COURSE CODE: CHM 101

COURSE TITLE: INTRODUCTORY PHYSICAL CHEMISTRY

NO OF UNITS: 03

COURSE DURATION: THREE HOURS PER WEEK

COURSE DETAILS:

PART TITLE: ATOMS, MOLECULES AND STRUCTUREELECTRONIC SPECTRA AND PERIODIC TABLE

COURSE LECTURER: DR. S. ADEWUYI

adsewuyis@unaab.edu.ng

COURSE REQUIREMENTS:

Students are expected to have a minimum of 75% attendance in this course before they could be allowed to write the examination

READING LIST:

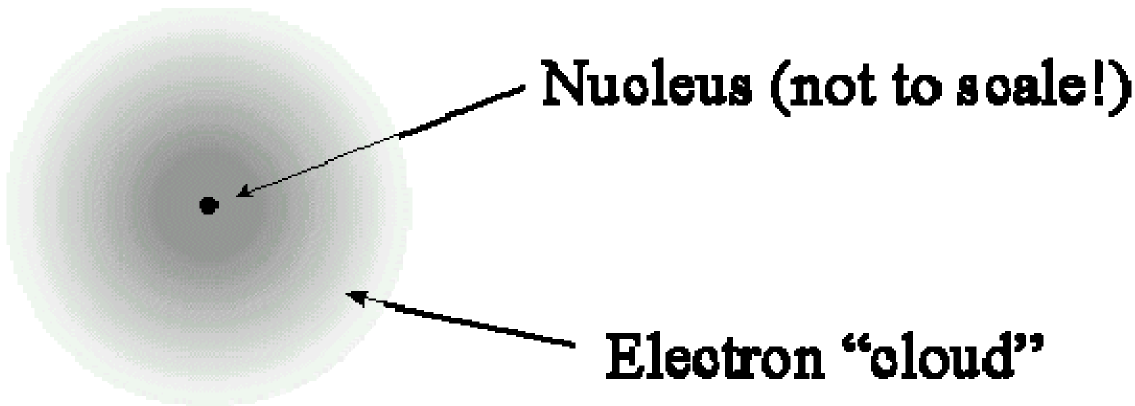
1. **Gross, J.M. and Wiseall, B.** ***Principle of physical chemistry*****. MacDonald and Evans Handbook series, 1972**
2. **Atkins, P.W.** ***Physical chemistry.*** **Oxford University Press, sixth edition, 1999**
3. **Bahl, A and Bahl, B.S.** ***Essentials of physical chemistry,*** **S.Chand and Company Ltd. 2007**

4. **Brown, T.L., Lemay, H.E., Bursten, B.E and Murphy, C.J *Chemistry: The central science.* Pearson Education, 11th Edition 2009.**
5. **Sharma, K.K. and Sharma, L.K. *Physical chemistry***

LECTURE NOTES

WHAT ARE THE COMPONENTS OF AN ATOM?

Atoms are composed of elementary particles:



Protons - "heavy," positive electric charge (mass = 1.6726 × 0^{-24}g)
neutrons - "heavy," electrically neutral (mass = 1.6749 × 10^{-24}g)
electrons - "light," negative electrical charge (mass = 9.1094 × 10^{-28}g)

It is easier to measure relative masses that to measure absolute masses so we introduce the atomic mass unit (amu) as 1/12 the mass of the carbon atom which has 6 protons and 6 neutrons in its nucleus. In amu the masses of the elementary particles are:
Protons – 1.007276 amu
neutrons – 1.008665 amu
electrons – 0.0005485799 amu
Electron and proton charges are equal and opposite - they "balance" each other.
Neutrons have zero charge.

In a neutral atom, # electrons = # protons = atomic number

Protons + # neutrons = mass number (or atomic mass number)

Mass of proton ≈ mass of neutron ≈ 1836 × mass of electron

(Most of the mass of an atom is in the protons and neutrons.)

The protons and neutrons are bound tightly together to form the nucleus. Most of the mass of the atom is concentrated in the nucleus.

The electrons surround the nucleus in a "charge cloud." Since the electrons are 1836 times lighter than protons and neutrons, only about 0.03% of the mass is in the electrons.

On the other hand, the nucleus is very small, most of the size (volume) of the atom is provided by the electrons.

Examples:

Hydrogen: The diameter of the nucleus (proton) is about 0.01 pm, but the diameter of the atom is about 104 pm.

Note: The mass of the electron is 9.1×10^{-28} g while the mass of the H atom is about 1.7×10^{-24} g.

Lead: The diameter of the nucleus is about 0.067 pm, but the diameter of the atom is about 350 pm.

Since the electrons surround the nucleus (they are on the "outside,") bonding must depend on what the electrons are doing.

(The electrons "shield" the nucleus – the electrons "bump into" stuff before the nucleus does.)

(There were other models of the atom, e.g. the "raisin muffin" model, but they didn't stand up to experimental test.)

We will look at what the electrons are doing (the electronic structure of the atom) later.

Structure of the atomic nucleus

The atomic number = # protons (always) = # electrons (in a neutral atom) determines what an element is.

O is atomic number 8
H is atomic number 1
He is atomic number 2
C is atomic number 6
Cl is atomic number 17
Ar is atomic number 18
etc.

Nuclear structure

Recall:

Mass Number = # neutrons + # protons
Atomic Number = # protons

We write a symbol of the atom which contains all of the nuclear and charge information as follows:

$${}^{MN}_{AN}X^{chg}$$

X is the symbol of the element and the subscript and superscripts tell us about the nuclear structure.
MN = mass number (Sometimes this is called the atomic mass number.)

AN = atomic number

chg = charge

Examples:

1 proton + 0 neutron + 1 electron = ${}^{1}_{1}H$

2 proton + 2 neutron + 1 electron = ${}^{4}_{2}He^{+}$

6 proton + 6 neutron + 7 electron = ${}^{12}_{6}C^{-}$

82 proton + 125 neutron + 80 electron = ${}^{207}_{82}Pb^{2+}$

and so on.

Isotopes

Although atoms of a given element always have the same number of protons, they can have different numbers of neutrons.

(Notice that this will not change the chemical properties, because the number of electrons will be the same.)

Atoms of an element with different mass numbers are called isotopes.

We use the same type symbols to specify the different isotopes:

Examples:

Isotopes of hydrogen - ${}^{1}_{1}H$, ${}^{2}_{1}H$, ${}^{3}_{1}H$
(${}^{2}_{1}H$ is called "deuterium" and ${}^{3}_{1}H$ is called "tritium." Not all isotopes have their own special name.)

Isotopes of helium - ${}^{3}_{2}He$, ${}^{4}_{2}He$
(${}^{3}_{2}He$ is called "helium – 3," and ${}^{4}_{2}He$ is called "helium – 4," and so on for the rest of the elements.)

Isotopes of carbon - ${}^{12}_{6}C$, ${}^{13}_{6}C$, ${}^{14}_{6}C$
(Carbon – 12, carbon – 13, and Carbon 14,etc.)
There are 7 known isotopes of C with mass numbers ranging from 10 to 16

Isotopes of oxygen - ${}^{16}_{8}O$, ${}^{17}_{8}O$, ${}^{18}_{8}O$
(There are 8 isotopes of oxygen: 13 to 20.)

and so on.

Notice that the symbol of the element and the atomic number subscript are redundant. We leave the subscript on when we are writing nuclear reactions, but sometimes the atomic number is left off.
For example, ${}^{235}_{92}U$ can also be written ${}^{235}U$.
Let's consider the nuclear structure of ${}^{235}_{92}U$.
This symbol tells us that the atom of uranium - 235 contains 92 protons and 235 - 92 = 143 neutrons. Since there is no charge indicated, we conclude that there are also 92 electrons in the atom.

Why are atomic weights or the atomic masses in the periodic table or in our element lists not integers?

Two reasons:

1. The neutrons and protons in different elements (and even different isotopes of the same element) do not all weigh the same and they do not weigh the same as free neutrons (1.008665 amu) and free protons (1.007276 amu).

These differences are small, but it means that we can't find the mass of a given isotope by adding the masses of all the neutrons and protons.

2. The main reason is that the atomic weight that we measure is the weighted average of the naturally occurring isotope masses.

Sometimes we call the weighted average of the naturally occuring isotopes the chemical atomic weight.

(With a mass spectrometer chemists can determine the natural abundance of the isotopes of an element and the actual mass of each isotope. The natural abundance is the percentage of each isotope in a sample of the naturally occurring element.)

Examples:

$^{12}_{6}C$ - 98.89%, mass = 12.0000 amu (by definition)

$^{13}_{6}C$ - 1.11%, mass = 13.00335 amu

The weighted average is:

$$\frac{98.89}{100}12.0000 + \frac{1.11}{100}13.00335 = 12.01 \text{ amu}$$

Another example:

$^{16}_{8}O$ - 99.759%, mass = 15.99491 amu

$^{17}_{8}O$ - 0.037%, mass = 16.995 amu

$^{18}_{8}O$ - 0.204%, mass = 17.9943 amu

The weighted average is:

$$\frac{99.759}{100}15.99494 + \frac{0.037}{100}16.995 + \frac{0.204}{100}17.9943 = 15.9994$$

One more example:

^{35}Cl - 75.53%, mass = 34.96885 amu

^{37}Cl - 24.47%, mass = 36.9671 amu

Weighted average is:

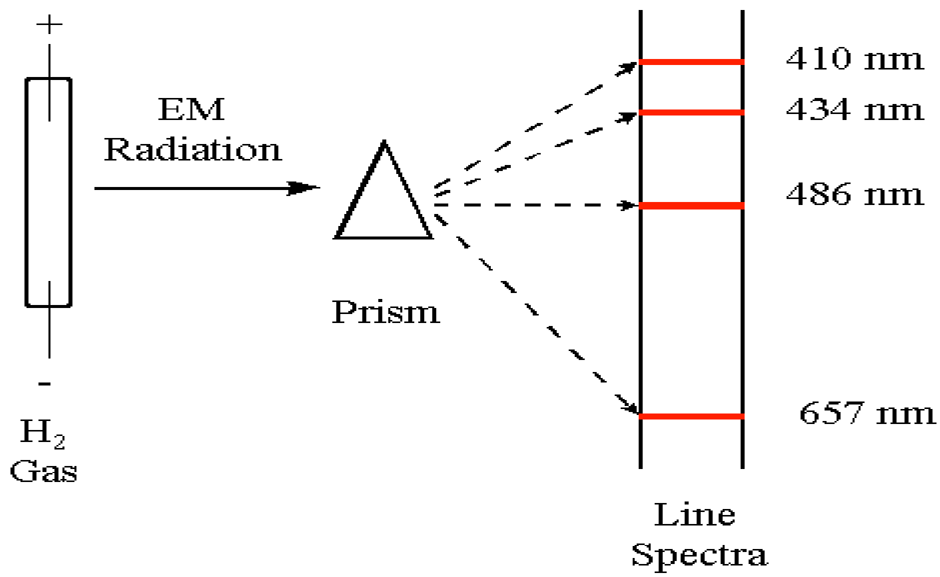
$$\frac{75.53}{100}34.96885+\frac{24.47}{100}36.94617=35.4527$$

The weighted average atomic masses are all in units of amu.

LINE SPECTRA

The emissions from hot gasses like H, and He was diffracted through a prism to give a series of lines of different colors instead of a smooth distribution of frequencies. They came to be known as line spectra.

Although there was no explanation for why this should be, people were able to find a formula which fit the wavelengths of light emitted by atomic hydrogen. The wavelengths fit into groups called "series."



One such group called the **Balmer series** fit an equation, called the Rydberg equation, of the form

$$\frac{1}{\lambda}=R\left(\frac{1}{2^2}-\frac{1}{n^2}\right),$$

Where n = 3, 4, 5…

The constant, R, is called the "Rydberg constant." It has the value, $1.0974 \times 10^{7}\ \text{m}^{-1}$ (or $0.010974\ \text{nm}^{-1}$).

People suspected that some of the other series might be obtained by changing the 2^2 to, maybe, 1^2, or 3^2, and so on. That is, maybe there were series whose wavelengths fit the equation

$$\frac{1}{\lambda} = R\left(\frac{1}{1^2} - \frac{1}{n^2}\right),$$ **Lyman series**

or $$\frac{1}{\lambda} = R\left(\frac{1}{3^2} - \frac{1}{n^2}\right),$$ **Paschen series**

and so on.

When people looked for spectral lines with wavelengths given by these formulas the found them!

The series from the first formula is called the Lyman series after the physicist who found it, and the series from the second formula is called the Paschen series after its discoverer. There were even series found with the 2^2 replaced with 4^2 and 5^2.

THE BOHR MODEL OF THE HYDROGEN ATOM

In 1913 Niels Bohr formulated a successful theory of the hydrogen atom. By his time

it was known that atoms consisted of a central heavy and relatively small core - called the nucleus - surrounded by the relatively light electrons.

The question is, how are the electrons arranged?

Bohr assumed that the electrons move around the nucleus like the planets move around the sun.

But the physics of Newton and Maxwell predicted that such an atom would radiate away its energy and collapse.

Bohr assumed that for unknown reasons the atom wouldn't radiate away its energy and collapse, but that there were only certain allowed orbits.

He assumed that Planck's formula for the energy of a photon was correct and then converted the Rydberg equation for wavelength into an equation in terms of energy rather than wavelength.

Remember that $\nu = c/\lambda$ and that Planck's equations gives the energy of a photon as $E = h\nu$, so $E = hc/\lambda$.

Multiply the Rydberg formula by hc and we get an equation that has units of energy.

$$\Delta E = \frac{hc}{\lambda} = hcR\left(\frac{1}{2^2} - \frac{1}{n^2}\right)$$

Bohr guessed that the energy of the photon emitted by an atom came from the electron in an atom dropping from a higher energy orbit to a lower energy orbit.

Based on his modified Rydberg formula, he deduced that the energy of the allowed orbits had to fit a formula of the form

$$E_m = -hcR\frac{1}{m^2},$$

where m can be an integer equal to 1, 2, 3, 4, and so on up to infinity. Then the energy of the photons in the Balmer series of spectral lines must come from an electron dropping from an orbit with m = 3, 4, 5, … down to the orbit with m = 2.

The New Quantum Theory

The above description of matter is sometimes now referred to as the "old quantum mechanics." The old quantum mechanics did very well in some cases but was not easily extendable to more complicated systems. It worked well for some aspects of the hydrogen atom, like line spectra, but failed in some areas. It couldn't be extended to more complicated atoms, like helium. Also, it did not provide a consistent explanation for Planck's and Einstein's applications of the theory.

In 1926 the Austrian physicist, Erwin Schrödinger, was pursuing de Broglie's idea that matter shared some of the properties of waves. In classical physics wave motion was well known and well studied. All wave motion could be described by an appropriate "wave equation."

The wave equation that describes the wave nature of particles is called Schrödinger's equation and it looks different from the classical wave equation.

Schrödinger's equation replaces the amplitude of the classical wave equation with a "wave function," usually symbolized by the Greek letter psi, ψ, or Ψ.

The wave function provides a description of the position of a particle in terms of probabilities.

The square of the magnitude of ψ tells us where the particle is. At a point in space where ψ is large the particle is likely to be there. At a point in space where ψ is small the particle is not likely to be there. And where ever in space ψ is zero the particle isn't there.

Schrödinger's equation does more than just tell us where the particle is likely to be. It also gives us the allowed energies of the system. There is an energy associated with each ψ_i which we indicate by E_i.

Schrödinger's quantum theory says that the electron in a hydrogen atom has three quantum numbers,

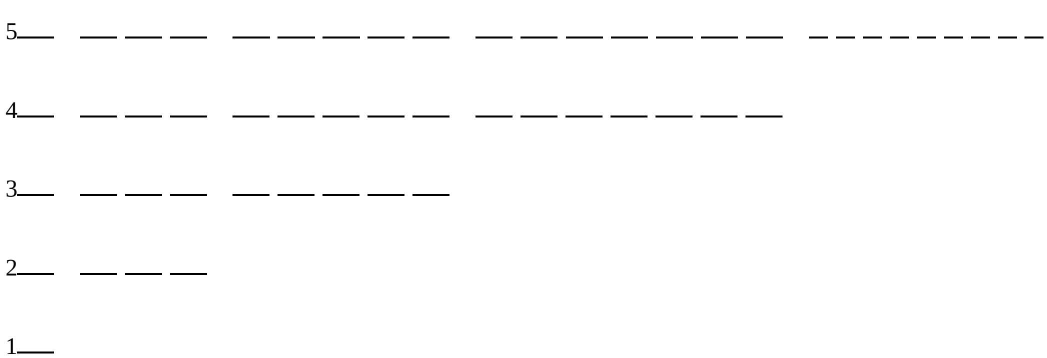
n, l, and m_l.

The allowed values of these quantum numbers are:
n (Principal quantum number) = 1, 2, 3, 4 … ∞,

l (Subsidiary quantum number) = 0, 1, 3 … $n - 1$

m_l (Magnetic quantum number) = $- l \leq 0 \leq l$.

In talking about the electronic structures of atoms it is usual to use the letters, *s, p, d, f, g*, … to refer to electrons with l = 0, 1, 2, 3, 4, … and so on. An energy level diagram for the hydrogen looks like the following diagram. Remember that on these diagrams energy is up. Notice that there is one lowest state - called the ground state. All other states are "excited states" of hydrogen.



s *p* *d* *f* *g*

In the hydrogen atom all the states with the same value of *n* have the same energy. We call this "degeneracy." That is, states with the same energy are degenerate.

Most of the spectroscopy of hydrogen can be explained using the above energy level diagram.

Each of the states corresponds to an atomic orbital. For hydrogen the orbitals are the wave function solutions of Schrödinger's equation.

We can think of the orbital as probability distribution clouds. Where the cloud is thick the electron is likely to be, where the cloud is thin the electron is not likely to be, and where there is no cloud there is no electron.

An electron can reside in any one of the orbitals. If the electron is in the orbital with quantum number, $n = 1$, we say that hydrogen is in its "ground state." If the electron is in an orbital with $n > 1$ we say that the hydrogen atom is in an excited state.

Atoms in excited states can drop down to a lower state or to the ground state by emitting a photon with the appropriate amount if energy.

It is useful to think of the orbitals as "places" where an electron can reside. So we can say things like, "the electron is in the 1*s* orbital, or "the electron is in the $3d_z{}^2$ orbital," and so on.

The electron in a hydrogen atom can be in any one of the orbitals, but if you leave the atom alone long enough the electron will drop down to the ground state and emit the excess energy as a photon. (It may drop down to the ground state in several steps.)

An input of energy is required to place the hydrogen atom in one of the excited states. This energy can come from the absorption of a photon (of the appropriate energy) or from heating the hydrogen gas to a high temperature.

An electron in an excited state will always drop down to the ground state by the emission of a photon (although it drops faster from some excited states than from others).

Electron Spin

The electron has another quantum number which could not be derived from Schrödinger's theory. This is the quantum which describes electron spin. Electron spin is a relativistic effect and could not be derived from Schrödinger's nonrelativistic theory.

Electron spin had been predicted on the basis of experimental work, but it was not understood theoretically until the relativistic theory of Paul Dirac.

The spin quantum number is labeled m_s and can take on only the values +1/2 and – 1/2.

The phenomenon is called spin because the electron acts like a tiny magnet and the only way a charged body can act like a magnet is if it is spinning. (The electron is probably not a small spinning body.)

So, to describe an electron in a hydrogen atom completely we need four quantum numbers:

n, l, m_l, and m_s.

There are times when we might want to be explicit about referring to the electron being in a state with $m_s = 1/2$ or $m_s = -1/2$, but it is more common to refer to $m_s = 1/2$ as "spin up" and $m_s = -1/2$ as "spin down." Further, in energy level diagrams we can indicate "spin up" by a vertical arrow pointing up, , and "spin down" by a vertical arrow pointing down

Atoms with More Than One Electron

Up to now we have been dealing with only one electron. Now we have to think about atoms with more than one electron.

The simplest thing to do is to assume that electrons added to the atom go into the same type orbitals we have been talking about. It turns out that this assumption is about 98% correct.

The question then becomes how do the electrons fill the orbitals? The answer to this question has been worked out over the years by a combination of experiment and theory. We now have a set of rules to describe filling the set of orbitals we use.

Rule 0: Atom has the same set of orbitals, 1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 3*d*, … etc., that are available in the hydrogen atom and the hydrogen-like ions.

However, the relative energies of the individual orbitals is not the same as in hydrogen. Some of the degeneracy has been lifted. The *p* orbitals are still degenerate among themselves, and the *d* orbitals are degenerate among themselves, and so on, but the 2*s* and 2*p* orbitals for, example, are no longer degenerate, and so on. The reason for this is that the electrons "screen" the nucleus from additional electrons. Each added electron sees not only the nucleus, but the nucleus and all the other electrons that have been added. The order of the energy of the electrons, starting from the lowest energy orbital, the 1*s*, is

1*s*, 2*s*, 2*p*, 3*s*, 3*p*, 4*s*, 3*d*, 4*p*, 5*s*, 4*d*, 5*p*, 6*s*, 4*f*, 5*d*, 6*p*, 7*s*, 5*f*, 6*d*, 7*p*, 8*s*, 5*g*,

and so on. (There are more than enough orbitals to take care of the elements that exist.)] Without worrying about the energy scale, the diagram looks like:

8*s*—

7*p*— — —

6*d*— — — — —

5*f*— — — — — — —

7*s*—

6*p*— — —

5*d*— — — — —

4*f*— — — — — — —

6*s*—

5*p*— — —

4*d*— — — — —

5*s*—

4*p*— — —

3*d*— — — — —

4*s*—

3*p*— — —

3*s*—

2*p*— — —

2*s*—

1*s*—

s *p* *d* *f* *g*

Rule 1: The "Aufbau" Principle (German for "build up."

The Aufbau principle says that we fill the lowest energy orbitals first (subject to the other rules to come).

Rule 2: The Pauli Exclusion Principle.

There can be no more than two electrons in each orbital. If there are two electrons they must have opposite spins. (That is, one must be spin up and the other must be spin down.) When there are two electrons in an orbital with opposite spins we say that the spins are "paired."

Rule 3: Hund's Rule

When filling degenerate orbitals keep the spins unpaired as long as possible.

Rule 4: Filled shells are particularly stable, but half filled shells also have a little extra stability.

By "shells" we mean all of a group of *p* orbitals at a particular level (with the same principal quantum number) or all of a group of *d* orbitals at a particular level.

Here are some examples:

O $1s^2 2s^2 2p^4$

Si $1s^2 2s^2 2p^6 3s^2 3p^2$

Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Cr $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

Br $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$

La $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 4f^1 5s^2 5p^6 6s^2$

OVERVIEW OF THE PERIODIC TABLE

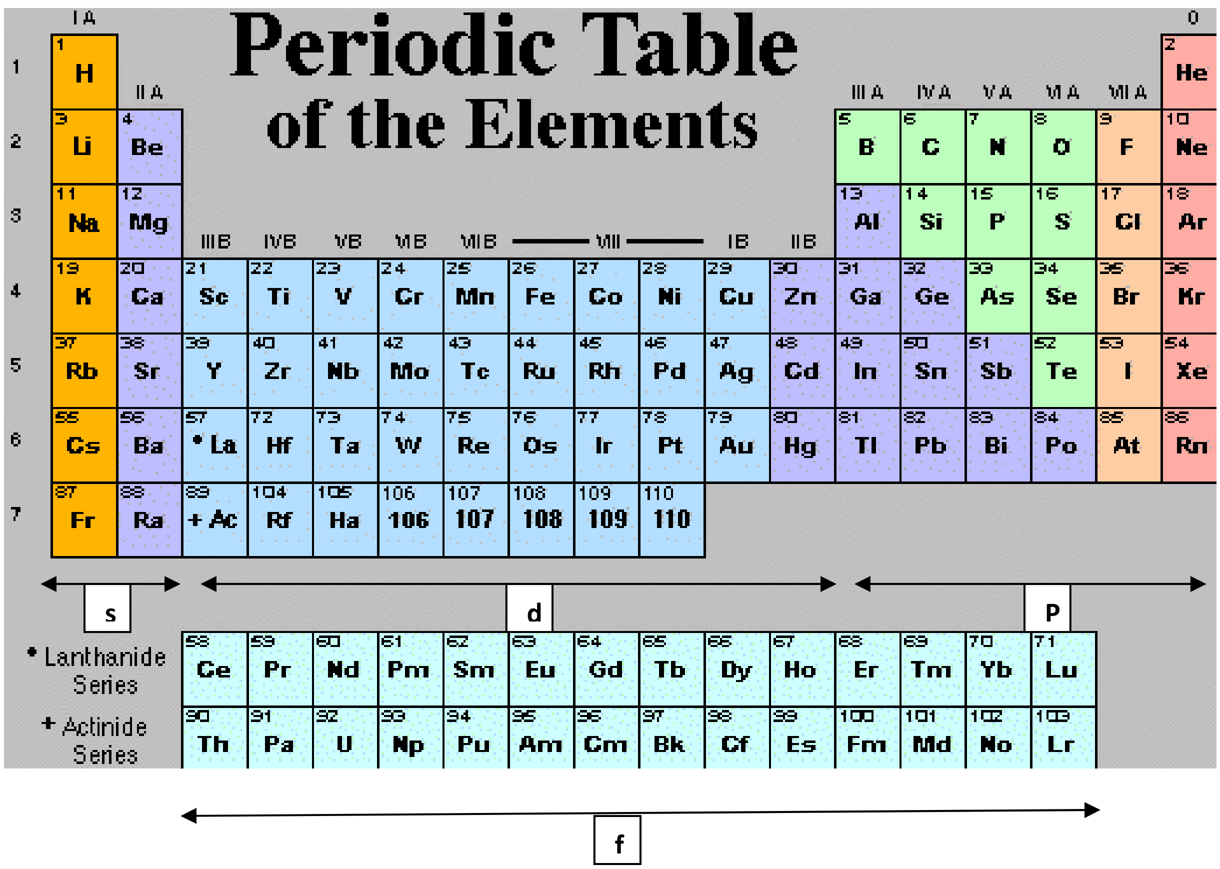
As more and more elements were discovered and characterized, efforts were made to see whether they could be **grouped**, or classified, according to their chemical behavior. This effort resulted, in 1869, in the development of the **Periodic Table**.

Certain elements show similar characteristics:

- Lithium (Li), Sodium (Na) and Potassium (K) are all soft, very reactive metals
- Helium (He), Neon (Ne) and Argon (Ar) are very non-reactive gasses

If the elements are arranged in order of increasing atomic number, their chemical and physical properties are found to show a repeating, or periodic pattern.

Note: This table lists the atomic number (number of protons) in the upper left corner of each box. The atomic number is formally placed as a subscript preceding the atom name.



As an example of the periodic nature of the atoms (when arranged by atomic number), each of the soft reactive metals comes immediately after one of the nonreactive gasses.

The elements in a column of the periodic table are known as a **family** or **group**. The labeling of the families are somewhat arbitrary, but are usually divided into the general groups of:

- Metals (everything on the left and middle region)
- Non-metals (upper diagonal on the right hand side - green, salmon and red)

- Metaloids (atoms in the boundary between the metals and metaloids: Boron(B), Silicon(Si), Germainium(Ge), Arsenic(As), Antimony(Sb), Tellurium(Te), Astatine(At)). These are some of the more useful materials for semi-conductors.

or, another convention is the 'A' and 'B' designators with column number labels (either in Roman or Arabic numerals). These columns have different types of classifications:

Group	Name	Elements
1A	Alkali metals	Li, Na, K, Rb, Cs, Fr
2A	Alkaline earth metals	Be, Mg, Ca, Sr, Ba, Ra
6A	Chalcogens ("chalk formers")	O, S, Se, Te, Po
7A	Halogens ("salt formers")	F, Cl, Br, I, At
8A	Noble gases (or inert, or rare gases)	He, Ne, Ar, Kr, Xe, Rn

The elements in a family of the periodic table have similar properties because they have the same type of **arrangement of electrons** at the **periphery** of their atoms.

The majority of elements are metals:

- high luster
- high electrical conductivity
- high heat conductivity
- solid at room temperature (except Mercury [Hg])

Note: hydrogen is a non-metal (at left hand side of the periodic table)

Non-metals

- solid, liquid or gas at room temp

General principles:

Elements in the same column have similar chemical properties.

Reactivity gets more varied as you go down a column.

Alkali metals
The elements in the far left column are called alkali metals. They are very reactive metals and tend to form ions with a charge of +1 in compounds.
Alkaline earth metals
The elements in the second from the left column are called alkaline earth metals. They are not as reactive as the alkali metals and tend to form ions with a charge of +2 in compounds.

Transition elements
The next ten columns comprise the transition elements. We will learn why they are called that in CHM 104. These elements are all metals.
B group

C group

For example, the elements of the boron group all form compounds with hydrogen with the formula, XH_4.
N group

O group

Halogens (acid formers)

noble gases

Tutorial Questions

Question #1: Which of the following cannot be broken down to anything simpler?

(A). water
(B). table salt
(C). silver
(D). sugar

Question #2: The first part of an atom to be discovered was the

(A). proton.
(B). neutron.
(C). electron.
(D). nucleus.

Question #3: The planetary model of an atom, with the nucleus playing the role of the Sun and the electrons playing the role of planets, is unacceptable because

(A). the electrical attraction between a proton and an electron is too weak.
(B). an electron is accelerating and would lose energy.
(C). the nuclear attraction between a proton and an electron is too strong.
(D). none of these because the planetary model is acceptable.

Question #4: Most of the volume of an atom is occupied by

(A). electrons.
(B). protons.
(C). neutrons.
(D). empty space

Question #5: The Bohr model of the atom was able to explain the Balmer series because

(A). larger orbits required electrons to have more negative energy in order to match the angular momentum.
(B). differences between the energy levels of the orbits matched the difference between energy levels of the line spectra.
(C). electons were allowed to exist only in allowed orbits and nowhere else.
(D). none of the above

Question #6: According to the equation de Broglie derived to describe matter waves, doubling the velocity of an electron would result in (A). less momentum.
(B). a greater mass.
(C). a smaller wavelength.
(D). an unchanged wavelength and mass.

Question #7: One reason the Bohr model of the atom failed was because it did not explain why

(A). accelerating electrons do not emit electromagnetic radiation.
(B). moving electrons have a greater mass.
(C). electrons in the orbits of an atom have negative energies.
(D). electrons in greater orbits of an atom have greater velocities.

Question #8: A hydrogen atom is in a ground state when its electron

(A). has moved away from the atom to the ground.
(B). has stopped moving.
(C). is moving back and forth through the nucleus.
(D). remains at the lowest energy level.

Question #9: What is the energy of a photon of black light (ultraviolet) that has a frequency of 2.00×10^{16} Hz?

(A). 1.33×10^{-17} J
(B). 13.26×10^{18} J
(C). 1.33×10^{51} J

(D). 3.05 X 10^{-19} J   Note: J = 6.63 x 10^{-34} Js

Question #10: What is the de Broglie wavelength of an electron with a velocity of 2.00 X 10^{7} m/s?

(A). 1.82 X 10^{-23} m
(B). 1.21 X 10^{-10} m
(C). 3.64 X 10^{-11} m
(D). 3.32 X 10^{-41} m   Note: Mass of electron = 9.11 x10^{-31}kg

Question #11: Compared to metals, nonmetals are

(A). more brittle as a solid.
(B). better electrical conductors.
(C). more ductile as a solid.
(D). better conductors of heat.

Question #12: Carbon cannot be broken down into anything simpler by chemical means, so carbon must be

(A). a heterogeneous mixture.
(B). a homogeneous mixture.
(C). an element.
(D). a compound.

Question #13: Isotopes are atoms of an element with identical chemical properties but with different

(A). numbers of protons.
(B). masses.
(C). numbers of electrons.
(D). atomic numbers.

Question #14: The weighted average of the masses of the stable isotopes of an element as they occur in nature is called the

(A). atomic number.
(B). atomic mass.
(C). atomic weight.
(D). mass number.

Question #15: The modern periodic law is based on

(A). atomic number.
(B). atomic mass.

(C). atomic weight.
(D). chemical activity.

Atomic Structure

Question #1: Which of the following cannot be broken down to anything simpler?

(A). water
(B). table salt
(C). silver
(D). sugar

#1 Answer (C). silver.

Water can be broken down to oxygen and hydrogen, table salt can be broken down to sodium and chlorine, and sugar can be broken down to carbon, hydrogen, and oxygen. Silver cannot be broken down to anything simpler because it is an element, as are oxygen, hydrogen, sodium, chlorine, and carbon.

Question #2: Elements combine in fixed mass ratios to form compounds. This must mean that elements
(A). are made up of continuous matter without subunits.
(B). are composed of discrete units called atoms.
(C). have unambiguous atomic numbers.
(D). are always chemically active.

#2 Answer (B). are composed of discrete units called atoms.

If matter were continuous there would be no reason for one amount to combine with another amount. Matter is made up of discrete units called atoms that combine in a fixed weight ratios.

Question #3: The first part of an atom to be discovered was the

(A). proton.
(B). neutron.
(C). electron.
(D). nucleus.

#3 Answer (C). electron.

The English physicist J. J. Thomson discovered the electron, a fundamental part of an atom, in 1897.

Question #4: The electron was discovered through experiments with

(A). electricity.
(B). light.
(C). radio waves.
(D). radioactivity.

#4 Answer (A). electricity.

The English physicist J. J. Thomson was working with a high-voltage electrical source connected to two metal plates in an evacuated glass tube when a greenish beam was observed to move from the cathode to the anode. Working with this beam between charged metal plates and a strong magnetic field, Thomson found it was made of negatively charged particles, or electrons.

Question #5: The nucleus was discovered through experiments with

(A). electricity.
(B). light.
(C). radio waves.
(D). radioactivity.

#5 Answer (D). radioactivity.

Rutherford and his co-workers studied alpha particle scattering from a thin gold foil. The alpha particles struck a detecting screen, producing a flash of visible light. Measurements of the angles between the flashes, the foil, and the radioactive source of the alpha particles showed that the particles were scattered in all directions, including straight back from the foil. These measurements gave Rutherford a means of estimating the size of the nucleus.

Question #6: The planetary model of an atom, with the nucleus playing the role of the Sun and the electrons playing the role of planets, is unacceptable because

(A). the electrical attraction between a proton and an electron is too weak.
(B). an electron is accelerating and would lose energy.
(C). the nuclear attraction between a proton and an electron is too strong.
(D). none of these because the planetary model is acceptable.

#6 Answer (B). an electron is accelerating and would lose energy.

According to understandings about the relationship between charged particles and electromagnetic radiation, an accelerating electric charge should emit electromagnetic radiation such as light. If an electron gave off light, it would lose energy. The energy loss would mean that the electron could not maintain its orbit, and would be pulled into the nucleus and the atom would collapse. Therefore the planetary model of an atom is unacceptable because atoms continue to exist.

Question #7: Most of the volume of an atom is occupied by

(A). electrons.
(B). protons.
(C). neutrons.
(D). empty space

#7 Answer (D). empty space

Rutherford was able to estimated the radius of the nucleus from his experiments with alpha particle scattering. The radius of the nucleus was found to be approximately 10^{-13} cm. Since the radius of the atom was found to be on the order of 10^{-8} cm, this means the electrons are moving around the nucleus at a distance 100,000 times the radius of the nucleus, meaning the volume of an atom is mostly empty space.

Question #8: The atomic number of an atom identifies the number of

(A). protons.
(B). neutrons.
(C). quantum orbits.
(D). excited states.

#8 Answer (A). protons.

The atomic number identifies the number of protons in the nucleus of an atom. A neutral atom also has negatively charged electrons that are equal in number to the protons.

Question #9: The Bohr model of the atom was able to explain the Balmer series because

(A). larger orbits required electrons to have more negative energy in order to match the angular momentum.
(B). differences between the energy levels of the orbits matched the difference between energy levels of the line spectra.
(C). electons were allowed to exist only in allowed orbits and nowhere else.
(D). none of the above

#9 Answer (B). differences between the energy levels of the orbits matched the difference between energy levels of the line spectra.

Question #10: The idea of matter waves, as reasoned by de Broglie, describes a wavelike behavior of any (A). particle, moving or not.
(B). particle that is moving.
(C). charged particle that is moving.
(D). particle that is stationary.

#10 Answer (B). particle that is moving.

Question #11: According to the equation de Broglie derived to describe matter waves, doubling the velocity of an electron would result in (A). less momentum.
(B). a greater mass.
(C). a smaller wavelength.
(D). an unchanged wavelength and mass.

#11 Answer (C). a smaller wavelength.

Question #12: A hydrogen atom has an electron in the sixth excited state so the principal quantum number of this electron is

(A). 7.
(B). 6.
(C). 5.
(D). 4.

#12 Answer (A). 7.

Question #13: One reason the Bohr model of the atom failed was because it did not explain why

(A). accelerating electrons do not emit electromagnetic radiation.
(B). moving electrons have a greater mass.
(C). electrons in the orbits of an atom have negative energies.
(D). electrons in greater orbits of an atom have greater velocities.

#13 Answer (A). accelerating electrons do not emit electromagnetic radiation.

Question #14: A hydrogen atom is in a ground state when its electron

(A). has moved away from the atom to the ground.
(B). has stopped moving.
(C). is moving back and forth through the nucleus.
(D). remains at the lowest energy level.

#14 Answer (D). remains at the lowest energy level.

Question #15: An atom of hydrogen emits a photon when its electron

(A). jumps from a lower-energy orbit to a higher-energy orbit.
(B). jumps from a higher-energy orbit to a lower-energy orbit.
(C). combines with a proton.
(D). combines with a neutron.

#15 Answer (B). jumps from a higher-energy orbit to a lower-energy orbit.

Question #16: What is the energy of a photon of black light (ultraviolet) that has a frequency of 2.00 X 10^{16} Hz?

(A). 1.33 X 10^{-17} J
(B). 13.26 X 10^{18} J
(C). 1.33 X 10^{51} J
(D). 3.05 X 10^{-19} J

#16 Answer (A). 1.33 X 10^{-17} J

$$
\begin{aligned}
E &= hf \\
&= \left(6.63 \times 10^{-34}\ \mathrm{Js}\right)\left(2.00 \times 10^{16}\ \frac{1}{\mathrm{s}}\right) \\
&= \left(6.63 \times 10^{-34}\right)\left(2.00 \times 10^{16}\right) \mathrm{Js} \times \frac{1}{\mathrm{s}} \\
&= 1.33 \times 10^{-17}\ \mathrm{J}
\end{aligned}
$$

Question #17: What is the de Broglie wavelength of an electron with a velocity of 2.00 X 10^{7} m/s?

(A). 1.82 X 10^{-23} m
(B). 1.21 X 10^{-10} m
(C). 3.64 X 10^{-11} m
(D). 3.32 X 10^{-41} m

#17 Answer (C). 3.64 X 10^{-11} m

$$\begin{aligned}
\lambda &= \frac{h}{mv} \\
&= \frac{6.63 \times 10^{-34}\,\mathrm{Js}}{\left(9.11 \times 10^{-31}\,\mathrm{kg}\right)\left(2.00 \times 10^{7}\,\frac{\mathrm{m}}{\mathrm{s}}\right)} \\
&= \frac{6.63 \times 10^{-34}\,\mathrm{Js}}{1.82 \times 10^{-23}\,\frac{\mathrm{kg\,m}}{\mathrm{s}}} \\
&= 3.64 \times 10^{-11}\,\frac{\frac{\mathrm{kg\,m^2}}{\mathrm{s^2}} \cdot \mathrm{s}}{\frac{\mathrm{kg\,m}}{\mathrm{s}}} \\
&= 3.64 \times 10^{-11}\,\mathrm{m}
\end{aligned}$$

Question #18: What is the frequency of a photon emitted when an electron in a hydrogen atom jumps from n = 3 to n = 2?

(A). 3.02 X 10^{-19} Hz
(B). 4.56 X 10^{14} Hz
(C). 7.29 X 10^{14} Hz
(D). 6.2 X 10^{14} Hz

#18 Answer (B). 4.56 X 10^{14} Hz

Question #19: What is the electron configuration for potassium (atomic number 19)?

(A). $1s^2 2p^6 3s^4 3p^6 4s^2$
(B). $1s^4 2p^6 3s^2 3p^6 4s^2$
(C). $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
(D).$1s^2 2s^2 2p^6 3s^2 3p^8$

#19 Answer (C). $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

First, note that an atomic number of 19 means a total of nineteen electrons. Second, refer to the order of filling matrix in Figure 9.21. Remember that only two electrons can occupy an orbital, but there are three orientations of the p orbital, for a total of 6 electrons. Starting at the lowest energy level, 2 electrons go in 1s, making $1s^2$; then 2 go in 2s, making $2s^2$. That is a total of 4 electrons so far. Next $2p^6$, $3s^2$, and $3p^6$ use 14 more electrons for a total of 18 so far. The remaining electron goes into the next sublevel, $4s^1$ and all the electrons have been used, giving a config - uration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

Question #20: What is the energy of a photon of red light with a frequency of 4.00 X 10^{14} Hz?

(A). 1.66 X 10^{-48} J
(B). 2.65 X 10^{-19} J
(C). 6.63 X 10^{-34} J
(D). 4.00 X 10^{14} J

#20 Answer (B). 2.65 X 10^{-19} J

Listing the known and unknown quantities:

Frequency	$f = 4.00 \times 10^{14}$ Hz
Planck's constant	$h = 6.63 \times 10^{-34}$ Js
Energy	$E = ?$

The relationship between the frequency (f) and energy (E) of a photon is found in equation, E = hf.

$$\begin{aligned} E &= hf \\ &= \left(6.63 \times 10^{-34}\ J \cdot s\right)\left(4.00 \times 10^{14}\ \frac{1}{s}\right) \\ &= \left(6.63 \times 10^{-34}\right)\left(4.00 \times 10^{14}\right)\ J \cdot s \times \frac{1}{s} \\ &= 2.65 \times 10^{-19}\ \frac{J \cdot s}{s} \\ &= 2.65 \times 10^{-19}\ J \end{aligned}$$

Elements and the Periodic Table

Question #1: Compared to metals, nonmetals are

(A). more brittle as a solid.
(B). better electrical conductors.
(C). more ductile as a solid.
(D). better conductors of heat.

#1 Answer

Question #2: Under ordinary, room temperature conditions, the greatest number of elements are

(A). gases.
(B). liquids.
(C). metallic solids.
(D). nonmetallic plasmas.

#2 Answer

Question #3: A solution of sugar dissolved in water is

(A). a heterogeneous mixture.
(B). a homogeneous mixture.
(C). an alloy.
(D). a pure substance.

#3 Answer

Question #4: Which of the following represents a physical change?

(A). Electricity is used to generate oxygen and hydrogen from water.
(B). Calcium carbonate is dissolved by stomach acid.
(C). Solid ice is melted into liquid water.
(D). Natural gas is burned as a heat source.

#4 Answer

Question #5: Carbon cannot be broken down into anything simpler by chemical means, so carbon must be

(A). a heterogeneous mixture.
(B). a homogeneous mixture.
(C). an element.
(D). a compound.

#5 Answer

Question #6: How many naturally occurring elements are found on the earth in significant quantities?

(A). 112
(B). 92
(C). 89
(D). 32

#6 Answer

Question #7: About 75 percent of the earth's solid surface is made up of

(A). silicon and oxygen.
(B). nitrogen and oxygen.
(C). hydrogen and oxygen.
(D). hydrogen, oxygen, and carbon.

#7 Answer

Question #8: About 99 percent of the earth's atmospheric air is made up of

(A). silicon and oxygen.
(B). nitrogen and oxygen.
(C). hydrogen and oxygen.
(D). hydrogen, oxygen, and carbon.

#8 Answer

Question #9: Isotopes are atoms of an element with identical chemical properties but with different

(A). numbers of protons.
(B). masses.
(C). numbers of electrons.
(D). atomic numbers.

#9 Answer

Question #10: The masses of all isotopes are based on a comparison to the mass a particular isotope of
(A). hydrogen.
(B). carbon.
(C). oxygen.
(D). uranium.

#10 Answer

Question #11: The sum of the number of protons and neutrons in the nucleus of an atom is called the

(A). atomic number.
(B). atomic mass.
(C). atomic weight.
(D). mass number.

#11 Answer

Question #12: The weighted average of the masses of the stable isotopes of an element as they occur in nature is called the

(A). atomic number.
(B). atomic mass.
(C). atomic weight.
(D). mass number.

#12 Answer

Question #13: The modern periodic law is based on

(A). atomic number.
(B). atomic mass.
(C). atomic weight.
(D). chemical activity.

#13 Answer

Question #14: Each family, or group of elements in a vertical column of the periodic table has elements with chemical characteristics that are

(A). exactly the same.
(B). similar.
(C). different.
(D). exactly opposite.

#14 Answer

Question #15: Which of the following belongs to the alkali metal family of elements?

(A). sodium
(B). calcium
(C). chlorine
(D). neon

#15 Answer

Question #16: Which of the following belongs to the halogen family of elements?

(A). sodium
(B). calcium
(C). chlorine
(D). neon

#16 Answer

Question #17: Which of the following belongs to the noble gas family of elements?

(A). sodium
(B). calcium
(C). chlorine
(D). neon

#17 Answer

Question #18: An atom of an element belonging to the alkali metal family has

(A). one outer shell electron.
(B). two outer shell electrons.
(C). all outer shell electrons but one.
(D). all outer shell electrons.

#18 Answer

Question #19: An atom of an element belonging to the halogen family has

(A). one outer shell electron.
(B). two outer shell electrons.
(C). all outer shell electrons but one.
(D). all outer shell electrons.

#19 Answer

Question #20: An atom of an element belonging to the noble gas family has

(A). one outer shell electron.
(B). two outer shell electrons.
(C). all outer shell electrons but one.
(D). all outer shell electrons.

#20 Answer
