Solution Manual for Introduction to Chemistry 3rd Edition Bauer Birk Marks 0073402672 9780073402673

Chapter 2 – Atoms, Ions, and the Periodic Table

2.1 (a) neutron; (b) law of conservation of mass; (c) proton; (d) main-group element; (e) relative atomic mass; (f) mass number; (g) isotope; (h) cation; (i) subatomic particle; (j) alkali metal; (k) periodic table

2.2 (a) transition element; (b) law of definite proportions; (c) electron; (d) anion; (e) diatomic molecule; (f) noble gas; (g) period; (h) atomic number; (i) atomic mass unit (amu); (j) group or family; (k) alkaline earth metal; (l) nucleus

2.3 Dalton used the laws of conservation of mass (Lavoisier) and definite proportions (Proust). Dalton essentially reasoned that, because pure substances were always composed of elements in some fixed ratios, matter must be composed of discrete units (atoms).

2.4 The scanning tunneling microscope (STM) is used to image atoms on the surface of a material (Figure 2.4).

2.5 The second postulate of Dalton's atomic theory, listed in section 2.1, states that atoms of different elements differ in their atomic masses and chemical properties.

2.6 The third and fourth postulates of Dalton's atomic theory, listed in section 2.1, can be used to explain conservation of mass. In essence, because atoms only rearrange *and* no new matter can be formed during chemical reactions, mass must be conserved. If elements could be changed into other elements during chemical reactions (as the alchemists were trying to do), then masses of atoms would change during reactions and mass would not be conserved.

2.7 Compounds contain discrete numbers of atoms of each element that form them. Because all the atoms of an element have the same relative atomic mass, the mass ratio of the elements in a compound is always the same (law of definite proportions).

2.8 The fourth postulate of Dalton's atomic theory (section 2.1) states that compounds are formed when elements combine in simple, whole-number ratios. In addition, for any pure substance that ratio must be fixed.

2.9 No. On the left side of the diagram there are four white atoms and two blue atoms (represented as diatomic molecules). On the right, however, there are six white atoms. Since the numbers of each atom are not conserved, mass is not conserved. For mass to be conserved, another white molecule is needed on the left side of the diagram.

2.10 Yes. There are four white atoms and four green atoms on the left side. On the right, the atoms have recombined to form new substances, but the numbers of each atom are the same. This means mass has been conserved.

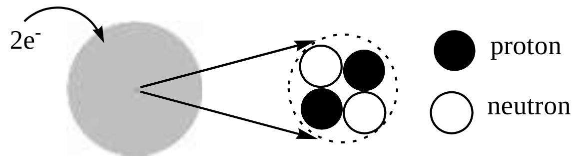
2.11 Thomson's cathode ray experiment.

2.12 Rutherford's gold foil experiment (Figure 2.8) revealed the nucleus of the atom.

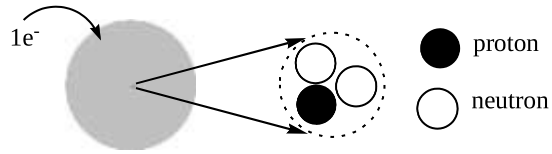
2.13 The electron is the subatomic particle with a negative charge. A summary of the subatomic particles and their properties is given in Table 2.1.

2.14 The proton is the subatomic particle with a positive charge. A summary of the subatomic particles and their properties is given in Table 2.1.

2.15 The nucleus of helium has two protons and two neutrons. Two electrons can be found in the cloud surrounding the nucleus.



2.16 The nucleus of hydrogen-3 ($^{3}_{1}H$) has one proton and two neutrons. Two electrons can be found in the cloud surrounding the nucleus:



2.17 The neutron and proton (Table 2.1) have approximately the same atomic mass (1 amu).

2.18 From Table 2.1 we find the mass of the proton and electron and calculate the ratio of their masses. From the calculation we find that the proton is about 2000 times the mass of the electron.

$$\frac{\text{proton mass}}{\text{electron mass}} = \frac{1.6726 \times 10^{-24}\,\text{g}}{9.1094 \times 10^{-28}\,\text{g}} = 1836.1$$

2.19 Carbon atoms have 6 protons. The relative atomic mass of a carbon atom is 12.01 amu indicating the presence of 6 neutrons. Protons and neutrons have approximately equal masses so the nuclear mass is approximately two times the mass of the protons.

2.20 Electrons and protons were discovered first because their deflection by electric and magnetic fields was relatively easy to detect. Neutrons were much more difficult to detect because they have no charge.

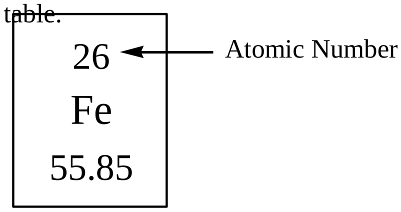
2.21 The atomic number (protons) is given on the periodic table. (a) 1 (element symbol: H); (b) 8 (element symbol: O); (c) 47 (element symbol: Ag)

2.22 The number of protons (atomic number) is given on the periodic table. (a) 2 (element symbol: He); (b) 18 (element symbol: Ar); (c) 82 (element symbol: Pb)

2.23 The number of protons determines the identity of an element.

2.24 The principle difference in isotopes of an element is the number of neutrons. The different numbers of neutrons causes the mass and mass number of each isotope to differ.

2.25 The atomic number of an atom is equal to the number of protons. If you know the name of the element, you can find the atomic number by finding the element on the periodic table. For example, for iron (Fe), you can find the atomic number, 26, listed with the element symbol in the fourth period of the periodic table.



2.26 The mass number (A) is the sum of the neutrons (N) and protons (Z): $A = N + Z$. For example, an isotope of boron (symbol B) has 6 neutrons. Since the atomic number of boron is 5, the mass number for that isotope is $A = 6 + 5 = 11$.

2.27 All atoms of an element have the same number of protons and electrons. Only (b), atomic number, is the same for different isotopes of an element. The mass number, neutron number, and mass of an atom are different for each isotope of an element.

2.28 The mass number (a), neutron number (c), and mass of an atom (d) are different for isotopes of an element. The atomic number is always the same for each isotope of an element.

2.29 The following table displays the atomic, neutron, and mass numbers for the isotopes of hydrogen:

	${}^{1}_{1}H$	${}^{2}_{1}H$	${}^{3}_{1}H$
Atomic number	1	1	1
Neutron number	0	1	2
Mass number	1	2	3

2.30 There are eight positively charged particles (protons) and ten neutral particles (neutrons). The identity of the element is determined by the atomic number (number of protons). Since the atomic number is 8, this element is oxygen, O.

	Equal to	Value for this isotope
Atomic number	number of protons	8
Neutron number	number of neutrons	10
Mass number	neutrons + protons	18

2.31 The mass number (A) is defined as the number of protons (Z) plus the number of neutrons (N), $A = Z + N$. To find the number of protons (Z) in the nucleus (atomic number), we need to find the element on the periodic table. This allows us to calculate the neutron number (N) using: $N = A – Z$. For example, the element oxygen (Ar) has an atomic number of 18. For an oxygen isotope with a mass number of 36, the number of neutrons is: $N = 36 – 18 = 18$.

(a) $Z = 18$, $A = 36$, and $N = 18$
(b) $Z = 18$, $A = 38$, and $N = 20$
(c) $Z = 18$, $A = 40$, and $N = 22$

2.32 There are four neutrons (red) and three protons (blue). Because there are three protons, this is an isotope of the element lithium. The atomic number is 3 and the neutron number is 4. You get the mass number by adding the neutrons and protons: $A = N + Z = 7$.

2.33 To find the number of protons (Z) in the nucleus (atomic number), we need to find the element on the periodic table. This allows us to calculate the number of neutrons (N) using $N = A – Z$, where A is the mass number. For example, the element oxygen (O) has an atomic number of 8. For an oxygen isotope with a mass number of 15, the number of neutrons is $N = 15 – 8 = 7$.

In an atom the number of protons is equal to the number of electrons.

	Protons	Neutrons	Electrons
(a) ${}^{15}_{8}O$	8	7	8
(b) ${}^{109}_{47}Ag$	47	62	47
(c) ${}^{35}_{17}Cl$	17	18	17

2.34 To find the number of protons (Z) in the nucleus (atomic number), we need to find the element on the periodic table. This allows us to calculate the number of neutrons (N) using $N = A - Z$, where Z is the mass number. For example, the element hydrogen (H) has an atomic number of 1. For a hydrogen isotope with a mass number of 1, the number of neutrons is $N = 1 - 1 = 0$.

In an atom the number of protons is equal to the number of electrons.

	Protons	Neutrons	Electrons
(a) ${}^{1}_{1}\mathrm{H}$	1	0	1
(b) ${}^{26}_{12}\mathrm{Mg}$	12	14	12
(c) ${}^{6}_{3}\mathrm{Li}$	3	3	3

2.35 Isotope symbols have the general format ${}^{\text{mass number}}_{\text{protons}}X^{\text{charge}}$ or ${}^{A}_{Z}X$ where A is the mass number (neutrons plus protons), Z is the atomic number (number of protons), and X is the element symbol.

(a) From the periodic table you find that the element with an atomic number of 1 is hydrogen, H. Since the isotope has 2 neutrons, the mass number is 3. The isotope symbol for hydrogen-3 is ${}^{3}_{1}\mathrm{H}$.

(b) The element with an atomic number of 4 is beryllium, Be. Since the isotope has 5 neutrons, the mass number is 9. The isotope symbol for beryllium-9 is ${}^{9}_{4}\mathrm{Be}$.

(c) The element with an atomic number of 15 is phosphorus, P. Since the isotope has 16 neutrons, the mass number is 31. The isotope symbol for phosphorus-31 is ${}^{31}_{15}\mathrm{P}$.

2.36 Isotope symbols have the general format ${}^{\text{mass number}}_{\text{protons}}X^{\text{charge}}$ or ${}^{A}_{Z}X$ where A is the mass number (neutrons plus protons), Z is the atomic number (number of protons), and X is the element symbol.

(a) From the periodic table you find that the element with an atomic number of 2 is helium, He. Since the isotope has 1 neutron, the mass number is 3. The isotope symbol for helium-3 is ${}^{3}_{2}\mathrm{He}$.

(b) The element with an atomic number of 47 is silver, Ag. Since the isotope has 62 neutrons the mass number is 109. The isotope symbol for silver-109 is ${}^{109}_{47}\mathrm{Ag}$.

(c) The element with an atomic number of 82 is lead, Pb. Since the isotope has 125 neutrons the mass number is 207. The isotope symbol for lead-207 is ${}^{207}_{82}\mathrm{Pb}$.

2.37 The atomic number is determined from the isotope symbol or by finding the element on the periodic table. For example, copper (Cu) has an atomic number of 29. The number of neutrons in an atom of copper-65 is $N = 65 - 29 = 36$.

	Protons (Z)	Neutrons (N) ($N = A - Z$)
(a) ${}^{56}_{26}\mathrm{Fe}$	26	$N = 56 - 26 = 30$
(b) ${}^{39}\mathrm{K}$	19	$N = 39 - 19 = 20$
(c) copper-65	29	$N = 65 - 29 = 36$

2.38 The atomic number is determined from the isotope symbol or by finding the element on the periodic table. For example, copper (Cu) has an atomic number of 29. The number of neutrons in an atom of copper-65 is N = 65 – 29 = 36.

	Protons (Z)	Neutrons (N) ($N = A - Z$)
(a) ${}^{11}_{5}\mathrm{B}$	5	$N = 11 - 5 = 6$
(b) ${}^{68}\mathrm{Zn}$	30	$N = 68 - 30 = 38$
(c) iodine-127	53	$N = 127 - 53 = 74$

2.39 The mass number is given by: $A = Z + N$, where Z is the number of protons and N is the number of neutrons. From the periodic table we find that nitrogen's atomic number is 7 so there are 7 protons. Since the mass number is 13, the number of neutrons is $Z = A - N = 13 - 7 = 6$.

2.40 The mass number is given by: $A = Z + N$, where Z is the number of protons and N is the number of neutrons. From the periodic table we find that phosphorus' atomic number is 15 so there are 15 protons. Since the mass number is 32, the number of neutrons is $Z = A - N = 32 - 15 = 17$.

2.41 The isotope symbol takes the form $^{\text{mass number}}_{\text{protons}}X^{\text{charge}}$ where the mass number is the neutrons plus protons and the charge is determined by the protons and electrons. X is the element symbol. All atoms are electrically neutral so the number of electrons and protons are the same.

Isotope Symbol	Number of Protons	Number of Neutrons	Number of Electrons
$^{23}_{11}Na$	11	12	11
$^{56}_{25}Mn$	25	31	25
$^{18}_{8}O$	8	10	8
$^{19}_{9}F$	9	10	9

2.42 The isotope symbol takes the form $^{\text{mass number}}_{\text{protons}}X^{\text{charge}}$ where the mass number is the neutrons plus protons and the charge is determined by the protons and electrons. X is the element symbol. All atoms are electrically neutral so the number of electrons and protons are the same.

Isotope Symbol	Number of Protons	Number of Neutrons	Number of Electrons
$^{14}_{6}C$	6	8	6
$^{23}_{12}Mg$	12	11	12
$^{30}_{16}S$	16	14	16
$^{30}_{14}Si$	14	16	14

2.43 They differ in the number of electrons. The identity of an atom or an ion is determined by the number of protons in the nucleus. However, ions have different numbers of electrons than protons. This is why ions are charged. For example, the ion N^{3-} is similar to the N atom because it has 7 protons, but the ion has 10 electrons. The three "extra" electrons give the ion the 3– charge.

2.44 The number of electrons changes. For example, the ion Ba^{2+} and the element Ba both have 56 protons. However, the ion has two fewer electrons (54). Since there are 56 protons and 54 electrons in the ion, the charge is 2+. A change in the number of protons would change the identity of the atom.

2.45 (a) When an atom gains one electron, an anion with a 1– charge is formed. For example, when a fluorine atom, with 9 protons and 9 electrons, gains 1 electron, there are 10 negative charges and 9 positive charges. This means that the resulting ion will have a 1– charge (F^-).

(b) When an atom loses two electrons, a cation with a 2+ charge is formed. For example, when a magnesium atom, with 12 protons and 12 electrons, loses 2 electrons, there are 12 positive charges and 10 negative charges. This means that the resulting ion will have a 2+ charge (Mg^{2+}).

2.46 (a) An ion with a 1+ charge has one more proton than it does electrons. When the ion receives an electron, the number of protons and electrons will be the same and the neutral atom is formed.

(b) An ion with a 1+ charge has one more proton than it does electrons. By losing two electrons, the ion will have three more protons than it does electrons. An ion with a 3+ charge will be formed.

2.47 (a) Zinc atoms have 30 protons and 30 electrons. When two electrons are lost there are still 30 positive charges, but only 28 negative charges. The ion that results has a 2+ charge. The symbol for the ion is Zn^{2+}. Positive ions are called cations.

(b) Phosphorus atoms have 15 protons and 15 electrons. When a P atom gains three electrons there will be three more negative charges (18 electrons) than positive charges (15 protons). The resulting ion will have a 3– charge. The symbol for the ion is P^{3-}. Negative ions are called anions.

2.48 (a) Selenium atoms have 34 protons and 34 electrons. When an Se atom gains two electrons there are two more negative charges (36 electrons) than positive charges (34 protons). As a result, the ion has a 2– charge. The symbol for the ion is Se^{2-}. Negative ions are called anions.

(b) Mercury atoms have 80 protons and 80 electrons. When an Hg atom loses two electrons there are two fewer negative charges (78 electrons) than positive charges (80 protons). The charge of the ion is 2+. The symbol for the ion is Hg^{2+}. Positive ions are called cations.

2.49 The number of protons is determined from the atomic symbol and the periodic table. For example, zinc (Zn) has 30 protons. The number of electrons is determined by looking at the charge on the ion. A Zn^{2+} ion has two fewer negative charges than positive charges. This means that there are 28 electrons (number of electrons = 30 – 2 = 28).

	Number of Protons	Number of Electrons
(a) Zn^{2+}	30	28
(b) F^{-}	9	10
(c) H^{+}	1	0

2.50 The number of protons is determined from the atomic symbol and the periodic table. For example, phosphorus (P) has 15 protons. The number of electrons is determined by looking at the charge on the ion. The P^{3-} ion has three more negative charges than positive charges. This means that there are 18 electrons (number of electrons = 15 + 3 = 18).

	Number of Protons	Number of Electrons
(a) P^{3-}	15	18
(b) Al^{3+}	13	10
(c) O^{2-}	8	10

2.51 The completed table is shown below.

(a) The number of protons is 17 as indicated by the isotope symbol. Since the mass number, *A*, is 37, the number of neutrons is 20 ($N = A - Z$). Since the charge is 1–, there must be 18 electrons.

(b) The number of protons is 12, so the element is Mg. The mass number, 25, is the sum of the protons and neutrons. Since there are two more protons than electrons, the ion has a charge of 2+.

(c) The number of protons is 7, so the element is N. The mass number is 13 (sum of protons and neutrons). The charge is 3– because there are three more electrons than protons.

(d) Since the element is calcium, the number of protons is 20. Because the mass number is 40, the number of neutrons is also 20. Since the charge is 2+, there must be 18 electrons.

Isotope Symbol	Number of Protons	Number of Neutrons	Number of Electrons
$^{37}_{17}Cl^{-}$	17	20	18
$^{25}_{12}Mg^{2+}$	12	13	10
$^{13}_{7}N^{3-}$	7	6	10
$^{40}Ca^{2+}$	20	20	18

2.52 The completed table is shown below.

(a) The number of protons is 35 as indicated by the isotope symbol. Since the mass number, *A*, is 81, the number of neutrons is 46 ($N = A - Z$). Since the charge is 1–, there must be 36 electrons.

(b) The number of protons is 38, so the element is Sr. The mass number, 88, is the sum of the protons and neutrons. Since there are two more protons than electrons, the ion has a charge of 2+.

(c) The number of protons is 1, so the element is H. The mass number is 2, the sum of the protons and neutrons. Since there is 1 more electron than proton, the charge is 1–.

(d) Since the element is hydrogen, the number of protons is 1. Because the mass number is 1, there must not be any neutrons in the nucleus (mass number and atomic number are the same). Since the charge is 1+, there are no electrons.

Isotope Symbol	Number of Protons	Number of Neutrons	Number of Electrons
$^{81}_{35}Br^{-}$	35	46	36
$^{88}_{38}Sr^{2+}$	38	50	36
$^{2}_{1}H^{-}$	1	1	2
$^{1}H^{+}$	1	0	0

2.53 Since the ion has a charge of 1+, there must be one more proton than electrons. Since there are 18 electrons, there must be 19 protons. The element is potassium, K.

2.54 Since the charge is 2–, there must be two more electrons than protons. Since there are 18 electrons, there must be 16 protons. The element is sulfur, S.

2.55 Since the charge is 2+, there must be two more protons than electrons. Since there are 27 electrons, there must be 29 protons. The element is copper, Cu.

2.56 Since the charge is 1+, there must be one more proton than electrons. Since there are 46 electrons, there must be 47 protons. The element is silver, Ag.

2.57 Lithium-7, ^{7}Li, has three protons, three electrons, and four neutrons. $^{7}Li^{+}$ has only two electrons, and ^{6}Li has only three neutrons. Otherwise they are the same as ^{7}Li. Lithium-6 differs the most in mass.

Isotope	Protons (*Z*)	Neutrons (*N*)	Mass number (*A*) $A = N + Z$	Electrons
^{7}Li	3	4	7	3
$^{7}Li^{+}$	3	4	7	2
^{6}Li	3	3	6	3

2.58 The differences in ${}^{79}_{35}\text{Br}^{-}$ and ${}^{81}_{35}\text{Br}$ are highlighted in the table below:

Isotope	Protons (*Z*)	Neutrons (*N*)	Mass number (*A*) $A = N + Z$	Electrons
${}^{79}_{35}\text{Br}^{-}$	35	44	79	36
${}^{81}_{35}\text{Br}$	35	46	81	35

${}^{81}_{35}\text{Br}$ has the greater mass.

2.59 From the periodic table we find that potassium has 19 protons. Since it has a 1+ charge, there must be one more proton than electrons in the atom. There are 18 electrons.

2.60 From the periodic table we find that calcium has 20 protons. Since it has a 2+ charge, there must be two more protons than electrons in the atom. There are 18 electrons.

2.61 The mass of carbon-12 is defined as exactly 12 amu. From this, the atomic mass unit is defined as 1/12th the mass of one carbon-12 atom.

2.62 The mass of a carbon-12 atom is defined as exactly 12 amu.

2.63 The approximate mass of an isotope is equivalent to its mass number. This is true since most of the mass of an atom comes from the protons and neutrons in the nucleus. (a) 2 amu; (b) 238 amu

2.64 Since each cobalt-59 atom has an approximate mass of 59 amu, we can calculate the mass as follows:

$$\text{Mass in amu} = 10 \text{ atom} \times \frac{59 \text{ amu}}{\text{atom}} = 590 \text{ amu}$$

2.65 Deuterium, an isotope of hydrogen with 1 neutron and 1 proton, has a mass number of 2. A molecule of D_2 has a mass of 4 amu and a molecule of H_2 has a mass of 2 amu. Therefore, the mass of D_2 is 2 amu greater (or two times greater) than the mass of H_2.

2.66 Deuterium, an isotope of hydrogen with 1 neutron and 1 proton, has a mass number of 2. One molecule of D_2O will have a mass of approximately 20 amu. A molecule of H_2O has a mass of 18 amu. Therefore, the mass of D_2O is approximately 2 amu greater than H_2O.

2.67 The mass of an atom is approximately equal to its mass number. The mass of a krypton-80 atom is about 40 amu greater **or twice the mass** of an argon-40 atom.

2.68 The mass of an atom is approximately equal to its mass number. The mass of a magnesium-24 atom is approximately 12 amu greater **or twice** the mass of a carbon-12 atom.

2.69 The numerical values of the masses of individual atoms are very small when measured on the gram scale. The size of the atomic mass unit allows us to make easier comparisons and calculations of masses of molecules.

2.70 One atomic mass unit is approximately 1.6606×10^{-24} g. The number of mass units in a gram is calculated as:

$$\text{Atomic mass units per gram} = \frac{1\text{amu}}{1.6606 \times 10^{-24} \text{ g}} = 6.0219 \times 10^{23} \text{ amu/g}$$

2.71 The mass number is the sum of the number of protons and neutrons in the nucleus and is always an integer value. The mass number, which is not the actual mass, is usually close to the actual mass because the proton and neutron weigh approximately 1 amu each. The mass of an atom is the actual measurement of how much matter is in the atom and is never exactly an integer value (except carbon-12).

2.72 The mass of carbon-12 is defined as exactly 12 amu.

2.73 A mass spectrometer is used to determine the mass of an atom (Figure 2.15). The mass number of an atom is the sum of the number of protons and neutrons.

2.74 The most likely mass is 61.9283461 amu because this number is closest to 62. The answer is not 62.0000000 amu because carbon-12 is the only isotope with a mass that is exactly the same as the mass number.

2.75 If there are only two isotopes, the relative mass will be closer to the mass of the isotope that is most abundant. Since the relative mass of calcium is 40.08 amu, we can assume that calcium-40 is the most abundant isotope.

2.76 The atomic mass for silicon is 28.09. Since there is only one isotope of silicon, the mass of the isotope will be close to its mass number. The symbol for silicon-28 is ${}^{28}_{14}\mathrm{Si}$.

2.77 To calculate the relative atomic mass we calculate the weighted average of the isotopes of X.

	Isotope mass	×	abundance	=	mass contribution from isotope
${}^{22}X$	21.995 amu	×	0.7500	=	16.50 amu
${}^{20}X$	19.996 amu	×	0.2500	=	5.00 amu
					21.50 amu (relative atomic mass of X)

2.78 To calculate the relative atomic mass we calculate the weighted average of the isotopes of Mg.

	Isotope mass	×	abundance	=	mass contribution from isotope
${}^{24}Mg$	23.985 amu	×	0.2000	=	4.797 amu
${}^{25}Mg$	24.985 amu	×	0.2000	=	4.997 amu
${}^{26}Mg$	25.983 amu	×	0.6000	=	15.59 amu
					25.38 amu (relative atomic mass of Mg)

2.79 (a) The tallest peak (nickel-58), with an abundance of 67.88%, is the most abundant isotope.
(b) The shortest peak (nickel-64), with an abundance of 1.08% is the least abundant isotope.
(c) The average mass will be closer to 58 because nickel-58 represents more than half of the stable isotopes.
(d) Each isotope has the same number of protons (28). Since each isotope has a 1+ charge, they must each have 27 electrons. The neutrons are obtained using $A = Z + N$ and solving for N:
nickel 58, $N = 30$; nickel-60, $N = 32$; nickel-61, $N = 33$; nickel-62, $N = 34$; nickel-64, $N = 36$.

2.80 (a) The tallest peak (magnesium-24), with an abundance of 78.70%, is the most abundant isotope.
(b) The shortest peak (magnesium-25), with an abundance of 10.13% is the least abundant isotope.
(c) The average mass will be closer to 24 because magnesium-24 represents more than half of the stable isotopes.
(d) Each isotope has the same number of protons (12). Since each isotope has a 1+ charge, they must each have 11 electrons. The neutrons are obtained using $A = Z + N$ and solving for N:
magnesium-24, $N = 12$; magnesium-25, $N = 13$; magnesium-26, $N = 14$.

2.81 The mass of 1000 boron atoms can be determined by multiplying the relative atomic mass by 1000.

$$\text{Total mass} = 1000\ \cancel{\text{atom}} \times \frac{10.81\ \text{amu}}{\cancel{\text{atom}}} = 10{,}810\ \text{amu or } 1.081 \times 10^4\ \text{amu}$$

2.82 The mass of 1000 mercury atoms can be determined by multiplying the relative atomic mass by 1000.

$$\text{Total mass} = 1000\ \cancel{\text{atom}} \times \frac{200.6\ \text{amu}}{\cancel{\text{atom}}} = 200{,}600\ \text{amu or } 2.006 \times 10^5\ \text{amu}$$

2.83 Since the relative atomic mass of mercury (200.6 amu) is much higher than that of boron (10.81 amu) there will be more boron atoms in 2500 amu (i.e. it takes more of them to add up to 2500). This can be demonstrated by calculating the number of atoms:

$$\text{Boron atoms} = 2500\ \cancel{\text{amu}} \times \frac{1\text{atom}}{10.81\ \cancel{\text{amu}}} = 231\ \text{atoms}$$

$$\text{Mercury atoms} = 2500\ \cancel{\text{amu}} \times \frac{1\text{atom}}{200.6\ \cancel{\text{amu}}} = 12\ \text{atoms}$$

2.84 Silver has a relative atomic mass of (107.9 amu) which is less than that of gold (197.0 amu). If you have equal masses of the metals, there would be more atoms in the sample of silver.

2.85 Most elements can be classified in multiple ways. Metals, metalloids, and nonmetals are distinguished in Figure 2.18. Many of the groups of the periodic tables have unique names.

Group	Name
“A” block	main group elements
“B” block	transition elements
IA (1)	alkali metals
IIA (2)	alkaline earth metals
VIIA (17)	halogens
VIIIA (18)	noble gases

(a) alkali metal: K
(b) halogen: Br
(c) transition metal: Mn
(d) alkaline earth metal: Mg
(e) noble gas: Ar
(f) main-group element: Br, K, Mg, Al, Ar

2.86 Most elements can be classified in multiple ways. Metals, metalloids, and nonmetals are distinguished in Figure 2.18. Many of the groups of the periodic tables have unique names.
(a) alkali metals: Na
(b) halogen: I
(c) transition metal: Zn
(d) alkaline earth metal: Ca
(e) noble gas: He
(f) main-group element: He, I, Ca, Na, Pb

2.87 Halogens are found in Group VIIA (17). The halogen found in period 3 is chlorine, Cl.

2.88 Alkaline earth metals are found in group IIA (2). The alkaline earth metal in period 5 is strontium, Sr.

2.89 Titanium, Ti, is found at the intersection of period 4 and group IVB (4).

2.90 Carbon, C, is found at the intersection of period 2 and group IVA (14).

2.91 The metals and nonmetals are shown in Figure 2.18. (a) Ca, metal; (b) C, nonmetal; (c) K, metal; (d) Si, metalloid

2.92 The metals and nonmetals are shown in Figure 2.18. (a) P, nonmetal; (b) Cr, metal; (c) As, metalloid; (d) Na, metal

2.93 The major groups of the elements are shown in Figure 2.18. (a) O, main group; (b) Mg, main group; (c) Sn, main group; (d) U, actinide; (e) Cr, transition metal

2.94 The major groups of the elements are shown in Figure 2.18. (a) S, main group; (b) Fe, transition metal; (c) Pu, actinide; (d) Ca, main group; (e) Xe, main group

2.95 Group VIIA (17), the halogen group, all occur as diatomic molecules (F_2, Cl_2, Br_2, I_2). Astatine has no stable isotopes, but is believed to form diatomic molecules. It is estimated that less than 1 oz of astatine exists in the earth's crust.

2.96 The seven elements that form diatomic compounds are: hydrogen, H_2; nitrogen, N_2; oxygen, O_2; fluorine, F_2; chlorine, Cl_2; bromine, Br_2; and iodine, I_2.

2.97 Neon does not occur as a diatomic molecule.

2.98 Sulfur does not occur as a diatomic molecule.

2.99 The noble gases, group VIIIA (18), all occur as gases of uncombined atoms.

2.100 The noble gases are sometimes called inert gases because they generally do not react to form compounds. The noble gases are unreactive and do not form ions because of the stability associated with the number of electrons they contain.

2.101 electrons

2.102 Noble gases, Group VIIIA (18), do not form ions.

2.103 Many main group elements gain or lose electrons to form ions with the same number of electrons as the closest noble gas. In general, metals lose electrons to form positively charged ions (cations) and nonmetals gain electrons to form negatively charged ions (anions).

(a) The atoms of the elements in group IA (1) lose one electron to form 1+ ions. For example, sodium atoms have 11 electrons. After losing one electron, a sodium ion will have a 1+ charge and the same number of electrons (10) as neon.
(b) The atoms of the elements of group IIA (2) lose two electrons to form 2+ ions. For example, the ion Ca^{2+} has the same number of electrons (18) as argon.
(c) The atoms of the elements of group VIIA (17) gain one electron to form 1– ions. For example, the ion I^- has the same number of electrons (54) as xenon.
(d) The atoms of the elements of group VIA (16) gain two electrons to form 2– ions. For example, the ion Se^{2-} has the same number of electrons (36) as krypton.

2.104 Many main group elements gain or lose electrons to form ions with the same number of electrons as the closest noble gas. In general, metals lose electrons to form positively charged ions (cations) and nonmetals gain electrons to form negatively charged ions (anions).

(a) Atoms of the elements of group VIA (16) will gain two electrons to form 2– ions.
(b) The halogens, group VIIA (17), gain one electron to form 1– ions.
(c) The alkali metals, group IA (1), lose one electron to form 1+ ions.
(d) The alkaline earth metals, group IIA (2) lose two electrons to form 2+ ions.

2.105 Metals will lose electrons to have the same number of electrons as the nearest noble gas. This means that they will form positively charged ions. Nonmetals gain electrons to have the same number of electrons as the nearest noble gas. Therefore, nonmetals form negatively charged ions.

(a) Sodium atoms, Na, lose one electron to match the number of electrons in neon. Since one electron is lost, the charge is 1+: Na^{+}.
(b) Oxygen atoms, O, gain two electrons to match the number of electrons in neon. Since two electrons are gained, the charge is 2–: O^{2-}.
(c) Sulfur atoms, S, gain two electrons to match the number of electrons in argon. Since two electrons are gained, the charge is 2–: S^{2-}.
(d) Chlorine atoms, Cl, gain one electron to match the number of electrons in argon. Since one electron is gained, the charge is 1–: Cl^{-}.
(e) Bromine atoms, Br, gain one electron to match the number of electrons in krypton. Since one electron is gained, the charge is 1–: Br^{-}.

2.106 Metals will lose electrons to have the same number of electrons as the nearest noble gas. This means that they will form positively charged ions. Nonmetals gain electrons to have the same number of electrons as the nearest noble gas, Therefore, nonmetals form negatively charged ions.
(a) Nitrogen atoms, N, gain three electrons to match the number of electrons in neon. Since three electrons are gained, the charge is 3–: N^{3-}.
(b) Phosphorus atoms, P, gain three electrons to match the number of electrons in argon. Since three electrons are gained, the charge is 3–: P^{3-}.
(c) Magnesium atoms, Mg, lose two electrons to match the number of electrons in neon. Since two electrons are lost, the charge is 2+: Mg^{2+}.
(d) Potassium atoms, K, lose one electron to match the number of electrons in argon. Since one electron is lost, the charge is 1+: K^{+}.
(e) Aluminum atoms, Al, lose three electrons to match the number of electrons in neon. Since three electrons are lost, the charge is 3+: Al^{3+}.

2.107 Reactivity is a periodic property. Since sodium is in group IA, you might expect other elements in the same group (lithium, potassium, rubidium, cesium and francium) to react the same way.

2.108 Reactivity is a periodic property. Since chlorine is in group VIIA, you might expect other elements in the same group (fluorine, bromine, iodine, and astatine) to react the same way.

2.109 The mass of oxygen added to form Fe_2O_3 causes an increase in the mass.

2.110 No. Since there are six white atoms before the reaction takes place and only four after, mass is not conserved.

2.111 To show that the data is in agreement with the law of definite proportions, the mass ratio of Zn/S must be calculated for each substance. If the ratio is the same for both substances, the law of definite proportions is obeyed. In the first sample of zinc sulfide we have a ratio of

$$\text{Zn/S} = \frac{67.1\text{ g zinc}}{32.9\text{ g sulfur}} = 2.04\text{ g Zn/g S}$$

For the second compound, we are given the mass of zinc and zinc sulfide, but must calculate the mass of the sulfur used by the reaction:

$$\text{Mass of sulfur} = 2.00\text{ g zinc sulfide} - 1.34\text{ g zinc} = 0.66\text{ g sulfur}$$

The Zn/S ratio is calculated as

$$\text{Zn/S} = \frac{1.34\text{ g zinc}}{0.66\text{ g sulfur}} = 2.0\text{ g Zn/g S}$$

The mass ratios of Zn/S are the same for the two samples (within the significant figures given).

2.112 Any 100-g sample of HgO will contain 92.6 g of mercury atoms and 7.4 g of oxygen atoms.

2.113 Electrons were discovered first because they have charge and were readily studied in cathode ray tubes. Thomson put cathode rays (beams of electrons) in magnetic and electric fields to determine the charge and mass-to-charge ratio of the electron.

2.114 (a) Thomson (1897) determined the mass-to-charge ratio of electrons using cathode ray tubes. In addition, he showed that electron beams were repelled by negatively charged plates, indicating that electrons have a negative charge.

(b) In 1909, Robert Millikan determined the charge of an electron by measuring the electric field required to suspend a charged oil drop. Building on Thomson's earlier data, he calculated the mass of an electron (9.1094×10^{-28} g).

(c) Rutherford, between the years of 1907 and 1911, postulated, and then experimentally showed, that the nucleus is a very small and dense region of the atom. By passing alpha particles through gold foil, he found that most of the particles went through undisturbed but that a very few bounced almost directly back at the alpha source. This could only have happened if the mass of the gold was concentrated in a very small region of space (otherwise more alpha particles would have bounced back).

2.115 Only nickel (Ni) has an atomic number of 28. Since its mass number is 60, the isotope must be nickel-60.

2.116 The mass number is the sum of the numbers of neutrons and protons. If the mass number was less than the atomic number, the neutron number would have to be negative, which is not possible.

2.117 Since potassium has 19 protons, potassium-39 must have 20 neutrons.

2.118 The most abundant isotope of hydrogen is hydrogen-1. Since the atomic number is 1 and the mass number is 1, there are no neutrons. In addition, atoms of all hydrogen isotopes have 1 electron, so the 1+ ion has no electrons either. In summary, the most common isotope of H^+ has 1 proton and no electrons or neutrons.

2.119 The most abundant isotopes of cobalt have masses greater than the masses of the most abundant isotopes of nickel. As a result, the relative atomic mass of cobalt is greater than that of nickel.

2.120 The only element this can be is hydrogen. The mass of tritium (hydrogen-3) is approximately three times the mass of hydrogen-1 (the most common isotope of hydrogen).

2.121 When there are many isotopes, some of the isotopes can be present in very low abundance. As a result, their masses cannot be determined as accurately, and their percentage contributions are also less well known. Both factors result in a decrease in the precision of the calculated relative atomic mass.

2.122 To calculate the number of carbon-12 atoms, we can convert amu to atoms using the atomic mass of carbon-12 (defined as exactly 12):

$$\text{Number of carbon-12 atoms} = \text{mass in amu} \times \frac{1 \text{ carbon-12 atom}}{12 \text{ amu}} = \text{carbon-12 atoms}$$

(a) $120\ \cancel{\text{amu}} \times \dfrac{1 \text{ carbon-12 atom}}{12\ \cancel{\text{amu}}} = 10 \text{ carbon-12 atoms}$

(b) $12{,}000\ \cancel{\text{amu}} \times \dfrac{1 \text{ carbon-12 atom}}{12\ \cancel{\text{amu}}} = 1.0 \times 10^{3} \text{ carbon-12 atoms}$

(c) $7.22 \times 10^{24}\ \cancel{\text{amu}} \times \dfrac{1 \text{ carbon-12 atom}}{12\ \cancel{\text{amu}}} = 6.02 \times 10^{23} \text{ carbon-12 atoms}$

2.123 The mass number is 127. Since there is only one isotope, the mass number should be very close to the relative atomic mass.

2.124 To calculate the number of iodine atoms, we can convert amu to atoms using the relative mass of iodine (126.9 amu).

$$\text{Number of iodine atoms} = \text{mass in amu} \times \frac{1 \text{ iodine atom}}{126.9 \text{ amu}} = \text{iodine atoms}$$

(a) $127\ \cancel{\text{amu}} \times \dfrac{1 \text{ iodine atom}}{126.9\ \cancel{\text{amu}}} = 1 \text{ iodine atom}$

(b) $12{,}700\ \cancel{\text{amu}} \times \dfrac{1 \text{ iodine atom}}{126.9\ \cancel{\text{amu}}} = 100 \text{ iodine atoms}$

(c) $7.22 \times 10^{24}\ \cancel{\text{amu}} \times \dfrac{1 \text{ iodine atom}}{126.9\ \cancel{\text{amu}}} = 5.69 \times 10^{23} \text{ iodine atoms}$

2.125 Since each carbon atom has less relative atomic mass than an iodine atom, there will be more carbon atoms.

2.126 Since the balloons contain equal numbers of atoms, the balloon with more massive atoms will be more dense. Since argon has a much higher relative atomic mass than helium, the argon-filled balloon will be more dense.

2.127 The relative atomic mass of boron is 10.81 amu. Since boron exists as either boron-10 or boron-11, the relative abundance of boron-11 will be much higher. Of the choices given only two make sense:

20% boron-10 and 80% boron-11 (correct answer)
5.0% boron-10 and 95.0% boron-11

The best answer can be found by calculating the relative atomic mass based on these percentages:

	Isotope mass	×	abundance	=	mass contribution from isotope
^{11}B	11.009 amu	×	0.800	=	8.81 amu
^{10}B	10.013 amu	×	0.200	=	2.00 amu
					10.81 amu (relative atomic mass of B)

2.128 Hydrogen is one of the diatomic molecules, so the formula for hydrogen gas is $H_2(g)$. You would expect a similar reaction with potassium because the potassium is in the same group as sodium. Copper and silver are in different groups so they would not be expected to react in the same way as sodium and potassium.

2.129 $Br_2(l)$

2.130 Nitrogen and oxygen are two of the diatomic elements. The formulas are written, $N_2(g)$ and $O_2(g)$, respectively.

2.131 Hydrogen is a nonmetal.

2.132 Carbon is the only nonmetal in group IVA (14). Silicon and germanium are metalloids and the rest are metals.

2.133 Adding a proton changes the element. So by adding a proton you would form a cation of a different element.

2.134 Mendeleev's periodic table was not arranged in order of increasing atomic number. Rather, he arranged the 63 known elements in order of increasing relative atomic mass and grouped elements with similar properties into columns and rows so that the properties varied in a regular pattern.

2.135 The subatomic particle described in (a) corresponds to a neutron because it has a zero charge. The subatomic particle in (b) has a mass that is four orders of magnitude less than a neutron. This subatomic particle is an electron and carries a charge of 1–. The subatomic particle described in (c) has a mass very close to the mass of neutron. This corresponds to a proton which carries a charge of 1+.

	Particle	**Mass (g)**	**Relative Charge**
(a)	neutron	1.6749×10^{-24}	0
(b)	electron	9.1094×10^{-28}	1–
(c)	proton	1.6726×10^{-24}	1+

2.136 The number of protons corresponds to the atomic number on the periodic table. The element with 28 protons is nickel.

2.137 Fluorine-18 is the designation for an isotope of this element with a mass number of 18. The mass number (A) is the sum of the neutrons (N) and protons (Z): $A = N + Z$. Since the atomic number of fluorine is 9, the neutron number is $N = A - Z = 18 - 9 = 9$. A neutral atom of fluorine would have 9 electrons to balance the 9 protons in the nucleus.

2.138 The atomic number of silver is 47. The number corresponds to the protons and electrons of a neutral atom of silver. When silver tarnishes and the Ag^+ ion is formed, the positive charge indicates that an electron has been lost. The Ag^+ ion contains 47 protons and 46 electrons.

2.139 Neutral Fe atoms differ from Fe^{2+} and Fe^{3+} ions in number of electrons. Each of these contains 26 protons. A neutral Fe atom has 26 electrons, an Fe^{2+} ion has 24 electrons, and an Fe^{3+} ion has 23 electrons.

2.140 The ice cube containing hydrogen-2 would have the greater density. The volume of the ice cube would not vary with the identity of the hydrogen isotope. However, the mass of hydrogen-2 is twice that of hydrogen-1. This would make the ratio of mass to volume (the density) greater for the ice cube containing hydrogen-2.

2.141 Iodine is in group VIIA (17) and period 5 of the periodic table. An element in group VIIA (17) is also known as a halogen.

2.142 In Thomson’s plum pudding model, the atom was thought to consist electrons (“plums”) scattered throughout a sphere of positive charge (“pudding”). If this model was correct, the alpha particles bombarding the gold foil (Rutherford’s experiment) would have all passed through the foil undeflected. Instead, some of the alpha particles were deflected indicating that the model of the atom needed to include a dense core that corresponds to the nucleus.

2.143 A relative atomic mass of 12.01 amu for carbon indicates that the predominant naturally occurring isotope must be carbon-12. That the relative atomic mass is just slightly greater than 12 indicates there is very little carbon-14 in natural abundance. The most likely abundance of carbon-14 from the possibilities listed is less than 0.1%.

2.144 Potassium-40 and argon-40 have the same mass number, indicating that the sum of protons and neutrons is the same. However, these isotopes differ in the number of protons and neutrons. Each potassium-40 atom contains 19 protons and 21 neutrons. An atom of argon-40 contains 18 protons and 20 neutrons.