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TEACHING CHEMISTRY PROGRESSIVELY: FROM SUBSTANCES, TO ATOMS AND MOLECULES, TO ELECTRONS AND NUCLEI

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ABSTRACT: In most chemistry courses today, students are introduced to atoms, molecules, ions, and electrons early in the course, and have to accept the teacher's word that these exist. A better method is to teach chemistry progressively, starting with observations at a macroscopic level, interpreting these at an atomic and molecular level, and then at an electronic and nuclear level. A modern way of doing this is described. [*Chem. Educ. Res. Pract. Eur.*: 2002, 3, 215-228]

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1. INTRODUCTION

Chemistry is a difficult subject to teach. Several educators have recently identified one reason for this (Tsaparlis, 1997; Jensen, 1998; Laing, 1999; Johnstone, 1999, 2000; Nelson, 1999). This is that chemists describe chemistry at several levels, only one of which can be readily observed. These levels are:

Level one: the macroscopic or bulk level

Level two: the atomic and molecular level

Level three: the electronic and nuclear level

Hydrogen, for example, is described at level one as a colourless, light, inflammable gas; at level two as comprising H₂ molecules; and at level three as comprising pairs of H⁺ nuclei held together by two electrons. Some authors add further levels corresponding to symbolic ways of representing the above.

Only level one can be readily observed; levels two and three have to be inferred from observations at level one. Scanning tunnelling microscopy can detect individual atoms adsorbed on surfaces, but we only know that they are atoms from other considerations. If we just looked at the scans, we could not tell whether the peaks correspond to atoms, or piles of fine-grained or continuous material.

This makes the teaching of chemistry difficult. We want to get to levels two and three as quickly as we can, but if we get to them too quickly we divorce them from the first level, which is the only one that can be perceived directly. This leaves students having to accept what we tell them, and able to do little more than manipulate chemical ciphers. This not only gives them a false impression of chemistry as a science, but also turns off those who do not

like abstract manipulations from doing chemistry. It also leaves those who go on to become chemists weak at level one.

Prior to the 1960s, chemistry in English-speaking countries was taught progressively, starting at level one and gradually working down to level three. In the 1960s, the teaching of chemistry was revolutionized, to bring the lower levels to the beginning. As a result, chemistry has become less popular, and students have become weak at level one. The literature contains numerous examples of this, from the US student who thought that silver chloride is a pale green gas (Davenport, 1970), to the UK graduate who did not know CO_2 reacts with sodium hydroxide (McRobbie, 1988). Johnstone (1999) has described the 1960s revolution, from the point of view of the psychology of learning, as “probably a disaster”.

One way of remedying this problem is to return to the approach taken before the 1960s, and present chemistry historically. This automatically secures the conceptually logical order, level one \rightarrow level two \rightarrow level three (chemical elements were established in the 18th century, the atomic theory in the 19th, and the electronic theory in the 20th). This approach has the advantage of introducing students to the pioneers of chemistry and the way they reasoned. Disadvantages are that it sets chemistry more in the past than the present, and can perpetuate misconceptions in the development of the subject which have subsequently to be unlearned.

An alternative is to use modern knowledge to develop a new approach, following the order: level one \rightarrow level two \rightarrow level three. I describe such an approach in this paper. While this follows the historical development quite closely, it departs from it where there were misconceptions. For example, the approach recognizes that compounds can be nonstoichiometric, and that many substances are nonmolecular. There are also some simplifications. For example, there is no reference to the phenomenon of reciprocal proportions, and Faraday’s laws of electrolysis come after the ionic theory. I have not made any reference to Berzelius’s theory of bonding, thereby allowing Avogadro’s hypothesis to be presented without raising the questions it did historically.

The resulting scheme is similar to those recently described by Hutchinson (2000) and Toomey, DePierro, and Garafalo (2001), but is more radical. Part of it (Sections 2.1–2.4) has been used in a new science foundation course at the Open University in the UK (Johnson, 1998). Georgiadou and Tsaparlis (2000) have proposed a “three-cycle” method of solving the same problem.

2. LEVEL ONE TO LEVEL TWO

2.1. Chemical change

The teacher begins by introducing students to substances and chemical change, described at level one. This leads to the classification of substances as “pure”, “mixtures”, or “solutions”, and of pure substances as “elements” or “compounds”. No reference is made at this stage to atoms or molecules (cf. Tsaparlis & Kampourakis, 2000). This material is a vital part of an introductory chemistry course, and should not be hurried (Bradley, 1988).

2.2. Basic phenomena

The next step is to introduce students to the following phenomena (Nelson, 1994b). These can be demonstrated with suitable experiments (see, e.g., Fowles, 1957; Sienko et al., 1984; Nelson, 1996a).

Law of conservation of mass: “In any chemical change there is no detectable change in mass.” “Chemical” here excludes “radiochemical”, i.e. changes in which high-energy radiation is involved.

Phenomenon of constant composition: Provisional statement: “Many compounds have the same composition by mass, irrespective of how they are made or what they are made from.” I have not called this a law because some compounds have variable composition. The statement is provisional because of isotopic variation.

Phenomenon of multiple proportions: Provisional statement: “When an element A forms more than one compound with another element B, the masses of B that combine with a fixed mass of A are often in the ratio of small whole numbers.”

Phenomenon of proportionate gaseous volumes: “At low pressures, the volumes of gases taking part in a chemical change, measured at the same temperature and pressure, are often in the ratio of small whole numbers.” Deviations occur at high pressures.

2.3. Atoms and molecules

There are several indications that matter may be made up of atoms:

- Many solids are crystalline. This can be explained in terms of the regular packing of small particles (cf. Partington, 1951).
- Gases are much more compressible than liquids or solids, and when they condense there is a large reduction of volume. These observations can be explained if gases comprise separate particles, which come together in the liquid or solid state. These may be atoms or clusters of atoms (molecules).
- When a small quantity of olive oil is poured on to a large pool of water, the oil only spreads over a limited area of the surface (cf. Giles, 1969).

These considerations, along with the phenomena of Section 2.2, lead to the following theory of matter (after Dalton and Avogadro):

1. Matter is made up of atoms.
2. The atoms of an element are all the same, and differ from those of other elements (provisional statement).
3. Chemical reactions involve changes in the way in which atoms are combined, but not in their numbers.
4. Atoms of different elements often combine in definite ratios.
5. These ratios are often of small whole numbers.
6. At low pressures, equal volumes of gases at the same temperature and pressure contain equal numbers of molecules (Avogadro’s hypothesis).

This theory explains the phenomena described in the previous section (2.2). For example, the phenomenon of constant composition follows from postulate 4, and of proportionate gaseous volumes from postulate 6. The latter is illustrated by the reaction between hydrogen and chlorine to form hydrogen chloride, for which the volumes are in the ratio 1:1:2. This is explained if hydrogen comprises H_{2m} molecules, chlorine Cl_{2n} , and hydrogen chloride H_mCl_n (n is not necessarily equal to m).

Postulate 6 enables the masses of molecules to be compared (Nelson, 1994b). For example, the density of hydrogen at STP (standard temperature and pressure) is 0.08988 g/l and of oxygen 1.4290 g/l. So:

$$\frac{\text{Mass of oxygen molecule}}{\text{Mass of hydrogen molecule}} \approx \frac{1.4290}{0.08988} = 15.899$$

The result is approximate because of the pressure; the limiting value at low pressures is 15.875. If the mass of a hydrogen molecule (μ) is provisionally made the unit of mass for atoms and molecules, the mass of an oxygen molecule is therefore about 16 μ .

2.4. Atomic compositions

To establish the atomic composition of a molecule, a further principle needs to be added to 1–6 (after Cannizzaro). This is:

7. The mass of an atom of an element is the smallest mass of the element found in any molecule containing it.

This principle is applied in Table 1. The masses of some molecules containing hydrogen, obtained from gas densities, are given in column two. The percentage by mass of hydrogen in each of these molecules, as determined by quantitative analysis of the substance concerned, is given in column three. The mass of hydrogen in each molecule (column two \times column three \div 100) is given in column four. The smallest value of this is 0.5 μ . If the table is extended to other substances containing hydrogen, the smallest value continues to be 0.5 μ . This is almost certainly, therefore, the mass of a hydrogen atom, since no subdivision of this mass occurs in any molecule. A hydrogen molecule (mass = 1.0 μ) thus contains two atoms (H_2).

TABLE 1: *Calculating the atomic mass of hydrogen.**

Substance	Mass of Molecule	Percentage by Mass of Hydrogen	Mass of Hydrogen in Molecule
Hydrogen	1.0 μ	100.0	1.0 μ
Methane	8.0 μ	25	2.0 μ
Ammonia	8.4 μ	17.7	1.5 μ
Steam	8.9 μ	11.2	1.0 μ
Hydrogen sulfide	16.9 μ	6.0	1.0 μ
Hydrogen chloride	18.1 μ	2.76	0.5 μ
Alcohol vapour	22.7 μ	13.3	3.0 μ
Ether vapour	36.4 μ	13.8	5.0 μ

* Source: Nelson (1994b)

This procedure can be extended to other elements (Nelson, 1994b). Thus, for example, the smallest mass of oxygen occurring in a molecule of any of its compounds is about 8 μ , making this the atomic mass of oxygen. A molecule of oxygen (mass $\approx 2 \times 8 \mu$) is therefore O_2 and of steam (mass $\approx 2 \times 0.5 \mu$ of hydrogen + 8 μ of oxygen), H_2O .

A more convenient unit than μ is the mass of a hydrogen atom, or a standard mass of similar magnitude to this. The standard mass now used is called the unified atomic mass unit (u). This makes the atomic mass of hydrogen 1.0079 u, and of oxygen, 15.999 u (Mills et al., 1993). The definition of u will be given later (Section 3.5).

From the masses of atoms, the atomic composition of any substance can be determined. For example, pure hydrogen peroxide contains about 6.0 per cent by mass of hydrogen and 94.0 per cent of oxygen. Thus the number of atoms of hydrogen in 100 g is about $6.0 \text{ g}/1.01 \text{ u} = 5.9 \text{ g/u}$ and of oxygen $94.0 \text{ g}/16.0 \text{ u} = 5.9 \text{ g/u}$, a ratio of 1:1. A molecule of hydrogen peroxide therefore has the formula $(\text{HO})_n$, and mass $17.0n \text{ u}$. The density of the vapour corresponds to a molecular mass of about 34 u, whence $n = 34/17 = 2$.

In the above calculation g/u is the Avogadro number. Stoichiometric calculations can be greatly simplified by using this as a unit (Nelson, 1991a). In practice, calculations are made more complicated by introducing "amount of substance" and the mole (Mills et al., 1993). Evaluation of the Avogadro number is not urgent, but a good estimate can be obtained from the area over which a measured quantity of oil spreads on water (Guggenheim & Prue, 1956).

The conclusion that matter is made up of atoms and molecules is supported by the results of the kinetic theory of gases. I have not included this theory because it is more physical than chemical, and the existence of atoms and molecules can be inferred without it. But the kinetic theory does provide confirmation of the chemical conclusions, and can usefully be added.

2.5. Molecular and nonmolecular substances

The distinction between molecular and nonmolecular substances is an important one in chemistry, and is best introduced at this point. This can be done without having to appeal to X-ray crystallography by classifying substances on the basis of volatility and solubility as follows (Nelson, 1996b).

Type 1: Volatile substances

These have a low boiling point ($\leq 100 \text{ }^\circ\text{C}$). If they are coloured, they usually have the same colour as a gas, liquid or solid (e.g. chlorine and bromine). Each dissolves in a range of other volatile substances — if coloured, usually without change in colour. These properties suggest that the molecules that are present in the gas (atoms in the case of the inert gases) are generally also present in the liquid and solid, and in solution. A further implication is that forces between molecules are relatively weak.

Type 2A: Semi-volatile and involatile substances that dissolve in volatile ones

These have similar properties to volatile substances but are less volatile. They dissolve in a range of volatile solvents, and preserve their colours in different phases (e.g. coloured organic compounds). Most organic and organometallic compounds fall into this class. Their properties suggest that substances of this type are also generally molecular in every phase (gas, liquid, solid, and in solution).

This conclusion is supported by the freezing points of solutions of type 1 and 2A substances. For dilute solutions, the depression of the freezing point (θ_f) is generally proportional to the fraction of solute molecules (x_{mol}) in the solution, calculated on the

assumption that the solute and solvent molecules in the solution are the same as in the gas phase:

$$\theta_f = k_f x_{\text{mol}} \quad (1)$$

The value of k_f is characteristic of the solvent. Some solutions do not satisfy equation (1), notably those conducting electricity while undergoing electrolysis (e.g. hydrogen chloride in water). These will be considered later (Section 3.2).

The boiling points of substances of types 1 and 2A generally increase with the number of atoms in the molecule, and with the masses of these atoms (though mass itself is not the cause of this). This implies that the forces between molecules of these substances generally increase with the number of atoms and their masses. Type 2A substances can nevertheless dissolve in type 1 substances because, in general, the more atoms a molecule has, the greater the number of solvent molecules it can interact with, and the heavier the atoms are, the stronger these interactions will be.

Type 2B: Semi-volatile and involatile substances that are generally insoluble in volatile ones

These are insoluble in all, or nearly all, volatile substances, other than by chemical reaction. If they are coloured, there is usually a change in colour between the solid or liquid and the gas (e.g. metals). Their boiling points are generally much higher than those of type 1 or 2A substances having similar gaseous molecules (e.g., Zn compared with Ar or Kr, NaCl compared with HCl or ICl, SiO₂ compared with CO₂ or SO₂). This class includes metals, salts, and adamantine substances (diamond, quartz, etc.). Salts generally dissolve in water, but their solutions conduct electricity while undergoing electrolysis, and do not obey equation (1).

The low volatility of these substances compared with type 1 or 2A substances having similar gaseous molecules implies that their gaseous molecules attract each other much more strongly than those of the latter. The change in colour of coloured substances of this type on condensation further implies that these interactions give rise to a different structure in the liquid or solid, one in which the identity of the molecules is lost. These observations suggest that molecules join together in the liquid or solid state to give a continuous network of atoms, with strong bonds throughout the network. Networks can be envisaged in one, two, or three dimensions.

For these substances to dissolve in a volatile solvent, sufficiently strong bonds have to be formed between units of the network and solvent molecules to overcome the bonding between the units. Since the bonds usually formed by molecules of volatile substances will be too weak for this, type 2B substances are usually insoluble. That salts dissolve in water implies a particularly strong kind of bonding between their units and water molecules. This is discussed further later.

In general, X-ray crystallography confirms the above inferences. Exceptions occur when a substance has molecular and nonmolecular forms of similar energy (e.g. dinitrogen pentoxide sublimes at 32 °C, but is NO₂⁺NO₃⁻ in the solid state). The number of exceptions is relatively small. Some substances have an intermediate character (e.g. water), but this can be discussed later in a course (Nelson, 2000).

So, in general, substances of types 1 and 2A can be described as “molecular” and those of type 2B as “nonmolecular”. Molecular substances made up of very large molecules may be described as “macromolecular”. Some authors also describe substances of type 2B as

“macromolecular”, but this term is best reserved for substances like proteins which, in bulk, still comprise many molecules.

The distinction between molecular and nonmolecular substances can be reinforced by formulating the latter as, for example, $(\text{NaCl})_\infty$ or (NaCl) . The formula NaCl goes back to the time when chemists believed sodium chloride to be molecular like ethyl chloride.

Descriptive chemistry

The teacher can now describe the chemistry of important elements at levels one and two. These elements are (Nelson, 1991b):

Nonmetals: H, C, N, O, F, Si, P, S, Cl

Metals: Na, Mg, Al, K, Ca, Cr, Mn, Fe, Ni, Cu, Zn

This material leads to the concept of valency.

2.6. Valency

Valency is a key concept in chemistry, modern theories of bonding notwithstanding (Nelson, 1997a). It can be introduced to students by showing them a chart of the atomic composition of selected substances, and inviting them to see if they can identify any patterns (Nelson, 1997b). With help, students should be able to detect, for example, that compounds of chlorine generally have the same composition as those of hydrogen [e.g., CH_4 , CCl_4 ; (NaH) , (NaCl) ; (CaH_2) , (CaCl_2) ; (AlH_3) , (AlCl_3)]. This implies that a chlorine atom generally has the same capacity for combining with other atoms as a hydrogen atom. Consistent with this, one chlorine atom combines with one hydrogen atom to form HCl .

This discussion leads naturally to the concept of valency, defined as “the capacity of an atom or radical to combine with other atoms or radicals in a family of compounds”. The appropriate valency at an introductory level is the “primary” or “classical” valency, ν (Nelson, 1997a). The value of ν for an atom or radical can be determined by assigning a value of one to hydrogen, and counting the number of hydrogen atoms that X will combine with to form a molecule XH_ν , or a unit of a nonmolecular compound (XH_ν) . Supplementary values can be obtained from the formulae of oxides using $\nu(\text{O}) = 2$.

From these values the composition of other compounds of X can be derived, by applying the principle that, in many compounds, the valencies of all the atoms or radicals are satisfied. This can be done either algebraically or graphically (Nelson, 1997a). The algebraic method is more general, allowing the number of bonds between atoms (n) to be nonintegral (e.g. $n = 1.5$ for the carbon-carbon bonds in benzene). The graphical method is simpler, as in the examples:



In the second example the strokes in the middle formula must not be joined up. In crystalline sodium chloride, each atom is bound to six neighbours, and the value of n is $1/6$.

2.7. Periodic Table

From the valencies and other chemical properties of the elements, the Periodic Table can now be derived. This can be done by writing these properties on cards, and arranging the cards in order of atomic mass (cf. Johnson, 1984). With few exceptions, a periodic repetition of properties is observed, especially of valency. The exceptions can be relieved by reversing the order of some of the pairs (e.g., potassium, atomic mass 39.098 u, and argon, 39.948 u). This derivation is better than one based on electronic configuration, an approach with which there are serious problems (Nelson, 1988). On the best form of table, see Jensen (1986) and Nelson (1987, 1989).

3. STEPS LEADING TO LEVEL THREE

To get from levels one and two to level three (electrons and nuclei) requires a consideration of the electrical aspects of chemistry. The following is a way of presenting these, leading to level three. The treatment presupposes that students are receiving a parallel course in physics, also taught progressively (the study of electricity beginning with the charges produced by rubbing some substances together, and the flow of these charges along a metal wire). If students are not receiving instruction in physics, the teacher will have to provide the necessary background as the course proceeds.

3.1. Electrolytes

The teacher begins by showing students that substances can be broadly divided into three types: those that do not conduct electricity (insulators), those that conduct electricity without chemical change (ordinary conductors), and those conducting electricity with chemical change at the electrodes (electrolytes). The teacher goes on to show students some simple electrochemical cells in which passage of an electric current produces chemical change (electrolysis) or chemical change produces an electric current (galvanism). This is best done after students have become thoroughly familiar with chemical changes (Section 2.1). The teacher describes what takes place entirely at a bulk level, e.g. "electrolysis of concentrated hydrochloric acid generates hydrogen at the negative electrode and chlorine at the positive electrode".

The subject can be left at this point, until after the development of the atomic and molecular theory (Sections 2.3 and 2.4). This gives students time to assimilate what are at first sight very puzzling phenomena.

3.2. Ions

In Section 2.5 we saw that molecular substances dissolve in other molecular substances, lowering their freezing point. In many cases, the lowering (θ_f) is proportional, at low concentrations, to the fraction of solute molecules (x_{mol}) in the solution [equation (1)]. This holds for solutions of nonelectrolytes in water (see, e.g., Webb & Lindsley, 1934). These give $k_f(\text{water}) = 103.1 \text{ }^\circ\text{C}$. For solutions of electrolytes, however, the picture is different.

Consider, for example, hydrochloric acid solutions. The freezing points of these at low concentrations (Randall & Vanselow, 1924) are higher than the values given by equation (1) with $k_f = 103.1 \text{ }^\circ\text{C}$. This is if the values of x_{mol} are calculated on the assumption that the solutions contain HCl molecules (call these values x_{HCl}). Thus if equation (1) is used to

calculate the number of solute particles in the solution from the observed values of θ_f and $k_f = 103.1$ °C, the values obtained (call them x_f) are greater than the values of x_{HCl} , as shown in Table 2. The limiting value of x_f/x_{HCl} at low concentrations is almost exactly two. This implies that there are actually two particles in the solution for every supposed HCl molecule.

TABLE 2: Fraction of solute molecules in dilute hydrochloric acid solutions.*

x_{HCl}	x_f	x_f/x_{HCl}
0.00002	0.0000396	1.98
0.00005	0.0000985	1.97
0.0001	0.0001956	1.96
0.0002	0.0003877	1.94

* Calculated from the data of Randall & Vanselow (1924).

Now what two particles can be derived from an HCl molecule in water? Whatever they are, they must be responsible for the electrical conductivity of hydrochloric acid solutions, as these increase with concentration (Randall & Vanselow, 1924). The particles must therefore be charged, and the only possibility is that they are charged hydrogen and chlorine atoms (“ions”), combined in some way with H_2O molecules.

The charges must be equal and opposite because the solutions are electrically neutral. Further, since, in electrolysis, chlorine is liberated at the positive electrode, chloride ions must take positive electricity from (or give negative electricity to) this electrode, and therefore must carry a negative charge. Hydrochloric acid solutions thus contain H^{q+} and Cl^{q-} ions, combined with H_2O molecules.

Similar conclusions can be reached for other electrolyte solutions. Some, however, have freezing points implying the presence of an equilibrium mixture of undissociated molecules and ions in the solution, e.g. acetic acid (Lewis et al., 1961) and copper sulfate (Brown & Prue, 1955). The latter is only about 85% dissociated at 0.001 M, 50 % at 0.01 M, and 20 % at 0.1 M. Many salts are incompletely dissociated in solution (Hawkes, 1996).

Teachers should encourage students at an early stage to think that interactions between ions and water molecules are considerable (cf. Section 2.5). The energy required to dissociate a molecule like HCl into ions has to be more than made up by the energy of these interactions. If students appreciate this they will then be prepared to learn later that hydrogen ions in water are H_3O^+ etc.

3.3. Ionic charges

The next step is to present the basic relations of electrolysis. This is most easily done after the concept of valency has been introduced (Section 2.6). The relations can then be stated as follows:

Relation 1: The mass of any substance produced or consumed in electrolysis is proportional to the quantity of electricity passed.

Relation 2: The masses of different substances produced or consumed by the same quantity of electricity are generally proportional to the masses of the atoms or radicals of which they are composed, divided by their valencies.

These relations are based on Faraday's laws. Relation 2 cannot be called a law because it does not hold for mixed conductors (Nelson, 1994a, 1997c).

Relations 1 and 2 can be combined to give, for the mass of a substance (m_s) produced or consumed by the passage of a quantity of electricity (Q):

$$m_s = kQm_x/v \quad (2)$$

where m_x is the mass of the atom or radical comprising the substance, v is its valency, and k is a constant. Equation (2) enables the charge carried by each atom or radical to be calculated. This is given by:

$$|q| = Q/N \quad (3)$$

where N is the number of atoms or radicals in the quantity of substance produced or consumed. Since the latter is given by m_s/m_x , equations (2) and (3) give

$$|q| = v/k = ve \quad (4)$$

The constant e is called the elementary charge. The ions in hydrochloric acid solutions are thus H^{+e} and Cl^{-e} , or H^+ and Cl^- .

The value of e can be determined from experimental data by using equation (2) with $k = 1/e$. This can be done without knowing the absolute value of m_x , by measuring this in atomic mass units, u (Section 2.4). The value of e is then 96500 C u/g (cf. Glasstone, 1940; Mills et al., 1993). Once students know how to determine the Avogadro number (g/u), e can be given the value 1.602×10^{-19} C, but the treatment that follows does not require this.

3.4. The electron

Once students have grasped the ionic theory, the teacher can move on to establish the existence of the electron. Students may have already met this particle in physics, but will nevertheless gain by seeing its existence established through chemistry.

The teacher begins by showing students what happens when electricity is passed through a gas at low pressures (equipment for doing this is available commercially). At moderately low pressures there is a glow, the colour of which is characteristic of the gas. If the cathode is perforated, rays of the same colour emerge through the holes ("canal rays"). At lower pressures, the tube becomes dark, and if there is a hole in the anode, the glass at the anode end of the tube glows, as if rays are being emitted from the cathode ("cathode rays").

The nature of the two kinds of ray can be established by passing them through an electric field produced by oppositely charged plates. Canal rays are deflected towards the negative plate and are positively charged; cathode rays are deflected towards the positive plate and are negatively charged.

If the rays comprise charged particles, there are several ways of determining the ratio of the mass (m) to the charge (q) of these particles. A simple one is to pass the rays through a uniform electric field (E), measure the radius (r) of the circular path they take, then use a uniform magnetic field (B) at right angles to the electric field to annul the deflection of the rays. Under the first conditions, the electric force on each particle (Eq) is equal to the particle's rate of change of momentum (mv^2/r); under the second, the electric force is balanced by the magnetic force (Bvq). From these relations

$$m/q = B^2r/E \quad (5)$$

TABLE 3: Values of m/q for some electric rays.*

Ray	m/q	Particle Identity	m
(a)			
Hydrogen ions in electrolysis	$\approx 1.044 \times 10^{-5}$ g/C	$\text{H}^+(\text{aq})$	1.007 u
Cathode rays in various gases	-5.69×10^{-9} g/C	e^-	5.49×10^{-4} u
Positive rays in hydrogen	1.044×10^{-5} g/C	H^+	1.007 u
	2.088×10^{-5} g/C	H_2^+	2.015 u
(b)			
Positive rays in neon	10.36×10^{-5} g/C	$^{20}\text{Ne}^{2+}$	19.99 u
	10.88×10^{-5} g/C	$^{21}\text{Ne}^{2+}$	20.99 u
	11.40×10^{-5} g/C	$^{22}\text{Ne}^{2+}$	21.99 u
	20.72×10^{-5} g/C	$^{20}\text{Ne}^+$	19.99 u
	21.76×10^{-5} g/C	$^{21}\text{Ne}^+$	20.99 u
	22.79×10^{-5} g/C	$^{22}\text{Ne}^+$	21.99 u
(c)			
Positive rays in helium	2.074×10^{-5} g/C	$^4\text{He}^{2+}$	4.002 u
	4.148×10^{-5} g/C	$^4\text{He}^+$	4.002 u
α -rays	2.074×10^{-5} g/C	$^4\text{He}^{2+}$	4.002 u
β -rays	-5.7×10^{-9} g/C	e^-	5.5×10^{-4} u

* Based on the data of Glasstone (1940) and Mills et al. (1993).

Values of m/q for cathode rays are the same for all gases, whereas those for the positive rays depend on the gas. This suggests that the particles in cathode rays are a common component of all substances. Some values of m/q are given in Table 3, and compared with the value for hydrogen ions in electrolyte solutions. For the latter, $q = +e = +96500$ C u/g and $m \approx m_{\text{at}} = 1.0079$ u (Section 2.4). The value of m/q for cathode rays is only a small fraction ($\approx 1/1835$) of the value for H^+ ions in electrolyte solutions, whereas the value for one of the positive rays from hydrogen is almost exactly the same (Table 3a). This suggests that the latter comprise free H^+ ions, and the former a relatively light particle of charge $-e$, a charge confirmed by other methods. The value of m/q for the second positive ray in hydrogen suggests that this comprises H_2^+ ions. Substitution of $q = \pm e$ gives the values of m in the table. [The value for $\text{H}^+(\text{aq})$ has been corrected for the removed electron.]

3.5. Isotopes

The treatment from this point is straightforward. Values of m/q for positive rays establish the existence of isotopes. A simple example is neon, which gives a set of rays corresponding to $^{20}\text{Ne}^+$, $^{21}\text{Ne}^+$, and $^{22}\text{Ne}^+$ ions, and a weaker set corresponding to $^{20}\text{Ne}^{2+}$, $^{21}\text{Ne}^{2+}$, and $^{22}\text{Ne}^{2+}$ (Table 3b). The intensities within each set are respectively 90.5%, 0.3%, and 9.2% of the total, from which the mean atomic mass is 20.18 u, the same value as given by the density of the gas (Section 2.3).

The existence of isotopes necessitates refinement of Table 3a: natural hydrogen contains a small fraction (0.015%) of ^2H . The existence of isotopes also requires a revision of the provisional statements made in Sections 2.2 and 2.3 (for the details, see Nelson, 1994b).

The modern definition of the atomic mass unit can now be given (cf. Section 2.4): $u = m_{\text{at}}(^{12}\text{C})/12$. Defining u in terms of one isotope allows accurate atomic masses to be determined.

TABLE 4: Nuclear charges determined by scattering of α -particles compared with atomic number.*

Element	q_n/e	Atomic Number
Copper	29.3 ± 0.5	29
Silver	46.3 ± 1.0	47
Platinum	77.4 ± 1.0	78

* Source: Chadwick (1920).

3.6. The nucleus and atomic structure

Values of m/q for α - and β -rays from radioactive substances, together with the observation that α -rays generate helium, enable the particles in these rays to be identified (Table 3c). Back scattering of α - or β -particles from metal targets (which can be demonstrated with commercially available equipment) leads to the nuclear structure of the atom. From the fraction of α - or β -particles scattered through known angles by metal foils, nuclear charges (q_n) can be calculated (Rutherford, 1911; Chadwick, 1920). Values obtained in this way (Table 4) support the proposition that q_n/e is equal to atomic number (Z), the ordinal number of an element in a chemically derived Periodic Table. An atom thus has Z electrons. The general arrangement of these electrons can be inferred from the valencies of the elements (Nelson, 1988).

4. CONCLUDING COMMENTS

In this article I have shown how chemistry at levels two and three can be derived from observations at level one. The programme I have described is not meant to be a complete course in introductory chemistry. There is more to chemistry than theory (Nelson, 1983). Students need to appreciate that chemistry is a practical subject, with wide applications. But the relationship between levels one, two, and three is an important part of chemistry, and needs to be presented well.

Feedback from students on the courses mentioned in the introduction indicate that the students on these courses like chemistry taught progressively (see, e.g., Hutchinson, 2000). These courses are run at a variety of institutions, and are taken by a range of students. I hope therefore that teachers will consider teaching chemistry in this way, and that examiners and publishers will support them.

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