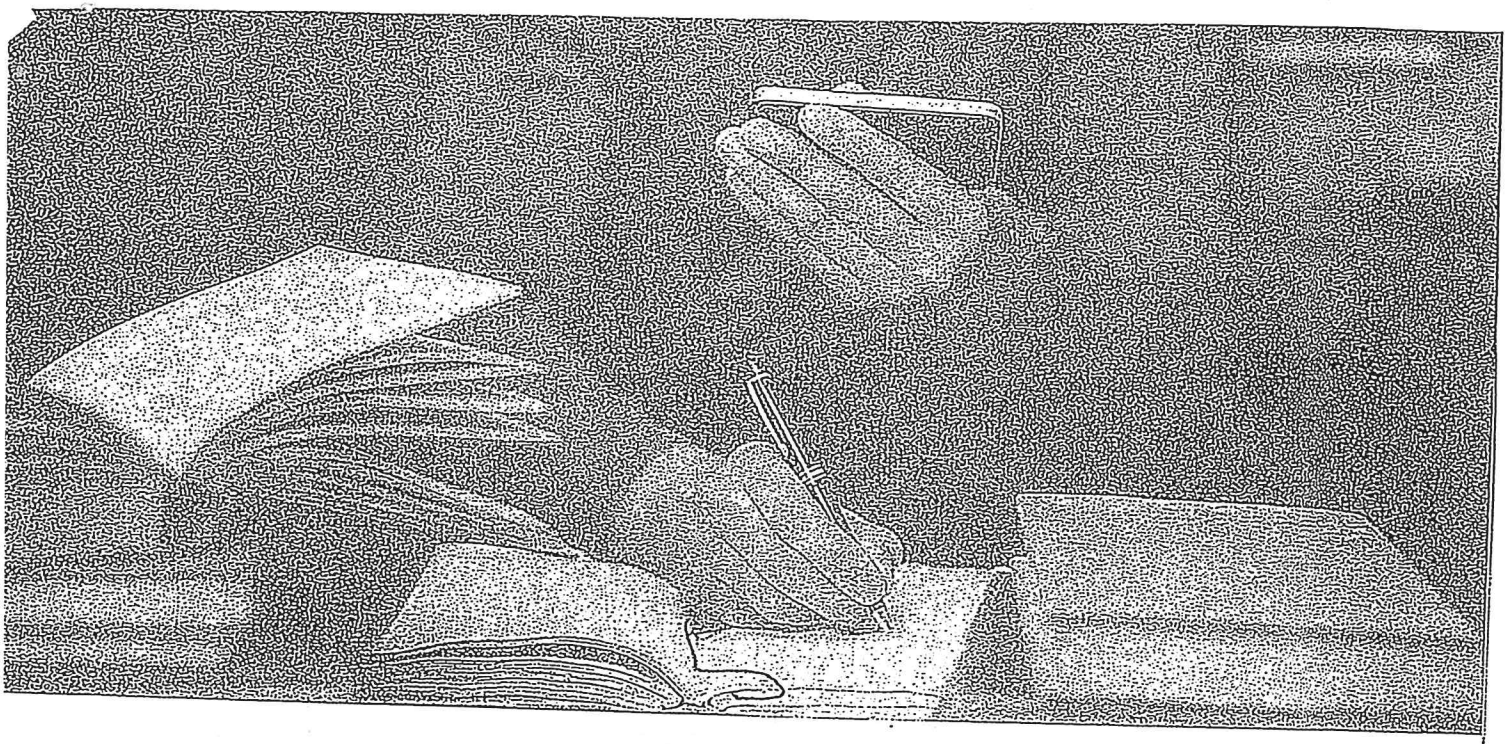


Chapter 3
Big Idea #1: Atoms, Elements, and the Building Blocks of Matter

The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

THE PERIODIC TABLE

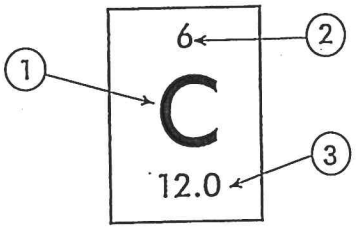
The most important tool you will use on this test is the Periodic Table of the Elements.

PERIODIC TABLE OF THE ELEMENTS

1 IA												13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1 H 1.008	2 IIA																2 He 4.00
3 Li 6.91	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.30	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8	9 VIIIB	10	11 IB	12 IIB	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57 *La 138.91	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.2	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.02	89 †Ac 227.03	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (271)	111 Rg (272)							

*Lanthanide Series	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.4	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
†Actinide Series	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 X (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

The Periodic Table gives you very basic but very important information about each element.



1. This is the symbol for the element; carbon, in this case. On the test, the symbol for an element is used interchangeably with the name of the element.
2. This is the atomic number of the element. The atomic number is the same as the number of protons in the nucleus of an element; it is also the same as the number of electrons surrounding the nucleus of an element when it is neutrally charged.
3. This number represents the average atomic mass of a single atom of carbon, measured in atomic mass units (amus). It also represents the average mass for a mole (see p. 70) of carbon atoms, measured in grams. Thus, one mole of carbon atoms has a mass of 12.01 g. This is called the molar mass of the element.

The horizontal rows of the periodic table are called periods.

The vertical columns of the periodic table are called groups.

Bad Joke Alert
To prepare for the AP Chemistry Exam, familiarize yourself with the periodic table of elements by looking at it periodically.

Groups can be numbered in two ways. The old system used Roman numerals to indicate groups. The new system simply numbers the groups from 1 to 18. While it is not important to know the specific group numbers, it is important to know the names of some groups.

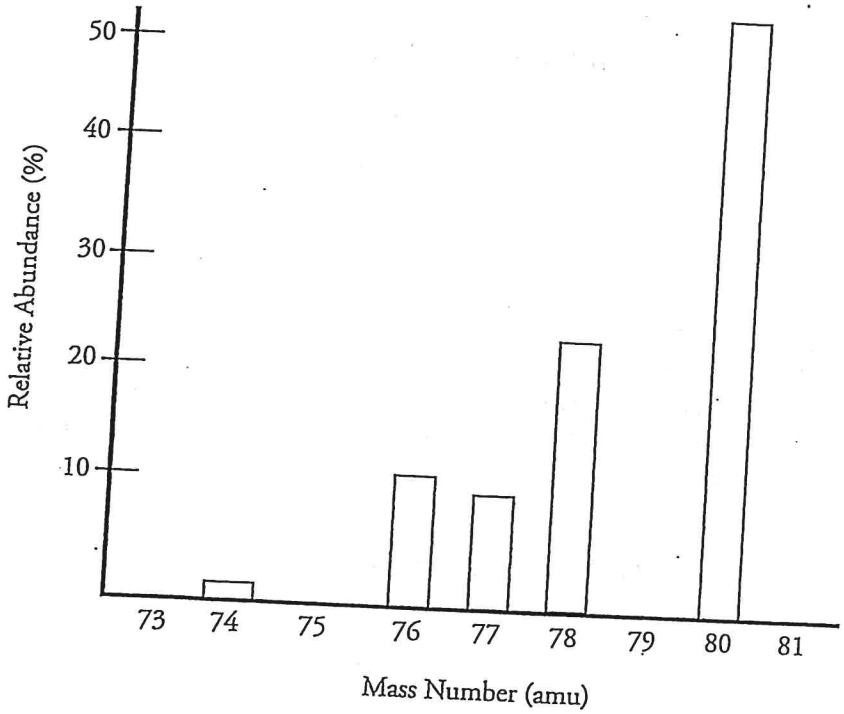
Group IA/1 – Alkali Metals
Group IIA/2 – Alkaline Earth Metals
Group B/3-12 – Transition Metals
Group VIIA/17 – Halogens
Group VIIIA/18 – Noble Gases

In addition, the two rows offset beneath the table are alternatively called the lanthanides and actinides, the rare Earth elements, or the inner transition metals.

The identity of an atom is determined by the number of protons contained in its nucleus. The nucleus of an atom also contains neutrons. The mass number of an atom is the sum of its neutrons and protons. Electrons have significantly less mass than protons or neutrons and do not contribute to an element's mass.

Atoms of an element with different numbers of neutrons are called isotopes; for instance, carbon-12, which contains 6 protons and 6 neutrons, and carbon-14, which contains 6 protons and 8 neutrons, are isotopes of carbon. The molar mass given on the periodic table is the average of the mass numbers of all known isotopes weighted by their percent abundance.

The mass of various isotopes of an element can be determined by a technique called mass spectrometry. A mass spectrum of selenium looks like the following:



As you can see, the most abundant isotope of selenium has a mass of 80, but there are four other naturally occurring isotopes of selenium. The average atomic mass is the weighted average of all five isotopes of selenium shown on this spectra.

The molar mass of an element will give you a pretty good idea of the most common isotope of that element. For instance, the molar mass of carbon is 12.01 and about 99 percent of the carbon in existence is carbon-12.

MOLES

The mole is the most important concept in chemistry, serving as a bridge that connects all the different quantities that you'll come across in chemical calculations. The coefficients in chemical reactions tell you about the reactants and products in terms of moles, so most of the stoichiometry questions you'll see on the test will be exercises in converting between moles and grams, liters, molarities, and other units.

Moles and Molecules

Avogadro's number describes how many atoms are in a single mole of any given element. Much like a dozen is always 12, Avogadro's number is always 6.02×10^{23}. While it technically can be used to count anything, due to its extremely large value it is usually only used to count extraordinarily small things. Within the confines of this book, it will be used to count atoms, molecules, electrons, or ions, depending on the problem.

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ particles}$$

$$\text{Moles} = \frac{\text{particles}}{\left(6.022 \times 10^{23}\right)}$$

Moles and Grams

Moles and grams can be related using the atomic masses given in the periodic table. Atomic masses on the periodic table are given in terms of atomic mass units (amu); however, they also signify how many grams are present in one mole of an element. So, if 1 carbon atom has a mass of 12 amu, then 1 mole of carbon atoms has a mass of 12 grams.

You can use the relationship between amu and g/mol to convert between grams and moles by using the following equation:

$$\text{Moles} = \frac{\text{grams}}{\text{molar mass}}$$

Moles and Gases

We'll talk more about the ideal gas equation in Chapter 4, but for now, you should know that you can use it to calculate the number of moles of a gas if you know some of the gas's physical properties. All you need to remember at this point is that in the equation $PV = nRT$, n stands for moles of gas.

$$\text{Moles} = \frac{PV}{RT}$$

P = pressure (atm)
V = volume (L)
T = temperature (K)
R = the gas constant, 0.0821 L·atm/mol·K

The equation above gives the general rule for finding the number of moles of a gas. Many gas problems will take place at STP, or standard temperature and pressure, where $P = 1$ atmosphere and $T = 273$ K. At STP, the situation is much simpler and you can convert directly between the volume of a gas and the number of moles. That's because at STP, one mole of gas always occupies 22.4 liters.

$$\text{Moles} = \frac{\text{liters}}{(22.4\text{L/mol})}$$

Molarity

Molarity (M) expresses the concentration of a solution in terms of volume. It is the most widely used unit of concentration, turning up in calculations involving equilibrium, acids and bases, and electrochemistry, among others. When you see a chemical symbol in brackets on the test, that means they are talking about molarity. For instance, "$[Na^+]$" is the same as "the molar concentration (molarity) of sodium ions."

$$\text{Molarity } (M) = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Percent Composition

Percent composition is the percent by mass of each element that makes up a compound. It is calculated by dividing the mass of each element or component in a compound by the total molar mass for the substance.

Calculate the percent composition of each element in calcium nitrate, $Ca(NO_3)_2$.

To do this, you need to first separate each element and count how many atoms are present. Subscripts outside of parentheses apply to all atoms inside of those parentheses.

Calcium: 1
Nitrogen: 2
Oxygen: 6

Then, multiply the number of atoms by the atomic mass of each element.

Ca: $40.08 \times 1 = 40.08$
N: $14.01 \times 2 = 28.02$
O: $16.00 \times 6 = 96.00$

Adding up the masses of the individual elements will give you the atomic mass of that compound. Divide each individual mass by the total molar mass to get your percent composition.

$$40.08 + 28.02 + 96.00 = 164.10 \text{ g/mol}$$

Ca: $40.08/164.10 \times 100\% = 24.42\%$
N: $28.02/164.10 \times 100\% = 17.07\%$
O: $96.00/164.10 \times 100\% = 58.50\%$

You can check your work at the end by making sure your percents add up to 100% (taking rounding into consideration).

24.42% + 17.07% + 58.50% = 99.99%. Close enough!

Empirical and Molecular Formulas

You will also need to know how to determine the empirical and molecular formulas of a compound given masses or mass percents of the components of that compound. Remember that the empirical formula represents the simplest ratio of one element to another in a compound (e.g., CH_2O), while the molecular formula represents the actual formula for the substance (e.g., $C_6H_{12}O_6$).

Let's take a look at the following example:

A compound is found to contain 56.5% carbon, 7.11% hydrogen, and 36.4% phosphorus.

a) Determine the empirical formula for the compound.

We start by assuming a 100 gram sample; this allows us to convert those percentages to grams. After we have that done, each element needs to be converted to moles.

$$\text{C: } 56.5\text{g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 4.71 \text{ mol C}$$

$$\text{H: } 7.11 \text{ g Cl} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 7.04 \text{ mol H}$$

$$\text{P: } 36.4 \text{ g P} \times \frac{1 \text{ mol P}}{30.97 \text{ g P}} = 1.18 \text{ mol P}$$

We then divide each mole value by the lowest of the values. In this example, that would be the phosphorus. It is acceptable to round your answers if they are close (within 0.1) to a whole number.

$$\text{C: } \frac{4.71 \text{ mol}}{1.18 \text{ mol}} = 4$$

$$\text{H: } \frac{7.04 \text{ mol}}{1.18 \text{ mol}} = 6$$

$$\text{P: } \frac{1.18 \text{ mol}}{1.18 \text{ mol}} = 1$$

Those values become subscripts, so the empirical formula for the compound is C_4H_6P.

b) If the compound has a molar mass of 170.14 g/mol, what is its molecular formula?

First, we determine the molar mass of the empirical formula.

$$(12.01 \times 4) + (1.01 \times 6) + 30.97 = 85.07 \text{ g/mol}$$

Then, we divide that mass into the molar mass.

$$\frac{170.14}{85.07} = 2$$

Finally, multiply all subscripts in the empirical formula by that value. So, the molecular formula is $C_8H_{12}P_2$.

Electron Configurations and the Periodic Table

The positively-charged nucleus is always pulling at the negatively-charged electrons around it, and the electrons have potential energy that increases with their distance from the nucleus. It works the same way that the gravitational potential energy of a brick on the third floor of a building is greater than the gravitational potential energy of a brick nearer to ground level.

The energy of electrons, however, is quantized. That's important. It means that electrons can exist only at specific energy levels, separated by specific intervals. It's kind of like if the brick in the building could be placed only on the first, second, or third floor of the building, but not in-between.

The Aufbau Principle

The Aufbau principle states that when building up the electron configuration of an atom, electrons are placed in orbitals, subshells, and shells in order of increasing energy.

The Pauli Exclusion Principle

The Pauli Exclusion Principle states that the two electrons which share an orbital cannot have the same spin. One electron must spin clockwise, and the other must spin counterclockwise.

Hund's Rule

Hund's Rule says that when an electron is added to a subshell, it will always occupy an empty orbital if one is available. Electrons always occupy orbitals singly if possible and pair up only if no empty orbitals are available.

Watch how the 2*p* subshell fills as we go from boron to neon.

	1*s*	2*s*	2*p*
Boron	↿⇂	↿⇂	↿
Carbon	↿⇂	↿⇂	↿ ↿
Nitrogen	↿⇂	↿⇂	↿ ↿ ↿
Oxygen	↿⇂	↿⇂	↿⇂ ↿ ↿
Fluorine	↿⇂	↿⇂	↿⇂ ↿⇂ ↿
Neon	↿⇂	↿⇂	↿⇂ ↿⇂ ↿⇂

COULOMB'S LAW

The amount of energy that an electron has depends on its distance from the nucleus of the atom. This can be calculated using Coulomb's Law:

$$E = \frac{k(+q)(-q)}{r}$$

E = energy

k = Coulomb's constant

$+q$ = magnitude of the positive charge (nucleus)

$-q$ = magnitude of the negative charge (electron)

r = distance between the charges

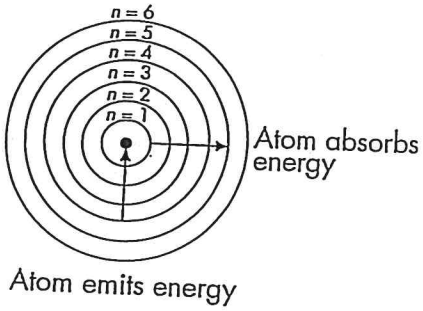
While on the exam, you will not be required to mathematically calculate the amount of energy a given electron has, you should be able to qualitatively apply Coulomb's Law. Essentially, the greater the charge of the nucleus, the more energy an electron will have (as all electrons have the same amount of charge). Coulombic potential energy is considered to be 0 at a distance of infinity. The Coulombic potential energy for a 1*s* electron is lower (more negative) than that of say, a 3*s* electron. The amount of energy required to remove a 1*s* electron, thereby bringing its Coulombic potential energy to zero, will thus be greater than the amount needed to remove a 3*s* electron. This removal energy is called the binding energy of the electron and is always a positive value.

Quantum Theory

Max Planck figured out that electromagnetic energy is quantized. That is, for a given frequency of radiation (or light), all possible energies are multiples of a certain unit of energy, called a quantum (mathematically, that's $E = h\nu$). So, energy changes do not occur smoothly but rather in small but specific steps.

The Bohr Model

Neils Bohr took the quantum theory and used it to predict that electrons orbit the nucleus at specific, fixed radii, like planets orbiting the Sun.



Each energy level is represented by a row on the periodic table. There are currently seven known energy levels, which correspond with $n = 1$ to $n = 7$. The closer an energy level is to an atom, the less energy electrons on that level have. While the Bohr model is not a perfect model of the atom, it serves as an excellent basis to understand atomic structure.

When atoms absorb energy in the form of electromagnetic radiation, electrons jump to higher energy levels. When electrons drop from higher to lower energy levels, atoms give off energy in the form of electromagnetic radiation.

The relationship between the change in energy level of an electron and the electromagnetic radiation absorbed or emitted is given below.

Energy and Electromagnetic Radiation

$$\Delta E = h\upsilon = \frac{hc}{\lambda}$$

ΔE = energy change
h = Planck's constant, 6.63×10^{-34} joule·sec
υ = frequency of the radiation
λ = wavelength of the radiation
c = the speed of light, 3.00×10^{8} m/sec

For a particular atom, the energy level changes of the electrons are always the same, so atoms can be identified by their emission and absorption spectra.

Frequency and Wavelength

$$c = \lambda\upsilon$$

c = speed of light in a vacuum (2.998×10^{8} ms^{-1})
λ = wavelength of the radiation
υ = frequency of the radiation

The frequency and wavelength of electromagnetic radiation are inversely proportional. Combined with the energy and electromagnetic radiation equation, we can see that higher frequencies and shorter wavelengths lead to more energy.

PHOTOELECTRON SPECTROSCOPY

If an atom is exposed to electromagnetic radiation at an energy level that exceeds the various binding energies of the electrons of that atom, the electrons can be ejected. The amount of energy necessary to do that is called the ionization energy for that electron. For the purposes of this exam, ionization energy and binding energy can be considered to be synonymous terms. When examining the spectra for electrons from a single atom or a small number of atoms, this energy is usually measured in electronvolts, eV (1 eV = 1.60×10^{-19} Joules). If moles of atoms are studied, the unit for binding energy is usually either kJ/mol or MJ/mol.

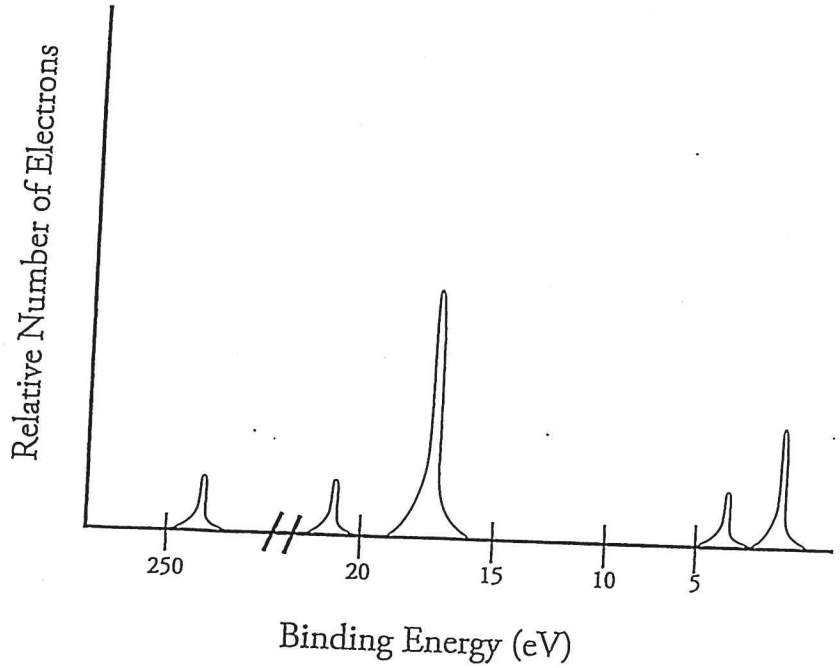
All energy of the incoming radiation must be conserved and any of that energy that does not go into breaking the electron free from the nucleus will be converted into kinetic energy (the energy of motion) for the ejected electron. So:

Incoming Radiation Energy =
Binding Energy + Kinetic Energy (of the ejected electron)

The faster an ejected electron is going, the more kinetic energy it has. Electrons that were originally further away from the nucleus require less energy to eject, and thus will be moving faster. So, by examining the speed of the ejected electrons, we can determine how far they were from the nucleus of the atom in the first place. Usually, it takes electromagnetic radiation in either the visible or ultraviolet range to cause electron emission, while radiation in the infrared range is often used to study chemical bonds. Radiation in the microwave region is used to study the shape of molecules.

Spectra

If the amount of ionization energy for all electrons ejected from a nucleus is charted, you get what is called a photoelectron spectra (PES) that looks like the following:



The *y*-axis describes the relative number of electrons that are ejected from a given energy level, and the *x*-axis shows the binding energy of those electrons. Unlike most graphs, binding energy (ionization energy) decreases going from left to right in a PES. The spectra on the previous page is for sulfur.

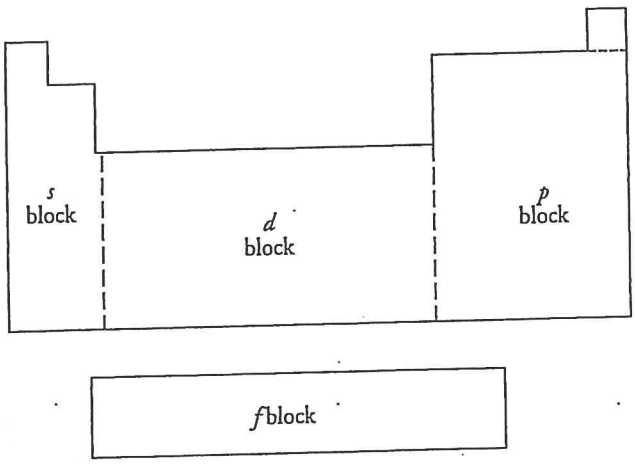
Each section of peaks in the PES represents a different energy level. The number of peaks in a section shows us that not all electrons at $n = 2+$ are located the same distance from the nucleus. With each energy level, there are subshells, which describe the shape of the space the electron can be found in (remember, we are in three-dimensional space here). The Bohr model is limited to two dimensions and does not represent the true positions of electrons due to that reason. Electrons do not orbit the nucleus as planets orbit the Sun. Instead, they are found moving about in a certain area of space (the subshell) a given distance (the energy level) away from the nucleus.

In all energy levels, the first subshell is called the *s*-subshell and can hold a maximum of two electrons. The second subshell is called the *p*-subshell and can hold a maximum of six electrons. In the spectra on the previous page, we can see the peak for the *p*-subshell in energy level 2 is three times higher than that of the *s*-subshell. The relative height of the peaks helps determine the number of electrons in that subshell.

In the area for the third energy level, the *p*-subshell peak is only twice as tall as the *s*-subshell. This indicates there are only four electrons in the *p*-subshell of this particular atom.

Electron Configuration

Studying the PES of elements allows scientists to understand more about the structure of the atom. In addition to the *s* and *p* subshells, two others exist: *d* (10 electrons max) and *f* (14 electrons max). The periodic table is designed so that each area is exactly the length of one particular subshell.



As you can see, the first two groups (plus helium) are in what is called the *s*-block. The groups on the right of the table are in the *p*-block, while transition metals make up the *d*-block. The inner transition metals below the table inhabit the *f*-block. The complete description of the energy level and subshell that each electron on an element inhabits is called its electron configuration.

Example 1: Determine the electron configuration for sulfur.

Sulfur has 16 electrons. The first two go into energy level 1 subshell *s*; this is represented by $1s^2$. The next two go into energy level 2 subshell *s*—$2s^2$. Six more fill energy level *p* ($2p^6$) then two more go into 3*s* ($3s^2$) and the final four enter into 3*p* ($3p^4$). So the final configuration is $1s^22s^22p^63s^23p^4$.

Using the periodic table as a reference can allow for the determination of any electron configuration. One thing to watch out for is that when entering the *d*-block (transition metals), the energy level drops by 1. The why behind that isn't important right now, but you should still be able to apply that rule.

Example 2: Determine the electron configuration for nickel.

$$1s^22s^22p^63s^23p^63d^84s^2$$

Electron configurations can also be written "shorthand" by replacing parts of them with the symbol for the noble gas at the end of the highest energy level which has been filled.

For example, the shorthand notation for Si is $[Ne]3s^23p^3$ and the shorthand for nickel would be $[Ar]4s^23d^8$.

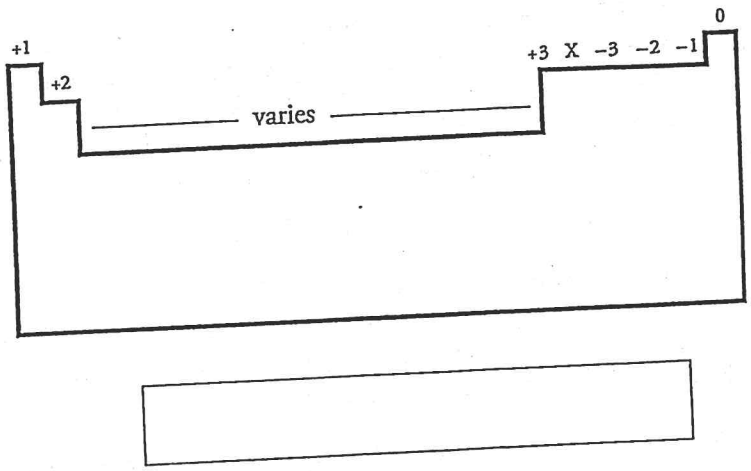
PREDICTING IONIC CHARGES

One of the great rules of chemistry is that the most stable configurations from an energy standpoint are those in which the outermost energy level is full. For anything in the *s* or *p* blocks, that means achieving a state in which there are eight electrons in the outermost shell (2 in the *s* subshell and 6 in the *p* subshell).

Elements that are close to a full energy level, such as the halogens or those in the oxygen group, tend to gain electrons to achieve a stable configuration. An **ion** is an atom which has either gained or lost electrons, while the number of protons and neutrons remains constant. Halogens need only gain one electron to achieve a stable configuration, and as such, typically form ions with a charge of negative one. F^-, Cl^-, etc. Any particle with more electrons than protons is called an **anion** (negatively charged ion). Those elements in the oxygen group need two electrons for stability; thus, they have a charge of negative 2.

On the other end of the table, the alkali metals can most easily achieve a full valence shell by losing a single electron, rather than by gaining seven. So, they will form positively charged ions (**cations**), which have more protons than electrons. Alkali metals typically have a charge of +1, alkaline Earth metals of +2, and so forth. The table below gives a good overview of the common ionic charges for various groups:

It's tempting to say that fluorine atoms "want" one electron or are "happy" with a full valence shell, but in reality, atoms don't have feelings and don't want anything! Instead, say fluorine atoms need only attract one electron to have a stable energy level. Assigning feelings and desires to inanimate particles won't do you any favors on the exam!



In general, transition metals can form ions of varying charges. All transition metals lose electrons to form cations, but how many electrons they lose will vary depending on the compound they are in. For example, in $CuBr_2$, we know what the charge on the bromide is negative one, and with two bromides the total negative charge present in the compound is –2. The total compound charge must be zero for the compound to be stable, so it stands to reason that the charge on the single copper ion is +2. That compound would be called copper (II) bromide. However, in CuBr, the charge on the copper would only need to be +1 to balance the single bromide, so that compound would be called copper (I) bromide.

There are a few transition metals that only form ions with one possible charge. Two important examples are zinc, which always forms ions with a charge of +2, and silver, which always forms ions with a charge of +1.

There's also a slightly different rule as to how transition metals lose electrons when forming those cations. Transition metals, when losing electrons, will lose their higher-level *s* electrons before losing any of the lower-level *d* electrons. Thus, while an iron atom has a configuration of $[Ar]4s^23d^6$, an iron ion that has lost two electrons (Fe^{2+}) would have a configuration of $[Ar]3d^6$. If it were to lose further electrons, those would then come from the *d*-orbital.

Names and Theories

Dalton's Elements

In the early 1800s, John Dalton presented some basic ideas about atoms that we still use today. He was the first to say that there are many different kinds of atoms, which he called elements. He said that these elements combine to form compounds and that these compounds always contain the same ratios of elements. Water (H_2O), for instance, always has two hydrogen atoms for every oxygen atom. He also said that atoms are never created or destroyed in chemical reactions.

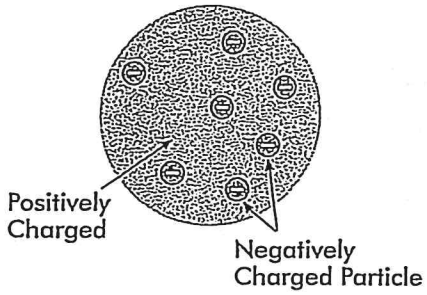
Development of the Periodic Table

In 1869, Dmitri Mendeleev and Lothar Meyer independently proposed arranging the elements into early versions of the periodic table, based on the trends of the known elements.

ıomson's Experiment

. the late 1800s, J. J. Thomson watched the deflection of charges in a cathode y tube and put forth the idea that atoms are composed of positive and negative .arges. The negative charges were called electrons, and Thomson guessed that ey were sprinkled throughout the positively charged atom like chocolate chips rinkled throughout a blob of cookie dough.

The Plum Pudding Model of an Atom

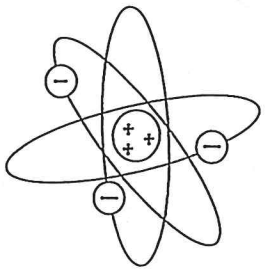


illikan's Experiment

ıbert Millikan was able to calculate the charge on an electron by examining the ıavior of charged oil drops in an electric field.

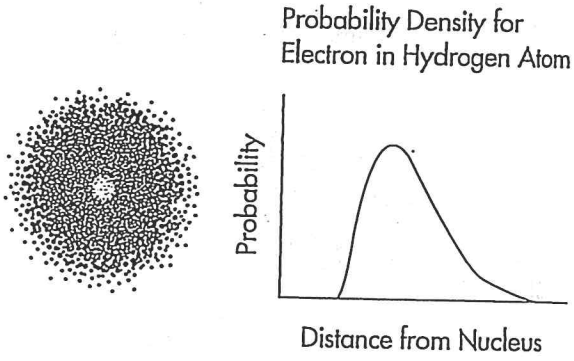
therford's Experiment

the early 1900s, Ernest Rutherford fired alpha particles at gold foil and observed w they were scattered. This experiment led him to conclude that all of the posi- ɛ charge in an atom was concentrated in the center and that an atom is mostly pty space. This led to the idea that an atom has a positively charged nucleus, ich contains most of the atom's mass, and that the tiny, negatively charged elec- ns travel around this nucleus.



The Heisenberg Uncertainty Principle

Werner Heisenberg said that it is impossible to know both the position and momentum of an electron at a particular instant. In terms of atomic structure, this means that electron orbitals do not represent specific orbits like those of planets. Instead, an electron orbital is a probability function describing the possibility that an electron will be found in a region of space.



PERIODIC TRENDS

You can make predictions about certain behavior patterns of an atom and its electrons based on the position of the atom in the periodic table. All the periodic trends can be understood in terms of three basic rules.

1. Electrons are attracted to the protons in the nucleus of an atom.
   a. The closer an electron is to the nucleus, the more strongly it is attracted.
   b. The more protons in a nucleus, the more strongly an electron is attracted.
2. Electrons are repelled by other electrons in an atom. So, if other electrons are between a valence electron and the nucleus, the valence electron will be less attracted to the nucleus. That's called shielding.
3. Completed shells (and to a lesser extent, completed subshells) are very stable. Atoms will add or subtract valence electrons to create complete shells if possible.

The atoms in the left-hand side of the periodic table are called metals. Metals give up electrons when forming bonds. Most of the elements in the table are metals. The elements in the upper right-hand portion of the table are called nonmetals. Nonmetals generally gain electrons when forming bonds. The metallic character of the elements decreases as you move from left to right across the periodic table. The elements in the borderline between metal and nonmetal, such as silicon and arsenic, are called metalloids.

Atomic Radius

The atomic radius is the approximate distance from the nucleus of an atom to its valence electrons.

Moving from Left to Right Across a Period (Li to Ne, for Instance), Atomic Radius Decreases

Moving from left to right across a period, protons are added to the nucleus, so the valence electrons are more strongly attracted to the nucleus; this decreases the atomic radius. Electrons are also being added, but they are all in the same shell at about the same distance from the nucleus, so there is not much of a shielding effect.

Moving Down a Group (Li to Cs, for Instance), Atomic Radius Increases

Moving down a group, shells of electrons are added to the nucleus. Each shell shields the more distant shells from the nucleus and the valence electrons get farther away from the nucleus. Protons are also being added, but the shielding effect of the negatively charged electron shells cancels out the added positive charge.

Cations (Positively Charged Ions) Are Smaller than Atoms

Generally, when electrons are removed from an atom to form a cation, the outer shell is lost, making the cation smaller than the atom. Also, when electrons are removed, electron–electron repulsions are reduced, allowing all of the remaining valence electrons to move closer to the nucleus.

Anions (Negatively Charged Ions) Are Larger than Atoms

When an electron is added to an atom, forming an anion, electron–electron repulsions increase, causing the valence electrons to move farther apart, which increases the radius.

Ionization Energy

Electrons are attracted to the nucleus of an atom, so it takes energy to remove an electron. The energy required to remove an electron from an atom is called the first ionization energy. Once an electron has been removed, the atom becomes a positively charged ion. The energy required to remove the next electron from the ion is called the second ionization energy, and so on.

Moving from Left to Right Across a Period, Ionization Energy Increases

Moving from left to right across a period, protons are added to the nucleus, which increases its positive charge. For this reason, the negatively charged valence electrons are more strongly attracted to the nucleus, which increases the energy required to remove them.

Moving Down a Group, Ionization Energy Decreases

Moving down a group, shells of electrons are added to the nucleus. Each inner shell shields the more distant shells from the nucleus, reducing the pull of the nucleus on the valence electrons and making them easier to remove. Protons are also being added, but the shielding effect of the negatively charged electron shells cancels out the added positive charge.

The Second Ionization Energy Is Greater than the First Ionization Energy

When an electron has been removed from an atom, electron–electron repulsion decreases and the remaining valence electrons move closer to the nucleus. This increases the attractive force between the electrons and the nucleus, increasing the ionization energy.

As Electrons Are Removed, Ionization Energy Increases Gradually Until a Shell Is Empty, Then Makes a Big Jump

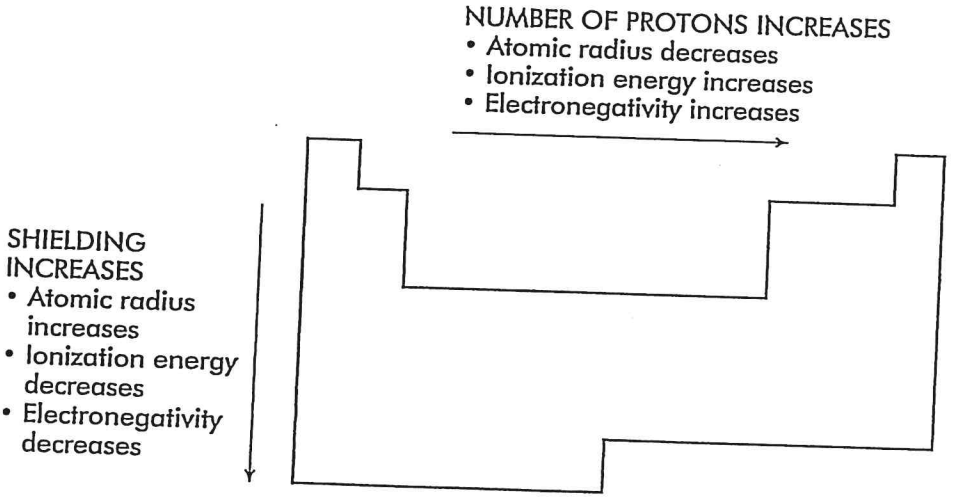
- For each element, when the valence shell is empty, the next electron must come from a shell that is much closer to the nucleus, making the ionization energy for that electron much larger than for the previous ones.
- For Na, the second ionization energy is much larger than the first.
- For Mg, the first and second ionization energies are comparable, but the third is much larger than the second.
- For Al, the first three ionization energies are comparable, but the fourth is much larger than the third.

Electronegativity

Electronegativity refers to how strongly the nucleus of an atom attracts the electrons of other atoms in a bond. Electronegativity is affected by two factors. The smaller an atom is, the more effectively its nuclear charge will be felt past its outermost energy level and the higher its electronegativity will be. Second, the closer an element is to having a full energy level, the more likely it is to attract the necessary electrons to complete that level. In general:

- Moving from left to right across a period, electronegativity increases.
- Moving down a group, electronegativity decreases.

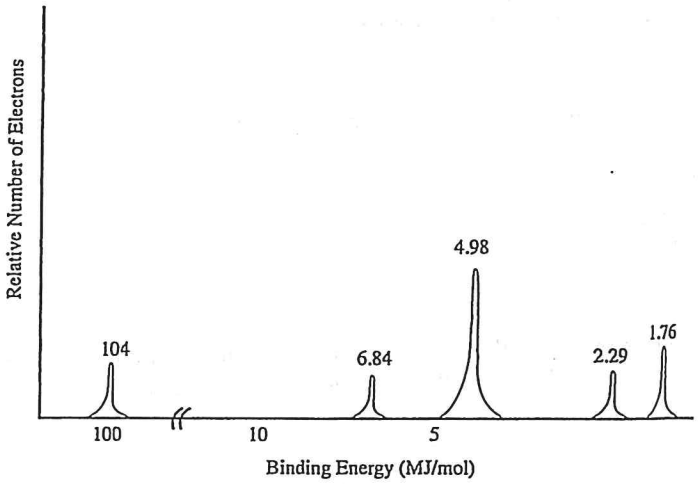
The various periodic trends are summarized in the diagram below. The primary exception to these trends are the electronegativity values for the three smallest noble gases. As helium, neon, and argon do not form bonds, they have zero electronegativity. The larger noble gases, however, can form bonds under certain conditions and do follow the general trends as outlined in this section. This will be discussed in more detail in the next chapter.



CHAPTER 3 QUESTIONS

Multiple-Choice Questions

Use the PES spectra below to answer questions 1-4.



1. What element does this spectra represent?

   (A) Boron
   (B) Nitrogen
   (C) Aluminum
   (D) Phosphorus

2. Which peak represents the 2*s* subshell?

   (A) The peak at 104 MJ/mol
   (B) The peak at 6.84 MJ/mol
   (C) The peak at 2.29 MJ/mol
   (D) The peak at 1.76 MJ/mol

3. An electron from which peak would have the greatest velocity after ejection?

   (A) The peak at 104 MJ/mol
   (B) The peak at 6.84 MJ/mol
   (C) The peak at 4.98 MJ/mol
   (D) The peak at 1.76 MJ/mol

4. How many valence electrons does this atom have?

   (A) 2
   (B) 3
   (C) 4
   (D) 5

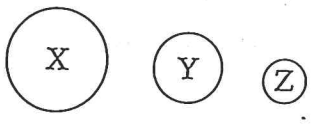
5. Why does an ion of phosphorus, P^{3-}, have a larger radius than a neutral atom of phosphorus?

   (A) There is a greater Coulombic attraction between the nucleus and the electrons in P^{3-}.
   (B) The core electrons in P^{3-} exert a weaker shielding force than those of a neutral atom.
   (C) The nuclear charge is weaker in P^{3-} than it is in P.
   (D) The electrons in P^{3-} have a greater Coulombic repulsion than those in the neutral atom.

6. Which neutral atom of the following elements would have the most unpaired electrons?

   (A) Titanium
   (B) Manganese
   (C) Nickel
   (D) Zinc

7. The diagram below shows the relative atomic sizes of three different elements from the same period. Which of the following statements must be true?



   (A) The effective nuclear charge will be the greatest in element X.
   (B) The first ionization energy will be greatest in element X.
   (C) The electron shielding effect will be greatest in element Z.
   (D) The electronegativity value will be greatest in element Z.

8. The first ionization energy for a neutral atom of chlorine is 1.25 MJ/mol and the first ionization energy for a neutral atom of argon is 1.52 MJ/mol. How would the first ionization energy value for a neutral atom of potassium compare to those values?

   (A) It would be greater than both because potassium carries a greater nuclear charge then either chlorine or argon.
   (B) It would be greater than both because the size of a potassium atom is smaller than an atom of either chlorine or argon.
   (C) It would be less than both because there are more electrons in potassium, meaning they repel each other more effectively and less energy is needed to remove one.
   (D) It would be less than both because a valence electron of potassium is farther from the nucleus than one of either chlorine or argon.

9. Which of the following isoelectric species has the smallest radius?

   (A) S^{2-}
   (B) Cl^{-}
   (C) Ar
   (D) K^{+}

10. What is the most likely electron configuration for a sodium ion?

   (A) $1s^2 2s^2 2p^5$
   (B) $1s^2 2s^2 2p^6$
   (C) $1s^2 2s^2 2p^6 3s^1$
   (D) $1s^2 2s^2 2p^5 3s^2$

11. Which of the following statements is true regarding sodium and chlorine?

(A) Sodium has greater electronegativity and a larger first ionization energy.
(B) Sodium has a larger first ionization energy and a larger atomic radius.
(C) Chlorine has a larger atomic radius and a greater electronegativity.
(D) Chlorine has greater electronegativity and a larger first ionization energy.

12. An atom of silicon in its ground state is subjected to a frequency of light that is high enough to cause electron ejection. An electron from which subshell of silicon would have the highest kinetic energy after ejection?

(A) $1s$
(B) $2p$
(C) $3p$
(D) $4s$

13. The wavelength range for infrared radiation is 10^{-5} m, while that of ultraviolet radiation is 10^{-8} m. Which type of radiation has more energy, and why?

(A) Ultraviolet has more energy because it has a higher frequency.
(B) Ultraviolet has more energy because it has a longer wavelength.
(C) Infrared has more energy because it has a lower frequency.
(D) Infrared has more energy because it has a shorter wavelength.

14. Which of the following nuclei has 3 more neutrons than protons? (Remember: The number before the symbol indicates atomic mass.)

(A) ^{11}B
(B) ^{37}Cl
(C) ^{24}Mg
(D) ^{70}Ga

15. Examining data obtained from mass spectrometry supports which of the following?

(A) The common oxidation states of elements
(B) Atomic size trends within the periodic table
(C) Ionization energy trends within the periodic table
(D) The existence of isotopes

16. In general, do metals or nonmetals from the same period have higher ionization energies? Why?

(A) Metals have higher ionization energies because they usually have more protons than nonmetals.
(B) Nonmetals have higher ionization energies because they are larger than metals and harder to ionize.
(C) Metals have higher ionization energies because there is less electron shielding than there is in nonmetals.
(D) Nonmetals have higher ionization energies because they are closer to having filled a complete energy level.

17. The ionization energies for an element are listed in the table below.

First	Second	Third	Fourth	Fifth
8 eV	15 eV	80 eV	109 eV	141 eV

Based on the ionization energy table, the element is most likely to be

(A) Sodium
(B) Magnesium
(C) Aluminum
(D) Silicon

Use the following information to answer questions 18-20.

The outermost electron of an atom has a binding energy of 2.5 eV. The atom is exposed to light of a high enough frequency to cause exactly one electron to be ejected. The ejected electron is found to have a KE of 2.0 eV.

18. How much energy did photons of the incoming light contain?
    (A) 0.50 eV
    (B) 0.80 eV
    (C) 4.5 eV
    (D) 5.0 eV

19. If the wavelength of the light were to be shortened, how would that effect the kinetic energy of the ejected electron?
    (A) A shorter wavelength would increase the kinetic energy.
    (B) A shorter wavelength would decrease the kinetic energy.
    (C) A shorter wavelength would stop all electron emissions completely.
    (D) A shorter wavelength would have no effect on the kinetic energy of the ejected electrons.

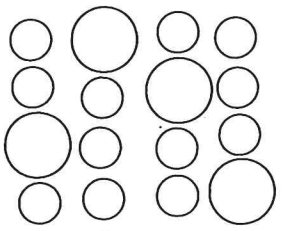
20. If the intensity of the light were to be decreased (that is, if the light is made dimmer), how would that affect the kinetic energy of the ejected electron?
    (A) The decreased intensity would increase the kinetic energy.
    (B) The decreased intensity would decrease the kinetic energy.
    (C) The decreased intensity would stop all electron emissions completely.
    (D) The decreased intensity would have no effect.

---

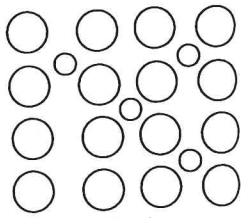
21. Which type of radiation would be most useful in examining the dimensionality of molecules?
    (A) Ultraviolet
    (B) Visible
    (C) Infrared
    (D) Microwave

22. Which of the following ions would have the most unpaired electrons?
    (A) Mn^{2+}
    (B) Ni^{3+}
    (C) Ti^{2+}
    (D) Cr^{6+}

23. Two alloys are shown in the diagrams below—bronze and steel. Which of the following correctly describes the malleability of both alloys compared to their primary metals?



Bronze



Steel

○ = Cu

● = Sn

○ = Fe

○ = C

(A) Bronze's malleability would be comparable to that of copper, but steel's malleability would be significantly lower than that of iron.

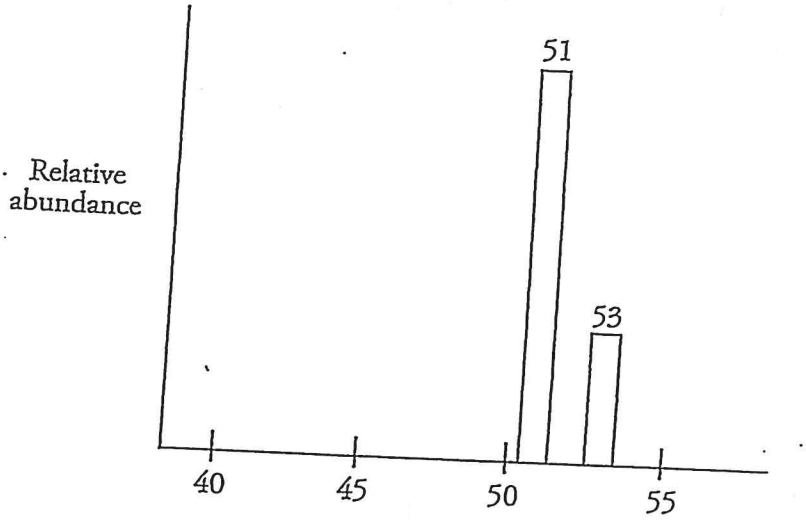
(B) Bronze's malleability would be significantly higher than that of copper, but steel's malleability would be comparable to that of iron.

(C) Both bronze and steel would have malleability values similar to those of their primary metals.

(D) Both bronze and steel would have malleability values lower than those of their primary metals.

Free-Response Questions

1. Explain each of the following in terms of atomic and molecular structures and/or forces.
   (a) The first ionization energy for magnesium is greater than the first ionization energy for calcium.
   (b) The first and second ionization energies for calcium are comparable, but the third ionization energy is much greater.
   (c) There are three peaks of equal height in the PES of carbon, but on the PES of oxygen the last peak has a height twice as high as all the others.
   (d) The first ionization energy for aluminum is lower than the first ionization energy for magnesium.



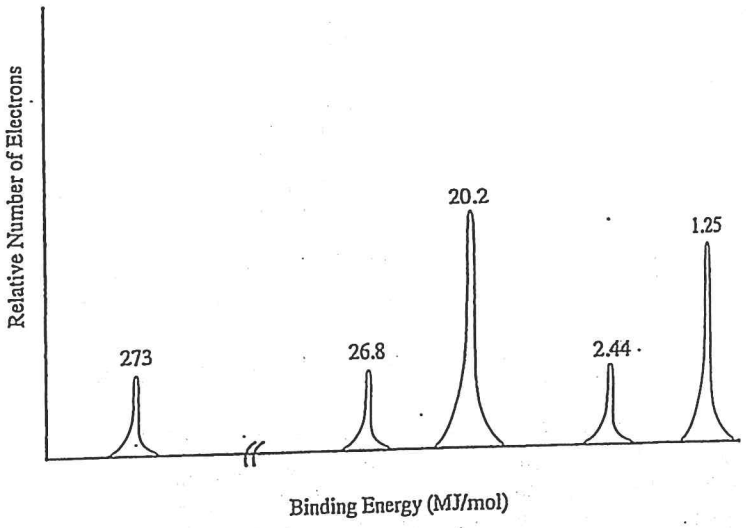
2. The above mess spectra is for the hypochlorite ion, ClO^-. Oxygen has only one stable isotope, which has a mass of 16 amu.
   (a) How many neutrons does the most common isotope of chlorine have?
   (b) Using the spectra, calculate the average mass of a hypochlorite ion.
   (c) Does the negative charge on the ion affect the spectra? Justify your answer.
   (d) The negative charge in the ion is located around the oxygen atom. Speculate as to why.

3. The table below gives data on four different elements, in no particular order:

Carbon, Oxygen, Phosphorus, and Chlorine

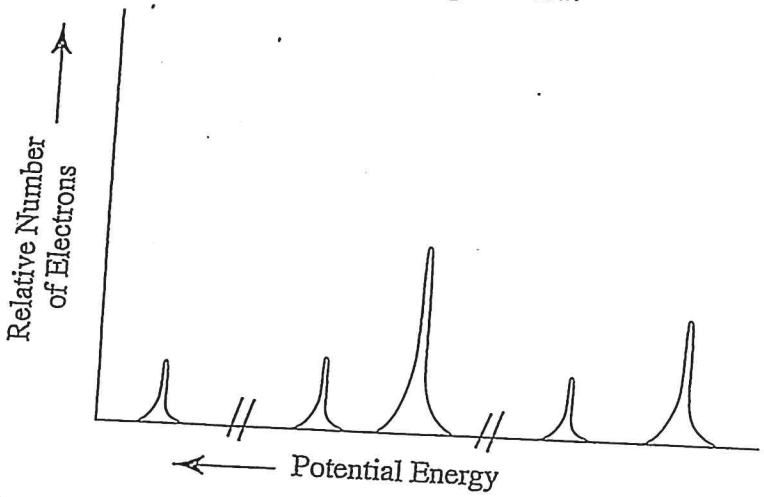
	Atomic radius (pm)	First Ionization Energy (kJ/mol^{-1})
Element 1	170	1086.5
Element 2	180	1011.8
Element 3	175	1251.2
Element 4	152	1313.9

(a) Which element is number 3? Justify your answer using both properties.
(b) What is the outermost energy level that has electrons in element 2? How many valence electrons does element 2 have?
(c) Which element would you expect to have the highest electronegativity? Why?
(d) How many peaks would the PES for element 4 have and what would the relative heights of those peaks be to each other?



4. The above PES belongs to a neutral chlorine atom.
   (a) What wavelength of light would be required to eject a 3*s* electron from chlorine?
   (b) For the PES of a chloride ion, how would the following variables compare to the peaks on the PES above? Justify your answers.
      (i) Number of peaks
      (ii) Height of the peaks

5. The photoelectron spectrum of an element is given below:



(a) Identify the element this spectra most likely belongs to and write out its full electron configuration.

(b) Using your knowledge of atomic structure, explain the following:
   (i) The reason for the three discrete areas of ionization energies
   (ii) The justification for there being a total of five peaks
   (iii) The relative heights of the peaks when compared to each other