weigh 12.095740 amu

But when nucleons come together, they give off energy (binding energy) and lose an equivalent amount of mass, called the *mass deficit*.
It is impossible to predict the mass of a nucleus by adding up the masses of its individual nucleons because they will lose some mass when they are brought together.
Therefore, the atomic mass of each atomic nucleus must be determined experimentally.

Since there is no absolute atomic mass scale, a *relative atomic mass* scale was devised.
The carbon-12 nucleus was defined to have a mass of 12.000000… amu.
All other atomic masses are determined relative to the 12.0000… amu of the C-12 nucleus.

Why does the periodic table show carbon having a mass of 12.011 amu.
Why does fluorine, which has only the F-19 isotope, have atomic mass 18.998404 amu?

<u>Periodic Table Layout</u>

Dmitri Mendeleev's Contribution

Mendeleev observed that some elements have similar properties and can be grouped in "families".
Mendeleev had no knowledge of *atomic numbers*.
Mendeleev arranged the known elements by atomic mass and found that properties repeated *periodically*
The resulting *periodic table* grouped similar elements in vertical groups.
Gaps in the periodic table predicted the existence of undiscovered elements.
Problem posed by I and Te.

Modern Periodic Table

Elements arranged by *Atomic Number* (much better criterion than atomic mass)
Grouping of elements by groups and families
(Solution of I and Te switch)
Periods of varying lengths
Separation of Metals and Nonmetals. (Metalloids)
General properties of metals and nonmetals
Numbering of groups by old and new systems
(No undiscovered elements await discovery)

Families of Elements

IA, Alkali Metals
IIA, "Alkaline Earths"
IIIA,
IVA, Carbon Group
VA,
VIA Oxygen Group,
VIIA, Halogens
VIIIA Noble Gases
Groups within the "Transition Elements"