

CHAPTER 1. Chemistry and the Scientific Method

1-1. Write 1.7201×10^{-3} as 172.01×10^{-5} and add it to 2.26×10^{-5} to obtain 174.27×10^{-5} or 1.7427×10^{-3} .

1-2. Consulting Table 1.2 we have that

(a) $400 \text{ nm} = 400 \times 10^{-9} \text{ m} = 4.00 \times 10^{-7} \text{ m}$

(b) $20 \text{ ps} = 20 \times 10^{-12} \text{ s} = 2.0 \times 10^{-11} \text{ s}$

1-3. (a) Using Equation 1.2,

$$t(\text{in } ^\circ\text{C}) = \left(\frac{5}{9}\right) (-90 - 32.0) = -68^\circ\text{C}$$

Using Equation 1.1,

$$T(\text{in K}) = -68 + 273.15 = 205 \text{ K}$$

(b) Let $x = t(\text{in } ^\circ\text{C}) = t(\text{in } ^\circ\text{F})$, then

$$x = \frac{5}{9}(x - 32.0) = \frac{5x}{9} - 17.8$$

$$9x = 5x - 160$$

$$4x = -160$$

$$x = -40$$

1-4. Using Equation 1.3,

$$d = \frac{m}{V} = \frac{22.55 \text{ g}}{1.667 \text{ mL}} = 13.53 \text{ g}\cdot\text{mL}^{-1} = 13.53 \text{ g}\cdot\text{cm}^{-3}$$

1-5. We must convert kilometers to meters

$$950 \text{ km} = 950 \times 10^3 \text{ m} = 9.50 \times 10^5 \text{ m}$$

and hours to seconds

$$1 \text{ hr} = 60 \text{ min} = 3600 \text{ s}$$

Thus, the velocity is

$$v = \frac{9.50 \times 10^5 \text{ m}}{3600 \text{ s}} = 263.9 \text{ m}\cdot\text{s}^{-1}$$

The kinetic energy is

$$\begin{aligned} E_{\text{K}} &= \frac{1}{2}mv^2 = \left(\frac{1}{2}\right) (3.22 \times 10^5 \text{ kg}) (263.9 \text{ m}\cdot\text{s}^{-1})^2 \\ &= 1.12 \times 10^{10} \text{ kg}\cdot\text{m}^2\cdot\text{s}^{-2} = 1.12 \times 10^{10} \text{ J} = 1.12 \times 10^7 \text{ kJ} \end{aligned}$$

1-6. We use Equation 1.5 to calculate the potential energy

$$E_{\text{p}} = mgh = (1 \text{ kg}) (9.81 \text{ m}\cdot\text{s}^{-2}) (200 \text{ m}) = 1.96 \times 10^3 \text{ J} = 1.96 \text{ kJ}$$

This energy is converted into kinetic energy as it falls.

1-7. The energy expended is given by ($1 \text{ kW} = 1 \text{ kJ}\cdot\text{s}^{-1}$)

$$\begin{aligned} E &= (7.5 \text{ kJ}\cdot\text{s}^{-1}) (55 \text{ min}) \left(\frac{60 \text{ s}}{1 \text{ min}}\right) \\ &= 24\,750 \text{ kJ} = 24\,750 \text{ kW}\cdot\text{s} \left(\frac{1 \text{ hr}}{3600 \text{ s}}\right) = 6.875 \text{ kW}\cdot\text{h} \end{aligned}$$

At a cost of \$0.10 per kW·h, we have \$0.69 to two significant figures.

1-8. The average of the measurements is given by

$$\text{ave} = \frac{(0.203 + 0.209 + 0.199) \mu\text{m}}{3} = 0.204 \mu\text{m}$$

The percentage error is given by (Equation 1.6)

$$\begin{aligned} \text{percentage error} &= \frac{(\text{average value}) - (\text{true value})}{\text{true value}} \times 100 \\ &= \frac{0.204 \mu\text{m} - 0.200 \mu\text{m}}{0.200 \mu\text{m}} \times 100 = 2\% \end{aligned}$$

See the next Section for a discussion of significant figures.

- 1-9. (a) 301 000 000; the last six zeros are not significant figures
 (b) 30006; all five digits are significant
 (c) 0.002 9060; the last five digits are significant; the leading zeros are not significant figures
 (d) 12 is exact because it is a counting number

1-10. $y = \frac{8.314 \times 298.15}{96\,485.3} = 0.02569$

The answer should be given to only four significant figures.

1-11. $y = \frac{7.2960}{8.9000} - 132.0 = 0.81978 - 132.0 = -131.2$

1-12. The percentage error is given by (Equation 1.6)

$$\begin{aligned} \text{percentage error} &= \frac{(\text{average value}) - (\text{true value})}{\text{true value}} \times 100 \\ &= \frac{3.03 - 2.70}{2.70} \times 100 = \frac{0.33}{2.70} \times 100 = 12\% \end{aligned}$$

1-13. Using the conversions given in Appendix B, we have

$$\begin{aligned} (72 \text{ km}) \left(\frac{1 \text{ mile}}{1.6093 \text{ km}} \right) &= 45 \text{ mile} \\ (1260 \text{ m}) \left(\frac{1.0936 \text{ yd}}{1 \text{ m}} \right) \left(\frac{3 \text{ ft}}{1 \text{ yd}} \right) &= 4130 \text{ ft} \end{aligned}$$

1-14. Using Table 1.2, the time to transfer the data is

$$(25 \text{ photos}) \left(\frac{1.2 \times 10^9 \text{ bytes}}{1 \text{ photo}} \right) \left(\frac{1 \text{ s}}{45 \times 10^6 \text{ bytes}} \right) = 670 \text{ s} = 11 \text{ min}$$

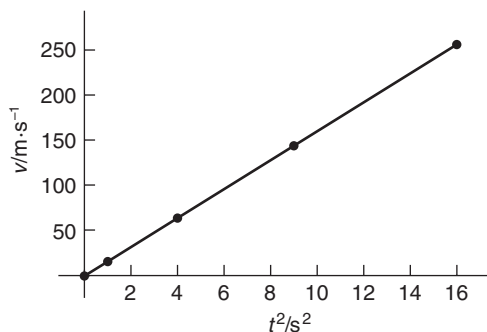
1-15. Express each piece of data as Charge/ 10^{-15} C as in the following table:

Charge/ 10^{-15} C
7.05
3.24
0.986

1-16. The appropriate table is

$v/\text{m}\cdot\text{s}^{-1}$	t^2/s^2
0	0
16	1.0
64	4.0
144	9.0
256	16.0

A plot of $v/\text{m}\cdot\text{s}^{-1}$ against t^2/s^2 is a straight line as shown below.



CHAPTER 2. Atoms and Molecules

2-1. We have that

$$\text{mass \% of an element} = \frac{\text{mass of the element}}{\text{mass of sample}} \times 100$$

Thus,

$$\text{mass of the element} = \frac{(\text{mass \% of an element})(\text{mass of sample})}{100}$$

The mass of each element is given by

$$\text{mass of K} = (5.650 \text{ g}) \left(\frac{38.67}{100} \right) = 2.185 \text{ g}$$

$$\text{mass of N} = (5.650 \text{ g}) \left(\frac{13.86}{100} \right) = 0.7831 \text{ g}$$

$$\text{mass of O} = 5.650 \text{ g} - 2.185 \text{ g} - 0.7831 \text{ g} = 2.682 \text{ g}$$

2-2. The atomic mass of O is 15.9994. The atomic mass of C based on O as exactly 16 would be

$$\text{atomic mass C} = (12.0107) \left(\frac{16}{15.9994} \right) = 12.0112$$

The difference is not significant to five significant figures.

2-3. See the text.

2-4. See the text.

2-5. We will use the atomic masses of bromine and fluorine to five significant figures. The molecular mass of BrF_5 is

$$\text{molecular mass} = (1)(79.904) + (5)(18.998) = 174.894$$

$$\text{mass \% Br} = \frac{79.904}{174.894} \times 100 = 45.687\%$$

$$\text{mass \% F} = \frac{5 \times 18.998}{174.894} \times 100 = 54.313\%$$

- 2-6. Assume that both the nucleus and the atom are spheres. The fraction of the volume of the atom that is occupied by the nucleus is given by

$$\text{fraction} = \frac{\text{volume of the nucleus}}{\text{volume of the atom}} = \frac{\frac{4}{3}\pi r_{\text{nuc}}^3}{\frac{4}{3}\pi r_{\text{atom}}^3} = \frac{r_{\text{nuc}}^3}{r_{\text{atom}}^3} = \left(\frac{r_{\text{nuc}}}{r_{\text{atom}}}\right)^3$$

We are given that

$$r_{\text{nuc}} = \frac{d_{\text{nuc}}}{2} = \frac{(1 \times 10^{-5})d_{\text{atom}}}{2} = 1 \times 10^{-5}r_{\text{atom}}$$

Thus, the fraction is given by

$$\text{fraction} = \left(\frac{1 \times 10^{-5}r_{\text{atom}}}{r_{\text{atom}}}\right)^3 = 1 \times 10^{-15}$$

The percentage of the atom occupied by the nucleus is $1 \times 10^{-13}\%$.

- 2-7. The atomic number of phosphorus is 15, so there are 15 protons in its nucleus and 15 electrons in the atom. The number of neutrons in the nucleus is $32 - 15 = 17$.
- 2-8. Let x be the percentage of lithium-7. Thus, the percentage of lithium-6 is $100 - x$. Then we have

$$6.941 = \left(\frac{100 - x}{100}\right)(6.015\ 1223) + \left(\frac{x}{100}\right)(7.016\ 0040)$$

Multiplying both sides of the equation by 100 gives

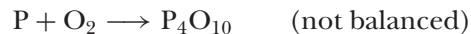
$$694.1 = (100 - x)(6.015\ 1223) + (x)(7.016\ 0040)$$

Solving for x gives $x = 92.5\%$. The natural abundance of lithium-7 is 92.5% and that of lithium-6 is $100\% - 92.5\% = 7.5\%$.

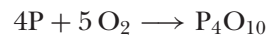
- 2-9. The atomic number of oxygen is 8. The number of electrons in the oxide ion is $8 + 2 = 10$ electrons. Examples of cations that are isoelectronic to O^{2-} are F^- , Na^+ , Mg^{2+} , and Al^{3+} .
- 2-10. (a) Silicon has an atomic number of 14. Therefore, there are 14 protons, $28 - 14 = 14$ neutrons, and $14 + 4 = 18$ electrons in ${}_{14}^{28}\text{Si}^{4-}$.
- (b) Tungsten has an atomic number of 74. Therefore, there are 74 protons, $186 - 74 = 112$ neutrons, and $74 - 5 = 69$ electrons in ${}_{74}^{186}\text{W}^{+5}$.
- (c) Uranium has an atomic number of 92. Therefore, there are 92 protons, $235 - 92 = 143$ neutrons, and 92 electrons in ${}_{92}^{235}\text{U}$.
- (d) Iron has an atomic number of 26. Therefore, there are 26 protons, $58 - 26 = 32$ neutrons, and $26 - 3 = 23$ electrons in ${}_{26}^{58}\text{Fe}^{+3}$.

CHAPTER 3. The Periodic Table and Chemical Periodicity

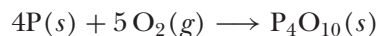
3-1. Step 1 is to write the following equation



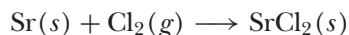
Using Step 2, there is one phosphorus atom on the left and four phosphorus atoms on the right, so we must add a 4 in front of P. Using Steps 3 and 4, there are 10 oxygen atoms on the right side, and so we add a 5 in front of O_2 to obtain



Finally adding the states of the reactants and product, we have

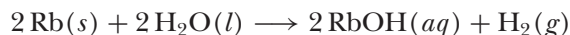


3-2. By analogy with the reaction of calcium with chlorine, so we predict (correctly) that



3-3. Iodine is in the same group as chlorine, so predict $\text{Cl}_4(s)$ and $\text{NI}_3(s)$.

3-4. Rubidium and sodium are in the same group. Using the reaction between sodium and water as an analogy, we write



3-5. See the text.

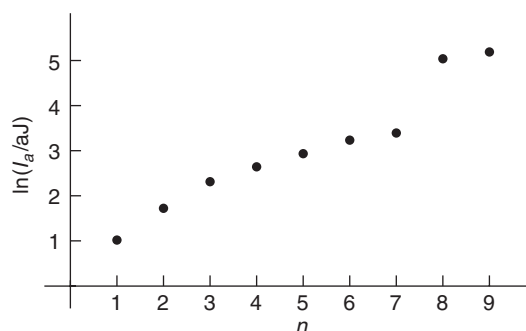
3-6. See the text.

CHAPTER 4. Early Quantum Theory

- 4-1. The energy required to produce B^{+3} from a boron atom is the sum of the first three ionization energies of a boron atom.

$$\text{energy} = (1.33 + 4.04 + 6.08) \text{ aJ} = 11.45 \text{ aJ}$$

- 4-2. We have the following figure:



- 4-3. We use Equation 4.1.

$$\begin{aligned}\lambda &= \frac{c}{\nu} = \frac{2.9979 \times 10^8 \text{ m}\cdot\text{s}^{-1}}{1.50 \times 10^{18} \text{ Hz}} = \frac{2.9979 \times 10^8 \text{ m}\cdot\text{s}^{-1}}{1.50 \times 10^{18} \text{ s}^{-1}} \\ &= 2.00 \times 10^{-10} \text{ m} = 0.200 \text{ nm}\end{aligned}$$

This wavelength corresponds to an X-ray.

- 4-4. The slope of the plot of E_k against ν gives the value of h . The frequencies corresponding to 0.6 aJ and 0.2 aJ are $14 \times 10^{14} \text{ s}^{-1}$ and $8 \times 10^{14} \text{ s}^{-1}$, respectively. Thus,

$$\begin{aligned}\text{slope} = h &= \frac{(E_k)_2 - (E_k)_1}{\nu_2 - \nu_1} = \frac{6 \times 10^{-19} \text{ J} - 2 \times 10^{-19} \text{ J}}{14 \times 10^{14} \text{ s}^{-1} - 8 \times 10^{14} \text{ s}^{-1}} \\ &= 6.7 \times 10^{-34} \text{ J}\cdot\text{s} = 7 \times 10^{-34} \text{ J}\cdot\text{s}\end{aligned}$$

to one significant figure.

4-5. We first use Equation 4.1

$$\nu = \frac{c}{\lambda} = \frac{2.9979 \times 10^8 \text{ m}\cdot\text{s}^{-1}}{656.3 \times 10^{-9} \text{ m}} = 4.568 \times 10^{14} \text{ s}^{-1}$$

The energy is given by Equation 4.3

$$\begin{aligned} E &= h\nu = (6.626 \times 10^{-34} \text{ J}\cdot\text{s})(4.568 \times 10^{14} \text{ s}^{-1}) \\ &= 3.027 \times 10^{-19} \text{ J} = 0.3027 \text{ aJ} \end{aligned}$$

4-6. We use Equation 4.7

$$\begin{aligned} \lambda &= \frac{h}{mv} = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s}}{(45.9 \text{ g}) \left(\frac{1 \text{ kg}}{1000 \text{ g}} \right) (120 \text{ mi}\cdot\text{h}^{-1}) \left(\frac{1.6093 \text{ km}}{1 \text{ mi}} \right) \left(\frac{1 \text{ hr}}{3600 \text{ s}} \right)} \\ &= 2.69 \times 10^{-31} \text{ m} \end{aligned}$$

A wavelength of $2.69 \times 10^{-31} \text{ m}$ is totally negligible relative to the diameter of a golf ball.

4-7. Using Equation 4.11, we have

$$\begin{aligned} \frac{1}{\lambda} &= (1.097 \times 10^7 \text{ m}^{-1}) \left(\frac{1}{3^2} - \frac{1}{5^2} \right) = 7.8009 \times 10^5 \text{ m}^{-1} \\ \lambda &= \frac{1}{7.8009 \times 10^5 \text{ m}^{-1}} = 1.282 \times 10^{-6} \text{ m} = 1282 \text{ nm} \end{aligned}$$

4-8. We use Equation 4.11 and the fact that $n_f = 2$ for the Balmer series

$$\begin{aligned} \frac{1}{656.3 \times 10^{-9} \text{ m}} &= (1.097 \times 10^7 \text{ m}^{-1}) \left(\frac{1}{4} - \frac{1}{n^2} \right) \\ 0.139 &= \frac{1}{4} - \frac{1}{n^2} \end{aligned}$$

or $n^2 = 9$ or $n = 3$. The state $n = 3$ is called the second excited state because it is two levels above the ground state of hydrogen ($n = 1$).

4-9. We use Equation 4.5 with $E = 2.18 \text{ aJ}$ (from Example 4-9)

$$\begin{aligned} \lambda &= \frac{hc}{E} = \frac{(6.626 \times 10^{-34} \text{ J}\cdot\text{s})(2.9979 \times 10^8 \text{ m}\cdot\text{s}^{-1})}{2.18 \times 10^{-18} \text{ J}} \\ &= 9.11 \times 10^{-8} \text{ m} = 91.1 \text{ nm} \end{aligned}$$

CHAPTER 5. Quantum Theory and Atomic Structure

5-1. We use Equation 5.1

$$\Delta p \geq \frac{h}{4\pi \Delta x} = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s}}{4\pi (1 \times 10^{-3} \text{ m})} = 5.3 \times 10^{-32} \text{ kg}\cdot\text{m}\cdot\text{s}^{-1}$$

The uncertainty in the speed is

$$\Delta v = \frac{\Delta p}{m} = \frac{5.3 \times 10^{-32} \text{ kg}\cdot\text{m}\cdot\text{s}^{-1}}{45.9 \times 10^{-3} \text{ kg}} \approx 1 \times 10^{-30} \text{ m}\cdot\text{s}^{-1}$$

The speed is

$$v = \left(\frac{200 \text{ km}}{1 \text{ hr}} \right) \left(\frac{1000 \text{ m}}{1 \text{ km}} \right) \left(\frac{1 \text{ hr}}{3600 \text{ s}} \right) = 60 \text{ m}\cdot\text{s}^{-1}$$

Thus, the percentage uncertainty in the speed is

$$\Delta v = \frac{1 \times 10^{-30} \text{ m}\cdot\text{s}^{-1}}{60 \text{ m}\cdot\text{s}^{-1}} \times 100 = 2 \times 10^{-30} \%$$

5-2. See the text.

5-3. See the text.

5-4. See the text.

5-5. See the text.

5-6. See the text.

5-7. The ground-state electron configuration of an O^{2-} ion (10 electrons) is $1s^2 2s^2 2p^6$. The first excited state occurs when one of the $2p$ electrons is promoted to the $3s$ state, giving $1s^2 2s^2 2p^5 3s^1$.

5-8. (a) A neon atom, with its ground-state electron configuration $[\text{He}]2s^2 2p^6$, has eight valence electrons.

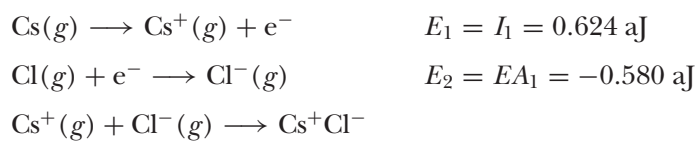
(b) The ground-state configuration of Al^{3+} is simply $[\text{Ne}]$ in the notation of Table 5.6. Being a cation, we say that Al^{3+} has no valence electrons.

- (c) The ground-state configuration of Mg^{2+} is simply [Ne] in the notation of Table 5.6. Being a cation, we say that Mg^{2+} has no valence electrons.
- (d) A phosphorus atom, with its ground-state electron configuration $[\text{Ne}]3s^23p^3$, has five valence electrons.
- (e) The ground-state configuration of Cl^- is $[\text{Ne}]3s^23p^6$ or [Ar] in the notation of Table 5.6. Being an anion, we say that Cl^- has eight valence electrons.

CHAPTER 6. Ionic Bonds and Compounds

- 6-1. Aluminum belongs to Group 13. An aluminum atom has three more electrons than a neon atom, or $[\text{Ne}]3s^23p^1$, and so it achieves a ground-state electron configuration of $[\text{Ne}]$ by losing three electrons to give Al^{3+} .
- 6-2. The ground-state electron configuration of a selenium atom is $[\text{Ar}]3s^23p^4$. It can achieve a krypton-like electron configuration by adding two electrons, giving Se^{2-} .
The ground-state electron configuration of a nitrogen atom is $[\text{He}]2s^22p^3$. It can achieve a neon-like electron configuration by adding three electrons, giving N^{3-} .
The ground-state electron configuration of a phosphorus atom is $[\text{Ne}]3s^23p^3$. It can achieve an argon-like electron configuration by adding three electrons, giving P^{3-} .
- 6-3. See the text.
- 6-4. See the text. Both Mg^{2+} and N^{3-} have a $[\text{Ne}]$ electron configuration ($1s^22s^22p^6$). To balance the total ionic charges, the formula of magnesium nitride is Mg_3N_2 , which is a solid at room temperature, as are all ionic compounds.
- 6-5. An indium atom has a ground-state electron configuration of $[\text{Kr}]5s^24d^{10}5p^1$. It can achieve an 18-electron configuration by losing its two $5s$ electrons and its one $5p$ electron to become $[\text{Ar}]4s^24p^64d^{10}$ or In^{3+} .
- 6-6. See the text.
- 6-7. The ground-state electron configuration of a Pd^{2+} ion is regular: $[\text{Kr}]4d^8$.
- 6-8. The ground-state electron configuration of these ions are as follows:
- $$\text{Cr(II)} : [\text{Ar}]3d^4 \quad \text{Cr(III)} : [\text{Ar}]3d^3 \quad \text{Cr(VI)} : [\text{Ar}]$$
- 6-9. Aluminum, magnesium, and sodium are in the same row of the periodic table (3rd row) and the ions are isoelectronic, and so the order of their sizes goes as $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+$. Bromine and iodine are in higher rows, and so we have $\text{Al}^{3+} < \text{Mg}^{2+} < \text{Na}^+ < \text{Br}^- < \text{I}^-$.

6-10. As in Example 6-10, we describe the reaction in three steps



$$E_3 = (231 \text{ aJ}\cdot\text{pm}) \left(\frac{(1)(-1)}{167 + 181} \right) = -0.664 \text{ aJ}$$

where we have used values from Tables 6.4 and 6.5. The total is

$$E_{\text{rxn}} = E_1 + E_2 + E_3 = -0.620 \text{ aJ}$$