Chapter 1

Elements and the Periodic Table

Section 1.1 The Nature of Atoms

Solutions for Practice Problems
Student Edition page 19

1. Practice Problem (page 19)
Chlorine exists naturally as 75.78% chlorine-35 (mass = 34.97 u) and 24.22% chlorine-37 (mass = 36.97 u). What is the average atomic mass of chlorine?

What Is Required?
You need to find the average atomic mass of chlorine, Cl.

What Is Given?
You know the mass and isotopic abundance of Cl-35: mass = 34.97 u and isotopic abundance = 75.78%
You know the mass and isotopic abundance of Cl-37: mass = 36.97 u and isotopic abundance = 24.22%

Plan Your Strategy
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

Act on Your Strategy
average atomic mass of Cl = contribution of \(_{17}^{35}\text{Cl}\) + contribution of \(_{17}^{37}\text{Cl}\)
= 34.97 u \times 0.7578 + 36.97 u \times 0.2422
= 26.5003 u + 8.9541 u
= 35.4544 u
= 35.45 u

The average atomic mass of chlorine is 35.45 u.
Check Your Solution
The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the larger isotopic abundance. Four significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

2. Practice Problem (page 19)
Boron exists naturally as boron-10 (mass = 10.01 u; isotopic abundance = 19.8%) and boron-11 (mass = 11.01 u; isotopic abundance = 80.2%). What is the average atomic mass of boron?

What Is Required
You need to find the average atomic mass of boron, B.

What Is Given?
You know the mass and isotopic abundance of B-10:
- mass = 10.01 u and isotopic abundance = 19.8%
You know the mass and isotopic abundance of B-11:
- mass = 11.01 u and isotopic abundance = 80.2%

Plan Your Strategy
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

Act on Your Strategy
\[
\text{average atomic mass of } B = \text{ contribution of } ^{10}B + \text{ contribution of } ^{11}B \\
= 10.01 \text{ u} \times 0.198 + 11.01 \text{ u} \times 0.802 \\
= 1.98198 \text{ u} + 8.83002 \text{ u} \\
= 10.81 \text{ u} \\
= 10.8 \text{ u}
\]

The average atomic mass of boron is 10.8 u.

Check Your Solution
The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the larger isotopic abundance. Three significant digits are appropriate, based on the number of significant digits in the isotopic abundances.
3. **Practice Problem (page 19)**
Lithium is composed of 7.59% lithium-6 (mass = 6.02 u) and 92.41% lithium-7 (mass = 7.02 u). Calculate the average atomic mass of lithium.

**What Is Required**
You need to find the average atomic mass of lithium, Li.

**What Is Given?**
You know the mass and isotopic abundance of Li-6:
- mass = 6.02 u and isotopic abundance = 7.59%
You know the mass and isotopic abundance of Li-7:
- mass = 7.02 u and isotopic abundance = 92.41%

**Plan Your Strategy**
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

**Act on Your Strategy**
\[
\text{average atomic mass of Li} = \text{contribution of } ^{6}\text{Li} + \text{contribution of } ^{7}\text{Li} \\
= 6.02 \text{ u} \times 0.0759 + 7.02 \text{ u} \times 0.9241 \\
= 6.4569 \text{ u} + 6.487 \text{ u} \\
= 6.9439 \text{ u} \\
= 6.94 \text{ u}
\]

The average atomic mass of lithium is 6.94 u.

**Check Your Solution**
The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the larger isotopic abundance. Three significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

4. **Practice Problem (page 19)**
Magnesium exists naturally as 78.99% magnesium-24 (mass = 23.99 u), 10.00% magnesium-25 (mass = 24.99 u), and 11.01% magnesium-26 (mass = 25.98 u). What is the average atomic mass of magnesium?

**What Is Required?**
You need to find the average atomic mass of magnesium, Mg.
What Is Given?
You know the mass and isotopic abundance of Mg-24: 
mass = 23.99 u and isotopic abundance = 78.99%
You know the mass and isotopic abundance of Mg-25: 
mass = 24.99 u and isotopic abundance = 10.00%
You know the mass and isotopic abundance of Mg-26: 
mass = 25.98 u and isotopic abundance = 11.01%

Plan Your Strategy
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

Act on Your Strategy
average atomic mass of Mg = contribution of $^{24}\text{Mg} + \text{ contribution of }^{25}\text{Mg} + \text{ contribution of }^{26}\text{Mg}$
= $23.99 \text{ u} \times 0.7899 + 24.99 \text{ u} \times 0.1000 + 25.98 \text{ u} \times 0.1101$
= $18.9497 \text{ u} + 2.4990 \text{ u} + 2.8604 \text{ u}$
= $24.3091 \text{ u}$
= $24.31 \text{ u}$

The average atomic mass of magnesium is 24.31 u.

Check Your Solution
The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the largest isotopic abundance. Four significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

5. Practice Problem (page 19)
Gallium exists naturally as gallium-69 (mass = 68.93 u; isotopic abundance = 60.1%) and gallium-71 (mass = 70.92 u). What is the isotopic abundance of gallium-71? What is the average atomic mass of gallium?

What Is Required?
You need to find the isotopic abundance of Ga-71 and the average atomic mass of gallium, Ga.
What Is Given?
You know the mass and isotopic abundance of Ga-69:
mass = 68.93 u and isotopic abundance = 60.1%
You know the mass of Ga-71:
mass = 70.92 u

Plan Your Strategy
Find the isotopic abundance of Ga-71 by subtracting the isotopic abundance of Ga-69 from 100%.
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

Act on Your Strategy
isotopic abundance of gallium-71 = 100% − isotopic abundance of gallium-69
= 100% − 60.1%
= 39.9%

average atomic mass of Ga = contribution of $^{69}$Ga + contribution of $^{71}$Ga
= 68.93 u × 0.601 + 70.92 u × 0.399
= 41.4269 u + 28.2971 u
= 69.72 u
= 69.7 u

The average atomic mass of gallium is 69.7 u.

Check Your Solution
The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the larger isotopic abundance. Three significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

6. Practice Problem (page 19)
Bromine exists naturally as bromine-79 (mass = 78.92 u; isotopic abundance 50.69%) and bromine-81 (mass = 80.92 u). What is the isotopic abundance of bromine-81? What is the average atomic mass of bromine?

What Is Required?
You need to find the isotopic abundance of Br-81 and the average atomic mass of bromine, Br.
What Is Given?
You know the mass and isotopic abundance of Br-79:
mass = 78.92 u and isotopic abundance = 50.69%
You know the mass of Br-81:
mass = 80.92 u

Plan Your Strategy
Find the isotopic abundance of Br-81 by subtracting the isotopic abundance of Br-79 from 100%.
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

Act on Your Strategy
isotopic abundance of Br-81 = 100% − isotopic abundance of Br-79
= 100% − 50.69%
= 49.31%

average atomic mass of Br = contribution of Br + contribution of Br
= 78.92 u × 0.5069 + 80.92 u × 0.4931
= 40.0045 u + 39.9016 u
= 79.9061 u
= 79.91 u

The average atomic mass of bromine is 79.91 u.

Check Your Solution
The calculated average atomic mass is about halfway between the atomic masses of the isotopes. This is reasonable since the abundances of the two isotopes are about equal. Four significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

7. Practice Problem (page 19)
A sample of rubidium is 72.17% rubidium-85 (mass = 84.91 u) and 27.83% rubidium-87 (mass = 86.91 u). Calculate the average atomic mass of rubidium.

What Is Required?
You need to find the average atomic mass of rubidium, Rb.
What Is Given?
You know the mass and isotopic abundance of Rb-85:
mass = 84.91 u and isotopic abundance = 72.17%
You know the mass and isotopic abundance of Rb-87:
mass = 86.91 u and isotopic abundance = 27.83%

Plan Your Strategy
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.

Act on Your Strategy
\[
\text{average atomic mass of Rb} = \text{contribution of } ^{85}\text{Rb} + \text{contribution of } ^{87}\text{Rb}
\]
\[
= 84.91 \text{ u} \times 0.7217 + 86.91 \text{ u} \times 0.2783
\]
\[
= 61.27955 \text{ u} + 24.18705 \text{ u}
\]
\[
= 85.46660 \text{ u}
\]
\[
= 85.47 \text{ u}
\]

The average atomic mass of rubidium is 85.47 u.

Check Your Solution
The calculated average atomic mass is closer to the atomic mass of Rb-85. This is expected since the abundance of Rb-85 is about three times that of Rb-87. Four significant digits are appropriate, based on the number of significant digits in the isotopic abundance.

8. Practice Problem (page 19)
The average atomic mass of nitrogen is 14.01 u. Nitrogen exists naturally as nitrogen-14 (mass = 14.00 u) and nitrogen-15 (mass = 15.00 u). What can you infer about the isotopic abundances for nitrogen?

What Is Required?
You must infer the isotopic abundances from the average atomic mass and the mass of each isotope.

What Is Given?
You know the mass of N-14:
mass = 14.00 u
You know the mass of N-15:
mass = 15.00 u
Plan Your Strategy
Examine the mass of each isotope and compare these values with the average mass of nitrogen.

Act on Your Strategy
Since the average atomic mass of nitrogen (14.01 u) is very close to the mass of the isotope nitrogen-14 (14.00 u), N-14 must make up almost all of the isotopic abundance of this element. In other words, the isotopic abundance of nitrogen-14 in nitrogen is very high, with just a trace amount of nitrogen-15.

Check Your Solution
The inference seems reasonable since the average atomic mass of nitrogen is almost the same as the mass of the most abundant isotope, N-14.

9. Practice Problem (page 19)
The following isotopes of rhenium are found in nature.
rhenium-185: \( m = 184.953 \text{ u} \); abundance = 37.4%
rhenium-187: \( m = 186.956 \text{ u} \); abundance = 62.6%
Analyze the data. Predict the average atomic mass and write down your prediction. Then calculate the average atomic mass and compare it with your prediction.

What Is Required?
You need to predict the average atomic mass of rhenium, Re, from the masses and abundances of the two isotopic forms of this element.

What Is Given?
You know the mass and isotopic abundance of Re-185:
mass = 184.953 u and isotopic abundance = 37.4%
You know the mass and isotopic abundance of Re-187:
mass = 186.956 u and isotopic abundance = 62.6%

Plan Your Strategy
Examine the mass and abundance of each isotope and predict the average atomic mass of rhenium.
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Add the contributions of the isotopes to determine the average atomic mass of the element.
**Act on Your Strategy**
The average atomic mass is predicted to be closer to 187 u (roughly 60% of the mass) than 185 u (roughly 40% of the mass). The average atomic mass should be around 186.2 u.

\[
\text{average atomic mass of Re} = \text{contribution of } {}^{185}\text{Re} + \text{contribution of } {}^{187}\text{Re}
\]
\[
= 184.953 \text{ u} \times 0.374 + 186.956 \text{ u} \times 0.626
\]
\[
= 69.17242 \text{ u} + 117.03446 \text{ u}
\]
\[
= 186.20688 \text{ u}
\]
\[
= 186.21 \text{ u}
\]
\[
= 186 \text{ u}
\]

The average atomic mass of rhenium is 186 u.

**Check Your Solution**
The prediction agrees with the calculated average atomic mass. The calculated average atomic mass is between the atomic masses of the isotopes, but closer to the atomic mass of the isotope that has the larger isotopic abundance. Three significant digits are appropriate, based on the number of significant digits in the isotopic abundances.

**10. Practice Problem (page 19)**
The average atomic mass of iridium is 192.22 u. If iridium-191 has an atomic mass of 190.961 u and an isotopic abundance of 37.3%, and iridium-193 is the only other naturally occurring isotope, what is the atomic mass of iridium-193?

**What Is Required?**
You need to find the isotopic abundance and atomic mass of Ir-193.

**What Is Given?**
You know the mass and isotopic abundance of Ir-191:
mass = 190.961 u and isotopic abundance = 37.3%
You know the average atomic mass of iridium:
mass = 192.22 u

**Plan Your Strategy**
Find the isotopic abundance of Ir-193 by subtracting the isotopic abundance of Ir-191 from 100%.
Let the atomic mass of Ir-193 = \( y \).
Multiply the mass of each isotope by its isotopic abundance, expressed as a decimal, to determine the contribution of each isotope to the average atomic mass.
Adding the contributions of the isotopes gives the average atomic mass of iridium. Solve the equation for the mass of Ir-193.
**Act on Your Strategy**

isotopic abundance of Ir-193 = 100% − isotopic abundance of Ir-191

= 100% − 37.3%

= 62.7%

Let \( y \) be the atomic mass of Ir-193.

average atomic mass of Ir = contribution of \( ^{191}\text{Ir} \) + contribution of \( ^{193}\text{Ir} \)

\[
192.22 \text{ u} = 190.961 \text{ u} \times 0.373 + (y) \text{ u} \times 0.627
\]

\[
= 71.2284 \text{ u} + 0.627y
\]

\[
0.627y = 192.22 \text{ u} - 71.2284 \text{ u}
\]

\[
0.627y = 120.9916 \text{ u}
\]

\[
y = \frac{120.9916 \text{ u}}{0.627}
\]

\[
y = 192.97 \text{ u}
\]

\[
y = 193 \text{ u}
\]

The atomic mass of iridium-193 is 193 u.

**Check Your Solution**

The given average atomic mass of iridium is closer to the calculated mass of Ir-193. Since Ir-193 has the greater isotopic abundance, this is a reasonable answer. Three significant digits are appropriate, based on the number of significant digits given for the isotopic abundance.

**Section 1.1 The Nature of Atoms**

**Solutions for Selected Review Questions**

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10. **Review Question (page 21)**

The atomic masses and isotopic abundances of the naturally occurring isotopes of silicon are given below. Calculate the average atomic mass of silicon.

\[ ^{28}\text{Si} : \text{mass} = 27.977 \text{ u}; \text{isotopic abundance} = 92.23\% \]

\[ ^{29}\text{Si} : \text{mass} = 28.976 \text{ u}; \text{isotopic abundance} = 4.67\% \]

\[ ^{30}\text{Si} : \text{mass} = 29.97 \text{ u}; \text{isotopic abundance} = 3.10\% \]

**What Is Required**

You need to find the average atomic mass of silicon, Si.
What Is Given?
You know the mass and isotopic abundance of Si-28:
mass = 27.977 u and isotopic abundance = 92.23%
You know the mass and isotopic abundance of Si-29:
mass = 28.976 u and isotopic abundance = 4.67%
You know the mass and isotopic abundance of Si-30:
mass = 29.97 u and isotopic abundance = 3.10%

Plan Your Strategy
Multiply the mass of each isotope by its isotopic abundance, expressed as a
decimal, to determine the contribution of each isotope to the average atomic
mass.
Add the contributions of the isotopes to determine the average atomic mass of
the element.

Act on Your Strategy

\[
\text{average atomic mass of Si} = \text{contribution of } ^{28}\text{Si} + \text{contribution of } ^{29}\text{Si} + \text{contribution of } ^{30}\text{Si}
\]
\[
= 27.977 \text{ u} \times 0.9223 + 28.976 \text{ u} \times 0.0467 + 29.97 \text{ u} \times 0.0310
\]
\[
= 25.803187 \text{ u} + 1.353179 \text{ u} + 0.929070 \text{ u}
\]
\[
= 28.085436 \text{ u}
\]
\[
= 28.08 \text{ u}
\]
\[
= 28.1 \text{ u}
\]

The average atomic mass of silicon is 28.1 u.

Check Your Solution
The calculated average atomic mass is between the atomic masses of the
isotopes, but closer to the atomic mass of the isotope that has the largest
isotopic abundance. Three significant digits are appropriate, based on the
number of significant digits in the isotopic abundances.

11. Review Question (page 21)
The atomic mass of yttrium-89 is 88.91 u. The average atomic mass of yttrium
that is reported in periodic tables is 88.91 u. Infer why these values are the
same.

Since the average atomic mass of yttrium and the mass of the isotope Y-89 are
the same, it can be inferred that yttrium exists in only one isotopic form, Y-89.
14. Review Question (page 21)
Animals consume carbon-containing compounds and incorporate them into their tissues and breathe out carbon dioxide. A very small amount of that carbon is radioactive carbon-14. When the animals die, they no longer exchange carbon with the environment. Based on this information, how do you think that scientists use the percentage of carbon-14 in a fossil to determine its age? (This technique is called carbon dating.)

In the environment, there is a constant decay/replenishment process between carbon-12 and radioactive carbon-14, and all living plants and animals have a constant ratio of C-14 to C-12. When an organism dies, its C-14 is no longer replenished. As the C-14 decays, the amount of this isotope decreases and the ratio of C-14 to C-12 decreases. The half-life of a radioactive element is the time it takes for half of the atoms to change or decay into another isotopic form or sometimes into a different element. Since the half-life of C-14 is known, the ratio between these two isotopes in a fossil’s remains can be related to the age of the fossil.

Section 1.2 The Periodic Table
Solutions for Selected Review Questions
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4. Review Question (page 30)
Sodium metal is usually stored in oil. Find sodium in the periodic table. Based on its position in the periodic table, why do you think that sodium must be stored in oil?

Sodium is a reactive alkali metal found in Group 1 of the periodic table. It is stored in oil to keep oxygen and water vapour from coming in contact with this metal.

5. Review Question (page 30)
The platter shown here is made from copper. Why could silicon not be used to make objects such as this?

Silicon is a brittle metalloid that would break apart if pressure were used in an attempt to change its shape. It has a very high melting point (1687 K) and could not be easily moulded into another shape.
8. **Review Question (page 30)**

List the following elements in order, from best to worst conductor of electric current: silicon, copper, iodine.

Listed from best conductor to worst: copper, silicon, iodine.

9. **Review Question (page 30)**

If the inner transition elements were inserted into the periodic table between Groups 3 and 4 in Periods 6 and 7, how many columns wide would the table be? Explain your answer.

Each of the inner transition metal groups has 14 elements. The number of columns would increase from 18 to 32.

12. **Review Question (page 30)**

Strontium is sometimes called a “bone seeker” because, when ingested, it is deposited in bones and remains there for long periods of time. Find strontium in the periodic table. Offer a possible explanation for why strontium is a “bone seeker.”

Strontium is in Group 2 of the periodic table. Like calcium, it has two valence electrons and the two elements are chemically similar, with strontium being slightly more reactive. Calcium accumulates in and is incorporated into the structure of bone. Strontium has this same property. If strontium compounds are taken into the body, the strontium will accumulate in the bone.

13. **Review Question (page 30)**

Examine each of the following pairs of elements. State which element is the most reactive, and explain how you made your choice.

a. Na or Mg
b. Br or Kr
c. H or He

a. Sodium (Na) is more reactive. Sodium is a Group 1 alkali metal, whereas magnesium (Mg) is a Group 2 alkaline earth metal. The alkali metals are the most reactive metals.

b. Bromine (Br) is more reactive. Bromine is a Group 17 halogen, the most reactive group of the non-metals. Krypton (Kr) is a noble gas in Group 18. These non-metals are unreactive.

c. Hydrogen (H) is more reactive. Although it is in Group 1, it is not an alkali metal; it is a reactive non-metal. Helium (He) is a Group 18 noble gas, the most unreactive of the non-metal elements.
14. **Review Question (page 30)**
The photograph below shows what happened when a sample of a pure element was dropped in water. Name two elements that could have caused this reaction.

![Image of reaction]

Rubidium and cesium will react violently with water.

**Section 1.3 Explaining Periodic Trends**

**Solutions for Selected Review Questions**

**Student Edition page 40**

2. **Review Question (page 40)**
List the following elements in order of increasing radius: Ba, Cs, O, Sb, Sn.

In order of increasing size: O, Sb, Sn, Ba, Cs

4. **Review Question (page 40)**
The following figures represent neutral atoms gaining or losing an electron to become ions. One of these figures represents a chlorine atom, Cl, gaining an electron to become a negatively charged ion, Cl⁻. The other figure represents a sodium atom, Na, losing an electron to become a positively charged ion, Na⁺. Identify which figure represents each element. Provide possible explanations for the change in size upon ionization. **Note:** The unit pm (picometre) is $10^{-12}$ m.
Figure A represents the sodium atom (Na, $r = 186$ pm), which loses an electron when it forms the Na$^+$ ion ($r = 102$ pm). The Na atom has 11 protons in the nucleus and 11 electrons in the outer electron cloud, whereas the Na$^+$ ion has 11 protons and only 10 electrons. The net attractive force of 11 protons on 10 electrons is greater than that of 11 protons on 11 electrons. Thus, the electron cloud is pulled closer to the nucleus in the ion, rendering the net radius of Na$^+$ to be smaller than that of Na.

Likewise, Figure B presents the Cl atom (Cl, $r = 100$ pm), which gains an electron when it forms the Cl$^-$ ion ($r = 181$ pm). The Cl atom has 17 protons in the nucleus and 17 electrons in the outer electron cloud, whereas the Cl$^-$ ion has 17 proton and 18 electrons. Since both have the same number of protons in their nuclei, but there are more electrons in the Cl$^-$ ion, the net repelling forces of the electrons against one another will be greater in Cl$^-$. Thus, the electrons in Cl$^-$ can "escape" farther away from the nucleus than the electrons in Cl can, rendering the net radius of Cl$^-$ to be greater.

7. **Review Question (page 40)**
   Write a chemical equation for an atom losing a third electron. Would the ionization energy for this reaction be larger or smaller than the ionization energy for the loss of the first or second electron? Explain.

   \[ \text{A}^{2+}(g) + \text{energy} \rightarrow \text{A}^{3+}(g) + e^- \text{ (third ionization)} \]
   The third ionization energy is greater than the first or second ionization energies. When an electron is removed, there are fewer negative charges repelling one another but the same number of positive charges attracting the electrons. The force of attraction of the protons on each electron increases with fewer electrons to hold. Consequently, more energy is required to remove each successive electron.

9. **Review Question (page 40)**
   List the following elements in order of increasing ionization energy for the first ionization:
   arsenic, cesium, fluorine, helium, phosphorus, strontium

   In order of increasing first ionization energy: Cs, Sr, As, P, F, He
15. **Review Question (page 40)**

Use the data in Figure 1.19 and Figure 1.20 on page 35 to answer the following questions. First ionization data can also be found in the data table in Investigation 1-B on page 42.

![First Ionization Energies of Elements](image)

**Figure 1.19** The data points on this graph represent the amount of energy required to remove one electron from a neutral atom of the various elements. The numbers on the scale must be multiplied by $10^{-18}$ to give the correct value in Joules. The dashed lines represent sections in which data points are missing.

![Periodic Table](image)

**Figure 1.20** The values in this table represent electron affinities. The values must be multiplied by $10^{-19}$ to get the correct values in Joules.

- **a.** What would the total change in energy be to remove an electron from a sodium atom and add the electron to a chlorine atom?
- **b.** Which of the following combinations would be more stable: a sodium atom and a chlorine atom, or a positively charged sodium ion and a negatively charged chloride ion? Explain your reasoning.
a. Equation 1: \[ \text{N}(g) + 8.23 \times 10^{-19} \text{ J} \rightarrow \text{Na}^+(g) + 1\text{e}^- \]
Equation 2: \[ \text{Cl}(g) + 1\text{e}^- \rightarrow \text{Cl}^-(g) + 5.78 \times 10^{-19} \]
Equations 1 + 2: \[ \text{Na}(g) + \text{Cl}(g) + 2.45 \times 10^{-19} \text{ J} \rightarrow \text{Na}^+(g) + \text{Cl}^-(g) \]

Net energy added = \(2.45 \times 10^{-19} \text{ J}\)

b. The combination of Na\(^+\) and Cl\(^-\) would be more stable because of the electrical attraction between opposite charges.