Interactive Chemistry Multimedia Courseware

Atomic Structure
Program Supplement

Atomic Structure
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Introduction & Modern Atomic Theory

Scene Number

Scene 1

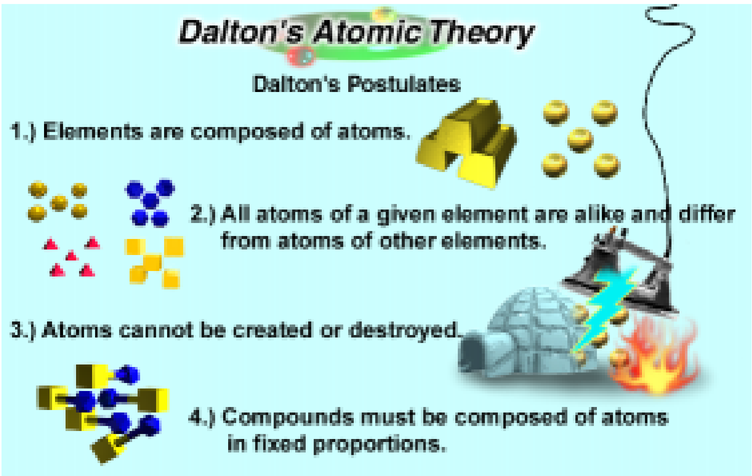
Welcome to Cyber Ed’s presentation of Atomic Structure. In this program, we will clearly define the concept of an atom, show how the structure of an atom has been determined through experimental investigation, and learn how Quantum Theory has shaped our conception of that structure. We will discuss the various ways the particles in an atom can be arranged and the implications of the instability of some atoms. We will use some of the ideas and definitions presented in the Program “Introduction to Chemistry”; you might want to go back and review that program before starting this one.

Scene 2

Everything is made up of elements. Elements can combine to form compounds, and compounds are usually found in mixtures -- all of these terms help us define what we call matter. In this program, we’re going to see what elements are made of. Long ago, some people made good guesses about elemental composition. The ancient Greek philosopher Democritus reasoned that there must be a smallest individual particle of an element that still has the characteristics of that element. He called that smallest particle an atom -- the Greek word for indivisible. In the Middle Ages, Galileo the astronomer knew of Democritus’ thinking, and he stated his belief that the creation of a new compound must involve rearrangement of these atoms.

Scene 3

In 1803, Dalton proposed his Atomic Theory of Matter. He stated that elements are composed of tiny particles, called atoms, and that all atoms of a given element are exactly alike and differ from atoms of other elements. He further stated that atoms cannot be created or destroyed, whether by lightning, heat, cold, or pressure, and that compounds must be composed of atoms in fixed proportions. This set of postulates, or hypotheses, was essentially correct, but it took scientists almost 100 years to prove it.



Scene 4

Around the year 1900, scientific studies finally confirmed that atoms are, indeed, the smallest elemental particles. More recently, physicists have actually photographed an atom! Not with an Instamatic, to be sure, since a thousand atoms would fit into a single wavelength of visible light, but using a device called a scanning tunneling microscope. By observing the way that a very fine-pointed electronic probe moves up and down as it crosses a copper atom, a topographic map of the surface of the atom can be produced.

Scene 5

Just like Democritus, Galileo, and Dalton, the scientists exploring the structure of the atom were faced with the task of understanding something they could not see. However, they could study the behavior of atomic particles subjected to electric fields and magnetic fields, and from this they concluded that there were both positive and negative particles in an atom. Also, they could determine how much deflection occurred when larger atoms are bombarded with smaller atoms, and from this conclude that most of the mass of the atom was concentrated in a tiny area near the center. On the basis of these experiments, scientists were able to create a series of hypotheses, or possible explanations, about the actual structure of the atom. The hypotheses described the atom as being mostly empty space, with a positively-charged dense central core surrounded by a negatively-charged shell of particles.

Scene 6

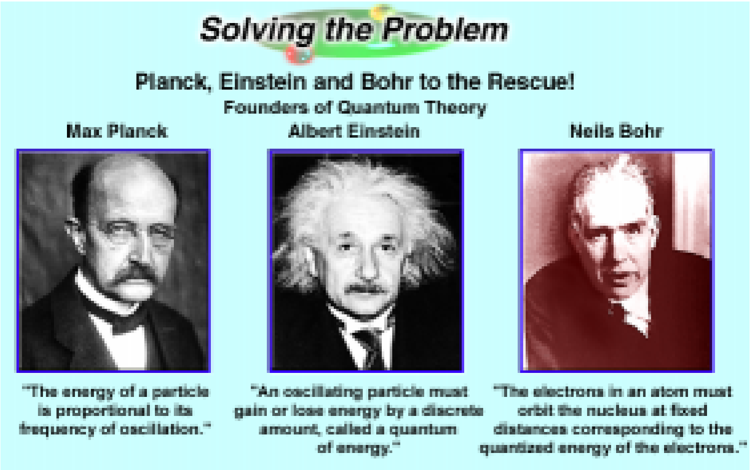
Additional experimental work confirmed and refined this hypothetical atomic structure, giving rise to the Planetary Theory of Atoms. This theory suggested that the structure of an atom could be compared to that of our Solar System -- a large, dense central core like the Sun, surrounded by small orbiting objects like planets.

Scene 7

Although the Planetary Theory presents a convenient visual picture of an atom, it has some problems. For instance, Classical Electromagnetic Theory says that particles of opposite charges will be attracted to one another and particles of like charges will repel one another. Why then don't those positive particles in the core fly apart? And why aren't the negative particles attracted toward the positive core? Furthermore, why do the different orbits contain two, or eight, or eighteen, or even thirty-two negative particles at approximately the same distance from the core? Unlike atoms, planets don’t have charges and don’t cluster in their orbits, so a better model of Atomic Structure was needed.

Scene 8

The discovery of the neutron, a particle similar to the other particle in the core of the atom but having no charge, answered the question about what keeps the core from flying apart. Neutrons, through a series of complex interactions, help to actually hold the positive particles together. As long as we are naming things, the positive particles we’ve shown in red are called protons, the negative ones in yellow are electrons, and the core of the atom is called the “nucleus”. Collectively, protons and neutrons are called nucleons.



Scene 9

Classical Electromagnetic Theory says that particles of opposite charges will be attracted to one another and particles of like charges will repel one another. Yet, as previously mentioned, negative electrons don’t spi ral into the positive nucleus and disintegrate, and multiple numbers of negative electrons can occupy the same orbit around the nucleus. These problems were addressed by Planck, Einstein, and Bohr during the first two decades of the twentieth century, when they developed the quantum theory of the Nuclear Atom -- often called the Bohr Model.

Scene 10

The Bohr Model of the atom hypothesizes that as long as the electron moves with a certain minimum speed and stays in a certain fixed orbit, the orbit will remain stable and the electron cannot be sucked into the nucleus. You might compare this to your spinning a lightstick on a string around your head. As long as the lightstick keeps moving at a minimum speed, the string stays taught, but if the lightstick slows down, it drops out of its orbit. The Bohr Model also accounts for multiple electrons occupying each of these fixed orbits. This is explained more fully in the program entitled "Electronic Structure", which goes into the details of Quantum Theory as it applies to electron orbits.

Scene 11

Although Quantum Theory resolved many of the aspects of atomic structure, there were some troublesome discrepancies between the theory and experimental observations of particle behavior. In the 1920's, de Broglie postulated that protons, and especially electrons, behave the way light does. That is, they possess both wave-like and particle-like properties. The existence of this duality remains one of the puzzles of Science, but it certainly helps explain some of the chemical and physical properties of atoms. Working from de Broglie's postulates, Schroedinger developed a model of the atom that incorporated both the particle and wave components; our atomic displays are based on the Schroedinger electronic orbitals, but we will explicitly show the Bohr electrons in those orbitals. For our purposes, we will consider atomic particles to be just that -- particles -- but we will show the shapes of electronic orbitals in accordance with their wave-like behavior.

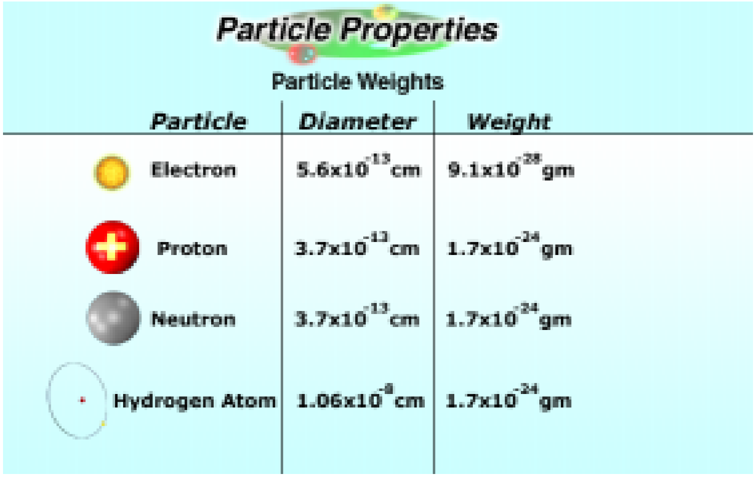
Scenes 12-24

Atomic Particles, Nuclear Chemistry, Summary & Conclusion

Scene Number

Scene 12

As mentioned earlier, atoms and their component particles are quite small. For comparison, we've superimposed a model of a neon atom onto the solar system by expanding the atom until it has the same diameter as the solar system. As you can see, the atom, like the solar system, is mostly empty space, but you can also see that the electron orbitals are nothing like planetary orbits. Note also that the weight of the electron is only about 1/2000 of the weight of a proton, although both have about the same diameter.



Particle Properties

Particle Weights

Particle	Diameter	Weight
Electron	$5.6x10^{-13}$cm	$9.1x10^{-28}$gm
Proton	$3.7x10^{-13}$cm	$1.7x10^{-24}$gm
Neutron	$3.7x10^{-13}$cm	$1.7x10^{-24}$gm
Hydrogen Atom	$1.06x10^{-8}$cm	$1.7x10^{-24}$gm

Scene 13

An atom in its elemental form is electrically neutral; that is, it always has exactly as many electrons as it has protons. The chemical properties of an element are determined entirely by the number of electrons that surround the nucleus. An atom can gain or lose one or more electrons to become an ion, which is the name we use for a charged atom. Since the number of electrons has changed, the chemical properties of an ion differ from those of the corresponding uncharged atom. However, the element that the ion represents remains the same, since the identity of a particular atom is determined solely by the number of protons. We refer to the number of protons as the atomic number of the element. In the program “An Introduction to Chemistry”, we showed how the elements are arranged in the Periodic Table by increasing atomic number.

Scene 14

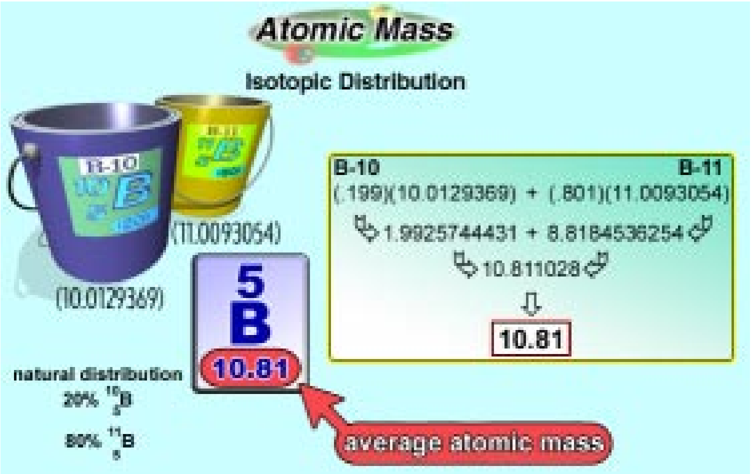
Although the atomic number, the number of protons in an atom of a particular element, is fixed, the number of neutrons is not. Atoms of the same element that have different numbers of neutrons are called “isotopes”. The number of neutrons does not affect the chemical properties of the element, since those properties are determined entirely by the number of electrons.

Scene 15

Although the isotopes of a particular element have differing numbers of neutrons, they have the same number of protons, or atomic number. In order to distinguish among isotopes, we need to measure something besides protons -- total number of nucleons, for instance. If we take the sum of the numbers of protons plus neutrons, which accounts for practically all the mass of the isotope, we get its mass number; this number is used to identify a particular isotope of an element. By agreeing that carbon-12, with six protons and six neutrons, is defined as having a mass of exactly 12, Chemists have established a scale of relative masses for each isotope. These masses are called Atomic Weights, and they are essentially identical with the Mass numbers. Note that the Mass Number does not uniquely identify the element -- both hydrogen-3 and helium-3 have the same number of nucleons and thus the same Mass Number, but they have different atomic Numbers and thus are different elements.

Scene 16

Although the mass number of an atom is a whole number, the atomic mass of an atom is in fact, fractional. You can get fairly close to calculating the atomic mass of an atom by adding up the masses of all the protons, neutrons, and electrons in that atom. However, energy is released when the nucleons bind together and since energy and mass are related, some of the mass of the atom is lost as energy when the nucleus forms. Therefore, an accurate value for the atomic mass cannot be calculated by simply adding together the masses of the individual subatomic particles, but rather must be obtained by experiments performed on the individual isotopes.



Notice that the atomic mass as reported on the periodic table is slightly different than the atomic mass for either isotope of boron. This is because the atomic mass as listed on the periodic table is actually an average, calculated from the atomic masses of all naturally occurring isotopes of that particular element. The atomic mass value as provided by the periodic table is referred to as the average atomic mass.

Scene 17

Quantum Theory states that only certain ratios of neutrons to protons in the nucleus can exist. Too few or too many neutrons will cause the nucleus to spontaneously decay; that is, it will eject one or more particles from the nucleus in order to achieve a stable ratio of neutrons to protons. We've shown a plot of neutrons vs protons for all of the stable isotopes of all the elements. You can see a very narrow Belt of Stability that favors having a few more neutrons than protons. Isotopes that are outside the belt, and even some that are inside, decay in such a way as to get closer to the center of the belt. The farther the isotope is from the belt, the faster that decay will occur. Elements that contain these unstable isotopes are described as radioactive, and chemists handling such elements must take special precautions to avoid the high-energy emissions that accompany the decay of these radioactive isotopes.

Scene 18

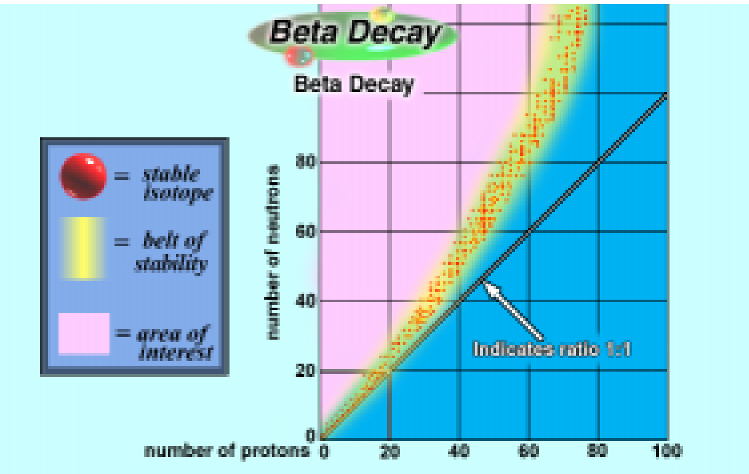
Depending on their neutron to proton ratio, radioactive isotopes decay in different ways. If the nucleus has fewer neutrons than protons, the nucleus either emits a positive particle, called a positron, or captures one of the orbital electrons. In either case, a proton is effectively converted to a neutron and the atom's atomic number has decreased by one.

Scene 19

If the isotope has more neutrons than protons, but still has too few neutrons to be in the Belt of Stability, or if it simply has too many nucleons to be stable, it ejects a particle composed of two protons and two neutrons, called an Alpha Ray, which ultimately captures two electrons to become a Helium atom. The remaining atom has decreased its atomic number by two and its atomic mass by four.

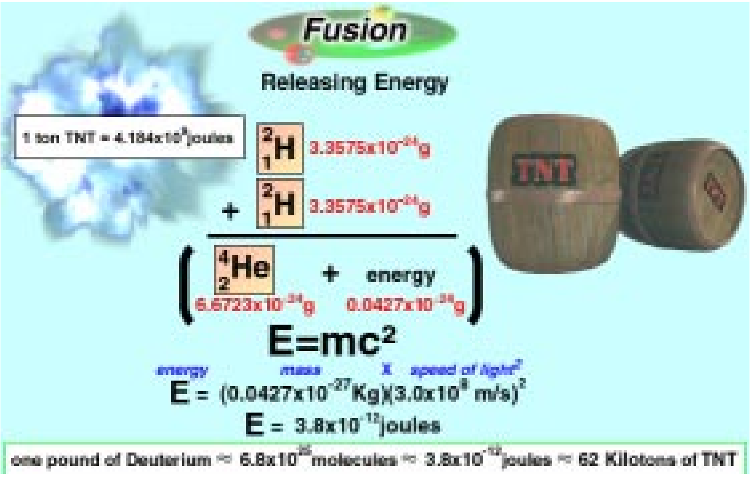
Scene 20

If the atom has too many neutrons, the decay usually involves the emission of an electron, called a beta ray, from the nucleus. Since the nucleus contains no electrons, where did the ejected one come from? An unstable neutron can escape not only the nucleus but the entire atom -- you might have noticed previously that the mass of a neutron exceeds the mass of a proton by almost exactly the mass of an electron. Since the atom now has one more proton, it has become ionized, and it needs to capture an electron so its electrical charge stays neutral. Notice how the atomic number has increased by one, even though the mass number stays the same. Some radioactive isotopes initiate a whole series of alpha and beta decays, such as the decay of uranium-238, which eventually produces a stable isotope of Lead.



Scene 21

Some isotopes, especially the heaviest ones, show another type of decay. These elements actually eject a neutron from the atomic core. This expulsion is accompanied by a burst of energy called a gamma ray. Gamma rays may also accompany very energetic alpha and beta decay processes. The released neutron is highly energetic, and it may strike other unstable isotopes which in turn, may strike other unstable isotopes, and so forth. This phenomenon is called a chain reaction. Another term for this self-destructive process is fission, the driving force in nuclear reactors and the Atomic Bomb.



Scene 22

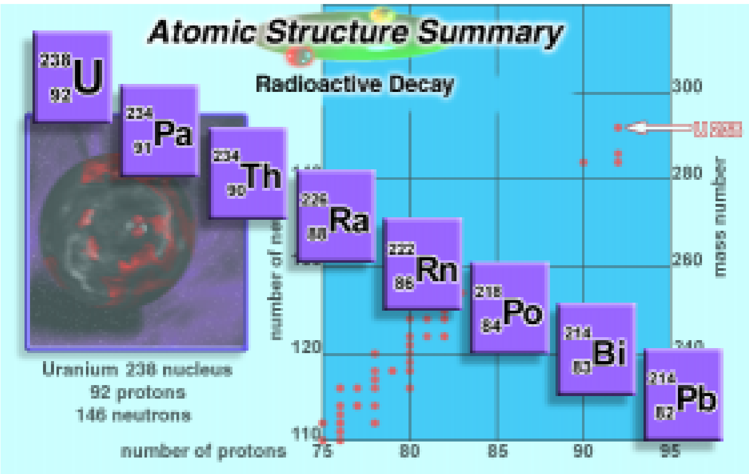
The reverse of the destructive fission process is called fusion, where two nuclei combine to produce a single nucleus. Beams of protons, Deuterium atoms, and Helium atoms are commonly used to generate fusion reactions. Almost invariably, the new nucleus weighs less than the sum of the two nuclei that join, and the difference in mass is converted into energy, as Einstein explained in his famous equation that calculates the amount of energy created to be the loss in mass times the speed of light squared. Fusion powers the sun, and a Hydrogen (actually Deuterium) Bomb is a fusion reaction.

Scene 23

In a scientific equivalent of throwing rocks at a rock-pile, physicists use particle accelerators to hurl nuclei at other nuclei. Their purpose is to produce nuclei that are far from the Belt of Stability and determine, from the way that they decay, what actually happens inside the nucleus. By choosing the proper set of rocks, physicists can even create new elements -- all of the elements with atomic numbers higher than Plutonium were built in this way. Most of the recently-created ones are so unstable that they decay in fractions of a second -- isolating even a few atoms of them long enough to study their properties has been one of the highlights of modern chemical accomplishment.

Scene 24

In this program, we've started with the model of an atom as a simple billiard ball. We then explored the Planetary Model,with a central positive nucleus surrounded by fixed orbits containing fixed numbers of electrons. This model was supplanted by the Bohr Model, where the various component parts of the atom behave as quantum particles instead of following classical electromagnetic theory. Finally, we arrived at the Schroedinger Model, where the electron behavior is expressed by quantum wave equations that describe variously shaped electronic orbitals. We've also looked briefly at characterization of atoms in terms of their atomic weights and atomic numbers and the effects of atomic instability on the structure of the atom. The next program in this series takes a detailed look at the influence of electronic structure on chemical properties.



Atomic Structure
Exam

1. The ancient Greek Philosopher that developed the idea of atoms was ________.

   A. Pythagoras
   B. Aristotle
   C. Democritus
   D. Sophocles

2. Galileo thought the creation of new compounds must involve _____.

   A. input of energy
   B. the Philosopher’s Stone
   C. rearrangement of atoms
   D. divine intervention

3. Dalton’s Fourth Postulate states that atoms combine ________.

   A. in fixed proportions
   B. when their colors match
   C. indiscriminately
   D. in the dark

4. Dalton’s Third Postulate states that atoms can be destroyed by _______.

   A. fire
   B. lightning
   C. pressure
   D. none of the above

5. The _________ microscope allows scientists to view individual atoms.

   A. enhanced optical
   B. scanning tunneling
   C. polarized electron
   D. focused neutrino

6. Most of the mass of the atom is located ________.

    A. in the middle
    B. on the circumfrence
    C. equally distributed throughout the atom
    D. trick question -- atoms have no mass

7. Beams of charged atomic particles _________.

    A. bend in an electrical field
    B. bend in a magnetic field
    C. light up a TV screen
    D. all of the above

8. The Planetary Theory of atoms suggests that ___________.

    A. electrons, like planets, can support life
    B. the nucleus shines as brightly as the Sun
    C. electrons, like planets, are of different sizes
    D. electrons revolve around the nucleus like planets around the sun

9. According to Classical Electromagnetic Theory, electrons should ________.

    A. repel one another
    B. spiral into the nucleus
    C. maintain their negative charge
    D. all of the above

10. _________ help hold the nucleus together.

    A. Gluons
    B. Neutrons
    C. Alpha particles
    D. Protons

11. The Bohr Model of the Atom resulted from the development of _______.

A. Planetary Theory
B. Quantum Theory
C. Quantum Mechanics
D. none of the above

12. The Schroedinger Model of the Atom treats electrons as ________.

A. particles
B. waves
C. both particles and waves
D. neither particles nor waves

13. Atoms of the same element such as gold all possess the same number of _______.

A. electrons
B. neutrons
C. protons
D. nuclei

14. A proton is approximately _______ times as heavy as an electron.

A. 2
B. 20
C. 200
D. 2000

15. The chemical properties of an atom are determined by its Atomic ______.

A. Mass
B. Symbol
C. Number
D. Diameter

16. An atom that has gained or lost electrons is _______.

A. a charged atom
B. an ion
C. still has the same number of protons
D. all of the above

17. All of the isotopes of any given element have ____.

A. the same number of nucleons
B. the same number of protons
C. the same number of neutrons
D. all of the above

18. The Atomic Weight scale is based upon international agreement that ______.

A. Hydrogen-1 weighs exactly 1 Atomic Mass Unit
B. Boron-11 weighs exactly 11 Atomic Mass Units
C. Carbon-12 weighs exactly 12 Atomic Mass Units
D. Oxygen-16 weighs exactly 16 Atomic Mass Units

19. The Atomic Weights of the elements contain fractional amounts because ____.

A. the electrons add to the weight
B. the proton and neutron have different weights
C. the isotopic distribution varies
D. all of the above

20. Radioactive decay occurs when the nucleus ejects ________.

A. an Alpha particle
B. a Beta particle
C. a neutron
D. all of the above

21. An Alpha particle is composed of _____________.

    A. one electron
    B. one proton
    C. one proton and one neutron
    D. two protons and two neutrons

22. An atom that contains _______.is always unstable.

    A. more neutrons than protons
    B. more protons than neutrons
    C. an odd number of nucleons
    D. all of the above

23. An atom with too many protons can decay by ________.

    A. capturing an orbiting electron
    B. emitting a positron
    C. both of the above
    D. neither of the above

24. Alpha decay is confined primarily to _______.

    A. the heavier elements
    B. the lighter elements
    C. the white elements
    D. the stable elements

25. A Beta particle is actually _________.

    A. a proton
    B. a neutron
    C. an electron
    D. a nucleon

26. A Gamma ray is ______ .

    A. a burst of energy
    B. associated with radioactive decay
    C. dangerous to humans
    D. all of the above

27. Another name for “atomic fission” is ______.

    A. nuclear chain reaction
    B. atomic fusion
    C. gamma radiation
    D. solar furnace

28. Fusion reactions occur in _______.

    A. an atomic bomb
    B. a hydrogen bomb
    C. a nuclear power plant
    D. none of the above

29. Particle accelerators are used to ________.

    A. build new elements
    B. disintegrate atoms
    C. study nuclear structure
    D. all of the above

30. A substance that weighs 10 grams and has a volume of 2 cm3 would have a density of ______.

    A. 5 g/cm^3.
    B. 10g/cm^3.
    C. 2g/cm^3.
    D. 15g/cm^3.

31. The terms Kelvin, Celsius and Fahrenheit all correspond to units of ___________.

    A. density
    B. hardness
    C. volume
    D. temperature

32. The _________ is the standard unit for measurements of length.

A. cetimeter
B. foot
C. meter
D. yard

33. One kilogram is equal to _________ pounds.

A. 1.2
B. 2.2
C. 3.2
D. 4.2

Atomic Structure

Exam Answer Key

1.	C	17.	B
2.	C	18.	C
3.	A	19.	D
4.	D	20.	D
5.	B	21.	D
6.	A	22.	B
7.	D	23.	C
8.	D	24.	A
9.	D	25.	C
10.	B	26.	D
11.	B	27.	A
12.	C	28.	B
13.	C	29.	D
14.	D	30.	A
15.	C	31.	D
16.	D	32.	C
		33.	B