Chapter 4
What are Atoms?

For over two thousand years, learned people wondered what matter was made of. In the 1800s, scientists found evidence that all matter was comprised of atoms. Now in the twenty-first century, images of atoms can be seen. Here we will explore the structure of the atom, a key component in understanding the fundamentals of chemistry.

4.1 Background

In 460 BCE, the Greek philosopher Democritus reasoned that if you keep cutting a piece of matter into smaller and smaller fragments, there will be some point at which the pieces cannot be made any smaller. He called these *a-tomos* which in English would be "un-cuttables". This proposal was just based on philosophy. It was not until the early 1800s that properties of matter were studied in a quantitative way and the only reasonable explanation was that matter consisted of particles. In this work, the pioneer was a British scientist, John Dalton. His primary interest was searching for patterns in the weather. Dalton was unsuccessful, but his proposals of atomic theory are still largely valid in modern chemistry.

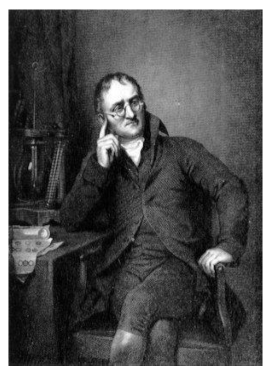


Figure 4.1 (left) Democritus, Greek philosopher (460-340 B.C.) and (right) John Dalton, British scientist (1766-1844)

Dalton proposed that:

1. All matter is composed of extremely small particles - atoms
2. Atoms cannot be subdivided
3. Atoms cannot be created or destroyed
4. All atoms of the same element are identical in mass, size, and chemical & physical properties
5. The properties of the atoms of one element differ from those of all other elements
6. Atoms combine in small whole-number ratios to form compounds.

However, in the early twentieth century, physicists discovered that atoms were not solid spheres as Dalton had implied, but that they contained charged particles. The most important individual discovery was that of Ernest Rutherford (1871-1937). Rutherford, who did much of his early research at McGill University in Montreal, showed that most of the atom was empty space. His research indicated that atoms had hard positively-charged 'cores' with negatively-charged particles in the remaining space. It is the Rutherford Model which will be the starting point for our study of the atom in the context of chemistry.

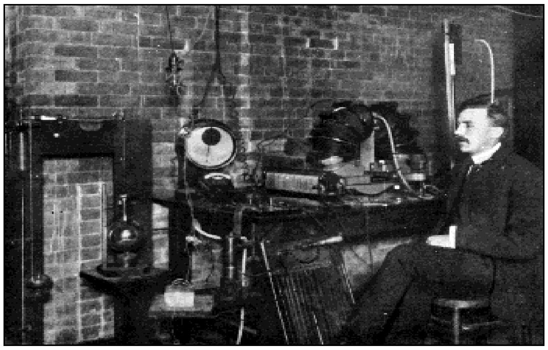


Figure 4.2 Ernest Rutherford at work in his laboratory.

4.2 The Structure of the Atom

We now know that there are three particles which comprise an atom. In the core – the nucleus – there are positive particles, called protons, and neutral particles, called neutrons. Around the nucleus, there are smaller and much lower mass negative particles – electrons. The properties of the particles are summarized below.

Table 4.1 The particles in the atom

Particle	**Symbol**	**Relative mass**	**Relative charge**
proton	p^+	1	+1
neutron	n	1	0
electron	e^-	0.0005	−1

The term ***atomic number*** is used for the number of protons in the atom. Each chemical element is a specific number of protons. In a neutral atom, the number of protons equals the number of electrons.

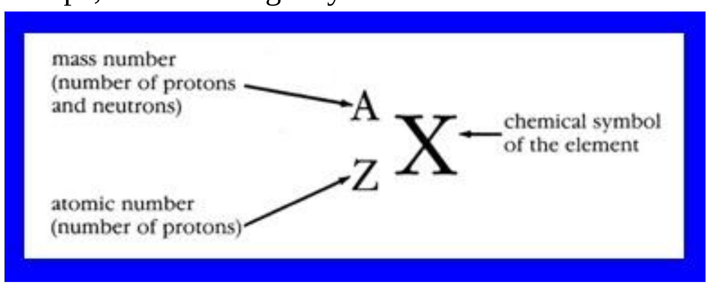
The sum of the number of protons and neutrons is called the ***mass*** number. The number of neutrons is usually greater than the number of protons. For some chemical elements, there is only one specific number of neutrons. For example, all the fluorine atoms found in nature have ten neutrons. Other elements are found with differing numbers of neutrons. For example, 75% of chlorine atoms in nature have 18 neutrons while 25% have 20 neutrons. There is no way of predicting the number of neutrons, it is necessary to check tables of data to find out this information.

When discussing a particular element, there is a special term – ***isotope*** – to indicate which number of neutrons that an atom possesses. Thus all fluorine atoms consist of one naturally-occurring isotope while chlorine is a mixture of two naturally-occurring isotopes.

Figure 4.3 (left) it was the British chemist, Frederick Soddy (1877-1956), who first proposed the existence of atoms of the same element with different numbers of neutrons while (right) it was Ellen Gleditsch (1879-1968) whose precise measurements confirmed the existence of isotopes.

To write a particular isotope, the following way is used:



Physicists, particularly, prefer the display method above. However, most chemists find it easier to write the mass number following the name of the element (such as fluorine-19) as the name of the element itself defines the number of protons.

Chemists are usually concerned with the stable isotopes that exist in nature. Many elements have isotopes, whose nucleus disintegrates over time, converting the atom to that of a different element. Such isotopes are said to be radioactive. A ***radioactive isotope*** is an isotope that spontaneously changes into another isotope (often of another element), usually with the emission of a sub-atomic particle and radiation.

Figure 4.4 (left) a sample of the stable cobalt-58; (right) A teletherapy machine containing radioactive cobalt-60, the emitted radiation is used to treat solid tumours

For some elements, such as uranium, all atoms are radioactive. The degree of danger from a radioactive isotope depends upon the ***half-life***, that is, the time that it takes for half of the atoms in a sample to change to another isotope. If the half-life is very long, then it is reasonably safe to work with. For example, uranium-238 has a half-life of 4.5×10^9 years, that is, it takes over a billion years for half the uranium atoms in a sample to change into other isotopes.

Figure 4.5 This cup is made of uranium-containing glass, it has a nice greenish-colour and it glows harmlessly under ultraviolet light. The very long half-life of uranium makes it safe to handle.

4.3 Atomic Mass

In Table 4.1, the protons and neutrons were listed as having masses close to one relative unit. The electron has such a small mass we can ignore that. So using this non-SI scale, one can say that the mass number, the sum of the number of protons and neutrons, will have a value close to the mass of the atom in these relative units. Many chemists use the term ***unified atomic mass units*** , symbol u, for these masses. It is called the unified atomic mass unit as previously there had been two slightly different atomic mass units, one used by physicists and one by chemists. Thus “unified” indicates that it is the value used by all sciences today (the unified atomic mass unit is also called a Dalton, symbol, Da). For example, an atom of naturally-occurring fluorine has nine protons and ten neutrons. It has a mass number of 19 and the relative mass of the atom would be approximately 19 u. From precise measurements using this scale, the mass is, in fact, 19.00 u.

AVERAGE ATOMIC MASS

As mentioned above, most elements are found as mixtures of isotopes. An instrument called a mass spectrometer can be used to separate the isotopes and to measure the percentage abundance of each. The percent abundance is the proportion of one isotope in 100 atoms of that element. Figure 4.5 shows the results of studies on the element, neon. Thus in a sample of 100 atoms of neon, 90.92 atoms will have a mass of 20, 0.26 will have a mass of 21, and 8.82 will have a mass of 22. As percentages, this would be 90.92%; 0.26%; and 8.82%.

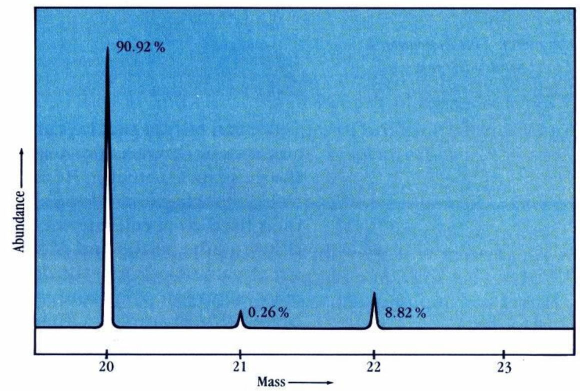


Figure 4.6 A plot of the abundance of the isotopes of neon

From such data, it is possible to determine the ***average atomic mass*** of an element, that is, the average mass of an atom. This value is of particular importance in chemistry. To illustrate, the average atomic mass of neon will be calculated.

EXAMPLE 4.1

Calculate the average atomic mass of neon. The precise masses and the abundances are as follows:

Isotope	Atomic Mass (u)	Abundance (%)
neon-20	19.99	90.92
neon-21	20.99	0.26
neon-22	21.99	8.82

Answer

First, each of the abundances has to be expressed as parts of one by dividing by 100%

This gives abundances of: 0.9092; 0.0026; and 0.0882.

Then each isotope atomic mass and the corresponding abundance are multiplied and the results added together.

Average atomic mass = (19.99 u)(0.9092) + (20.99 u)(0.0026) + (21.99 u)(0.0882)

= (18.17 u) + (0.055 u) + (1.94 u)

= 20.17 u

A value close to, and just above 20 would be expected as such a high proportion of the atoms have a mass of 20.

THE DIFFERENCE BETWEEN MASS NUMBER AND ATOMIC MASS

The mass number (sometimes called the nucleon number) represents the sum of the number of protons and neutrons in an atom. It is always an integer. The atomic mass is the mass of one atom of an element. On the unified atomic mass scale, both a proton and a neutron have a mass of about 1 u while an electron has a negligible mass of 0.0005 u. Thus it might be expected that, for a particular isotope, the mass number would be essentially identical to the atomic mass.

This is not quite true. For example, in nature, all atoms of manganese, atomic number 25, have 30 neutrons to give a mass number of 55. However, the atomic mass of manganese is 54.94 u. The explanation for the 'missing' 0.06 u is to be found in the well-known equation, $E = mc^2$, where E is the energy corresponding to mass m. When the 25 protons and 30 neutrons combine to form the nucleus of a manganese-55 atom, energy is released, a quantity of energy equivalent to the loss of 0.06 u of mass for every atom. As a result, for atoms of nearly all elements, the atomic mass is slightly less than the mass (nucleon) number.

4.4 The Mole

Atoms are incredibly tiny and, just as eggs are counted in dozens, so atoms are counted in big groupings called the amount of substance. The unit for the amount of substance is the mole, symbol, mol. The numerical value is $\boldsymbol{6.02 \times 10^{23}}$ and it is known as Avogadro's number, named after the Italian scientist, Amadeo Avogadro (1776-1856), and it is sometimes given the symbol N.

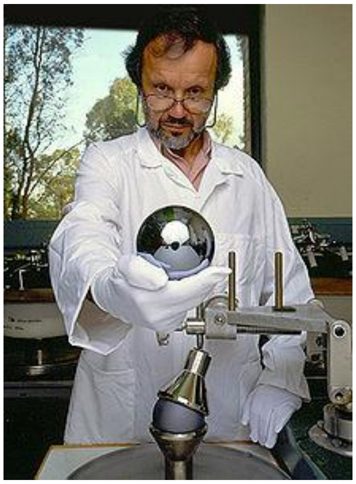


Figure 4.7 A one-kilogram sphere of isotopically-pure silicon-28 is now used to measure Avogadro's constant to high precision, the current value being $6.02214078 \times 10^{23}$ mol^{-1}

It is such a significant number in chemistry that science classes at some grade schools celebrate the day – often by having a mole day starting at 6:02 a.m. on 23rd October. Below is one of the commercial Mole Day cards available to send to your chemistry friends.

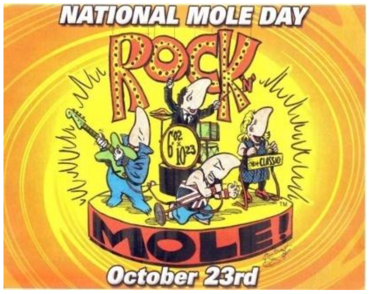


Figure 4.8 A postcard to celebrate Mole Day

Using the relationship that 1 mol ≡ 6.02×10^{23} atoms, it is straightforward to convert moles to numbers of atoms and the reverse.

Example 4.2
Calculate the number of atoms of carbon in 2.67 mol of carbon.
Answer

Strategy	Relationship
mol → # of atoms	1 mol ≡ 6.02×10^{23} atoms

$$\text{Number of atoms carbon} = 2.67\ \text{mol} \times \left(\frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}\right) = 1.61 \times 10^{23}\,\text{atoms}$$

Example 4.3
Calculate the moles of sulfur in 5.97×10^{22} atoms of sulfur.
Answer

Strategy	Relationship
# of atoms → mol	1 mol ≡ 6.02×10^{23} atoms

$$\text{Mol of sulfur} = 5.97 \times 10^{22}\ \text{atom} \times \left(\frac{1\ \text{mol}}{6.02 \times 10^{23}\ \text{atoms}}\right) = 9.87 \times 10^{-2}\ \text{mol sulfur}$$

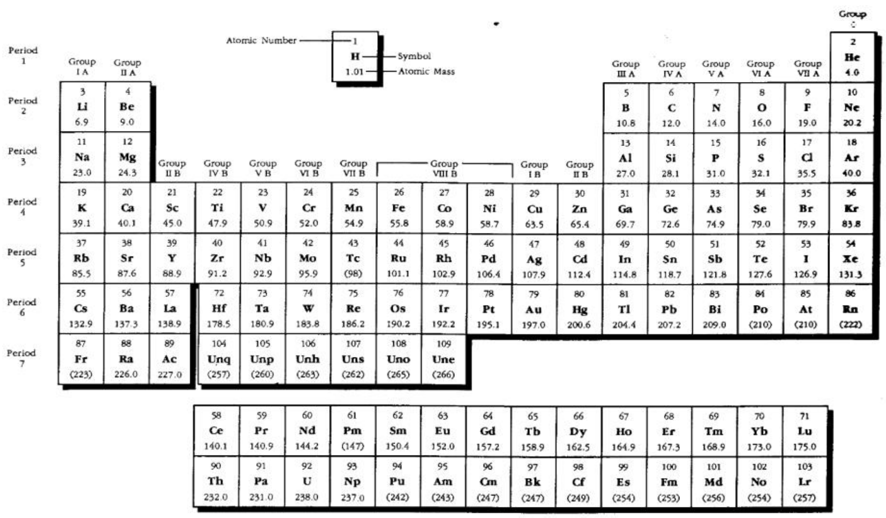
4.5 Molar Mass

Though it is useful to be able to convert moles of an element to the number of atoms, unfortunately, there are no balances that measure in moles. However, there is a way of taking the mass measurement from a balance and converting it to units of moles. The unit factor differs between elements and it is so arranged as to be the same numerical value in grams as the mass of one atom is in unified atomic mass units. The standard is the isotope carbon-12. Thus:

1 atom of carbon-12 has a mass of exactly 12 unified atomic mass units
1 mol of carbon-12 atoms has a mass of exactly 12 grams

Figure 4.9   1 mol (6.02×10^{23} atoms) of carbon-12 has a mass of exactly 12 g

Thus the molar mass is the mass of one mole of an element expressed in grams. So for each element, one must know the average atomic mass in order to calculate the number of moles. Fortunately, these numerical values are displayed in the standard Periodic Table. The details of the Periodic Table will be described in the next chapter. However, here it is important to realize that the number underneath the symbol of an element is the atomic mass (in unified atomic mass units) and, at the same time, the molar mass (in grams). Though many tables list atomic/molar masses to very high precisions, usually three significant figures are ample.



Atomic Number — 1
Symbol — H
Atomic Mass — 1.01

	Group I A	Group II A	Group II B	Group IV B	Group V B	Group VI B	Group VII B	Group VIII B			Group I B	Group II B	Group III A	Group IV A	Group V A	Group VI A	Group VII A	Group 0
Period 1																		2 He 4.0
Period 2	3 Li 6.9	4 Be 9.0											5 B 10.8	6 C 12.0	7 N 14.0	8 O 16.0	9 F 19.0	10 Ne 20.2
Period 3	11 Na 23.0	12 Mg 24.3											13 Al 27.0	14 Si 28.1	15 P 31.0	16 S 32.1	17 Cl 35.5	18 Ar 40.0
Period 4	19 K 39.1	20 Ca 40.1	21 Sc 45.0	22 Ti 47.9	23 V 50.9	24 Cr 52.0	25 Mn 54.9	26 Fe 55.8	27 Co 58.9	28 Ni 58.7	29 Cu 63.5	30 Zn 65.4	31 Ga 69.7	32 Ge 72.6	33 As 74.9	34 Se 79.0	35 Br 79.9	36 Kr 83.8
Period 5	37 Rb 85.5	38 Sr 87.6	39 Y 88.9	40 Zr 91.2	41 Nb 92.9	42 Mo 95.9	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
Period 6	55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (210)	85 At (210)	86 Rn (222)
Period 7	87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Unq (257)	105 Unp (260)	106 Unh (263)	107 Uns (262)	108 Uno (265)	109 Une (266)									

58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (147)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.2	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (249)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (254)	103 Lr (257)

Figure 4.10 In the Periodic Table, the average atomic mass is displayed underneath the symbol of the element.

EXAMPLE 4.4

Calculate the moles of zinc in 7.04 g of zinc.

Answer

<u>Strategy</u>
mass Zn → mol Zn

<u>Relationship</u>
1 mol ≡ 65.4 g Zn

$$\text{Mol of zinc} = 7.04\ \text{g} \times \left(\frac{1\ \text{mol}}{65.4\ \text{g}}\right) = 0.108\ \text{mol zinc}$$

EXAMPLE 4.5

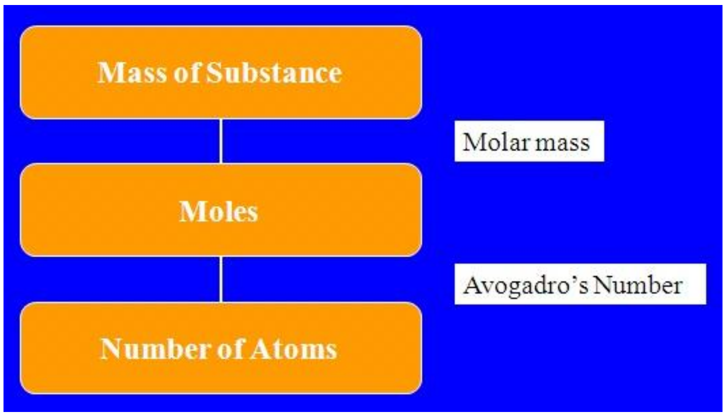
Calculate the mass of 5.00×10^{-2} mol of silicon.

Answer

Strategy	Relationship
Mol of Si $\rightarrow$ mass Si	1 mol $\equiv$ 28.1 g Si

$$\text{Mass of silicon} = 5.00 \times 10^{-2}\ \text{mol Si} \times \left(\frac{28.1\ \text{g}}{1\ \text{mol}}\right) = 1.41\ \text{g silicon}$$

The two calculations: mass$\leftrightarrow$mol and mol$\leftrightarrow$atoms, can be combined as is shown in the following flow chart.



EXAMPLE 4.7

Calculate the number of atoms in 14.8 g of phosphorus.

Answer

Strategy	Relationship
Mass of P $\rightarrow$ mol P	1 mol $\equiv$ 31.0 g P
Mol of P $\rightarrow$ # of atoms P	1 mol $\equiv 6.02 \times 10^{23}$ atoms

$$\text{Mol of P} = 14.8\ \text{g P} \times \left(\frac{1\ \text{mol}}{31.0\ \text{g}}\right) = 0.477\ \text{mol P}$$

$$\#\ \text{of atoms P} = 0.477\ \text{mol P} \times \left(\frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \text{mol}}\right)$$

$$= 2.87 \times 10^{23}\ \text{atoms phosphorus}$$

EXAMPLE 4.8

Calculate the mass of 2.50×10^{22} atom of magnesium.

Answer

Strategy	Relationship
# of atoms Mg → mol Mg	1 mol ≡ 6.02×10^{23} atoms
Mol of Mg → mass Mg	1 mol ≡ 24.3 g

$$\text{Mol Mg} = 2.50 \times 10^{22}\ \text{atom Mg} \times \left(\frac{1\ \text{mol}}{6.02 \times 10^{23}\text{atoms}}\right) = 4.15 \times 10^{-2}\ \text{mol Mg}$$

$$\text{Mass of Mg} = 4.15 \times 10^{-2}\ \text{mol Mg} \times \left(\frac{24.3\ \text{g}}{1\ \text{mol}}\right) = 1.01\ \text{g magnesium}$$

Where Next?

In this chapter, the focus has been on the protons and neutrons in the atom. In the next chapter, you will discover that it is the tiny electrons that control the chemistry of each element.