A.1 INTRODUCTION

In this appendix, we will look at some of the quantitative relationships associated with the mass of elements and compounds. This subset of chemistry is called stoichiometry, a word derived from the Greek word “stoikheion”, meaning element. Before we begin, we need to say a few words about the approach that we will take to the calculations in this appendix.

A.2 THE CONVERSION FACTOR APPROACH TO CALCULATIONS

If someone told you that she was “six”, you might have a little trouble deciding what was meant. That person could be six years old, but if she were a college student, that would probably not be correct. She could weigh six tons or be six inches tall, but probably not. She is more likely six feet tall. The point is that there are two parts to a measurement or a piece of quantitative information; the “number” and the “unit”. Six inches, six feet and six meters all have the same number, but are clearly different lengths. In scientific measurements or calculations, we must pay attention to both the number and the unit.

In order to convert the height of six feet to inches, most of us would say “multiply by 12” to give an answer of 72 inches. Although the result is correct, we did not multiply by 12; we actually multiplied by 1! Here’s how:

We know the following equality: 12 in = 1 ft

\[
\frac{12 \text{ in}}{1 \text{ ft}} = \frac{1 \text{ ft}}{1 \text{ ft}} = 1
\]

The fraction in the box is called a “conversion factor” and it is equal to 1. In converting six feet to inches, we actually performed the following operation:

\[
6 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} = 72 \text{ in}
\]

In the above, the distance in feet is multiplied by a conversion factor to produce a distance in inches.

Notice that we did the operation on the numbers (six times twelve divided by one equals seventy-two) and on the units as well (feet times inches divided by feet equals inches; feet “cancel out”). Both the number and the unit changed, but the height did not (going from six feet to 72 inches, the person did not grow or shrink). This is what we would expect upon multiplication by 1. Although this is a trivial example, we will use this same conversion factor or factor label approach for nearly all of the stoichiometric calculations in this book.

A.3 MOLAR MASSES AND ATOMIC WEIGHTS OF THE ELEMENTS

The number under the symbol of an element on the periodic table is the element’s atomic weight. It represents the “average atomic weight” or “average atomic mass” of the element because it is determined from the masses and abundance of the different isotopes of the element. Although there is a technical difference between weight and mass (weight depends on the gravitational force where you do the measurement), the two terms are often used interchangeably. We will use the term mass here, although your instructor may refer to “atomic weight”. The mass of a single atom is the element’s atomic weight expressed in units of amu, atomic mass units. The mass of a mole of atoms is the atomic weight expressed in units of grams/mol. The latter is referred to as the molar mass of the element. In this book, we use the symbol \( M_m \) to represent molar mass.

A.4 RELATING GRAMS, MOLES AND MOLAR MASS

Molar mass can also be used as a conversion factor. Using carbon as an example, the molar mass can be expressed as a fraction:

\[
M_m = \frac{12.01 \text{ g}}{1 \text{ mol}}
\]

The above is a conversion factor and is also equal to 1 because 12.01 grams of carbon and one mole of carbon are the same amount of carbon. The molar mass of an element can be used to convert between grams and moles.
Example 1

How many grams of sulfur are present in 0.250 moles of sulfur?

Solution:
The molar mass of sulfur from the periodic table is 32.07 g/mol. We start with the given information and apply the molar mass as a conversion factor.

\[
0.250 \text{ moles} \times \frac{32.07 \text{ g}}{1 \text{ mol}} = 8.02 \text{ g}
\]

Example 2

How many moles of copper are contained in 525 g of copper?

Solution:
The molar mass of copper from the periodic table is 63.55 g/mol. We again start with the given information and apply the molar mass as a conversion factor, but this time we use it in its reciprocal form (turn it “upside down”) with moles in the numerator and grams in the denominator such that grams cancel out.

\[
\frac{525 \text{ g}}{63.55 \text{ g}} = 8.26 \text{ mol}
\]

Comment:
In examples 1 and 2, a conversion factor was applied, that changed the number and the unit, but not the amount of substance (which is what you would expect upon “multiplying by 1”). 525 g of copper and 8.26 moles of copper are the same amount of copper, expressed in two different units.

A.5 COUNTING INDIVIDUAL ATOMS

Avogadro’s number, which is \(6.02 \times 10^{23}\) mol\(^{-1}\), is the number of items present in a mole. Whether you are counting individual atoms, molecules, or trees, Avogadro’s number can be used to convert between the number of items and the number of moles of items; i.e., it is just another conversion factor.

\[
\begin{align*}
6.02 \times 10^{23} \text{ atoms} & \quad 6.02 \times 10^{23} \text{ molecules} & \quad 6.02 \times 10^{23} \text{ trees} \\
1 \text{ mole of atoms} & \quad 1 \text{ mole of molecules} & \quad 1 \text{ mole of trees}
\end{align*}
\]

Example 3

How many calcium atoms are in 2.25 moles of calcium?

Solution:
Start with the known information and apply Avogadro’s number as a conversion factor:

\[
2.25 \text{ mol Ca} \times \frac{6.02 \times 10^{23} \text{ Ca atoms}}{1 \text{ mol Ca}} = 1.35 \times 10^{24} \text{ Ca atoms}
\]

Example 4

How many calcium atoms are in 2.25 grams of calcium?

Solution:
Avogadro’s number tells us how many calcium atoms are in a mole of calcium, but the given information in this example is grams of calcium. Therefore, we must use the molar mass of calcium from the periodic table (40.08 g/mol) to convert grams to moles of calcium. We then apply Avogadro’s number to obtain the number of atoms.

\[
2.25 \text{ g Ca} \times \frac{1 \text{ mol Ca}}{40.08 \text{ g Ca}} \times \frac{6.02 \times 10^{23} \text{ Ca atoms}}{1 \text{ mol Ca}} = 3.38 \times 10^{22} \text{ Ca atoms}
\]

Comment:
Here we have combined two separate calculations (grams to moles and moles to number of atoms) by stringing together two conversion factors. We could, of course, have done the two separate calculations on two separate lines.

Example 5

What is the mass of 1.00\(\times\)10\(^{22}\) bromine atoms?

Solution:
We apply Avogadro’s number to determine the number of moles of bromine in the given number of atoms. Next we use, the molar mass of bromine (79.90 g/mol from the periodic table), to convert moles into grams.

\[
1.00 \times 10^{22} \text{ Br atoms} \times \frac{1 \text{ mol Br}}{6.02 \times 10^{23} \text{ Br atoms}} \times \frac{79.90 \text{ g Br}}{1 \text{ mol Br}} = 1.33 \text{ g Br}
\]

Comment:
Again, two separate calculations were combined in one step.
A.6 CHEMICAL FORMULAS OF COMPOUNDS

A compound is a pure substance that is made up of more than one element. Compounds can be ionic (CAMS Chapter 4) or covalent as described in (CAMS Chapter 5). Covalent compounds are said to be molecular because they exist as discrete molecules, but ionic compounds exist as extended three-dimensional arrays of ions and not as discrete molecules.

The molecular formula of a compound tells us how many atoms of each element are in one molecule. A carbon dioxide molecule, which has the formula CO₂, contains one carbon atom and two oxygen atoms. This information is contained in the subscripts after each element. A molecule of sucrose (C₁₂H₂₂O₁₁) has 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms. The subscripts also indicate the ratios of the elements. A dozen CO₂ molecules have one dozen carbon atoms and two dozen oxygen atoms. A million CO₂ molecules have one million carbon atoms and two million oxygen atoms. A mole of CO₂ molecules (we usually just say “a mole of CO₂”) has one mole of carbon atoms and two moles of oxygen atoms. The atom ratio and the mole ratio of the elements are identical!

The chemical formula of an ionic compound does not tell us the number of atoms in a molecule because ionic substances are not molecular. However, it still gives the mole ratio of the elements. One mole of NaCl contains one mole of Na⁺ ions and one mole of Cl⁻ ions. Ionic compounds with polyatomic ions are somewhat more complicated. One mole of sodium sulfate, Na₂SO₄, contains two moles of Na⁺ ions and one mole of SO₄²⁻ ions or two moles of sodium, one mole of sulfur and four moles of oxygen. An additional complexity comes from the way we write formulas of compounds containing polyatomic ions. Iron (III) nitrate has the formula Fe(NO₃)₃. This tells us that for every iron(III) ion (Fe³⁺), there are three nitrate ions (NO₃⁻). Each nitrate ion contains one nitrogen atom and three oxygen atoms. Therefore, one mole of iron(III) nitrate contains one mole of iron, three moles of nitrogen and nine moles of oxygen.

A.7 MOLAR MASSES, MOLECULAR WEIGHTS AND FORMULA WEIGHTS OF COMPOUNDS

The molar mass of a compound can be determined from its chemical formula and the periodic table. The number obtained is sometimes referred to as the molecular weight or the formula weight. Molecular weight refers to the weight or mass of one molecule, in units of amu. Formula weight refers to the same quantity, but can be applied to substances that are not molecular, e.g., NaCl. Thus, even though iron(III) nitrate is not molecular, we can still talk about the mass of one formula unit, that is, one iron, three nitrogens and nine oxygens. All three terms are often used interchangeably, although molar mass is technically the only one that is in units of grams/mole.

To calculate a molar mass, we simply sum up the contributions of each element or atom. For carbon dioxide, CO₂, one carbon atom contributes 12.01 g/mol, the two oxygens together contribute (2)(16.00) = 32.00 g/mol. The molar mass is then 12.01 + 32.00 = 44.01 g/mol.

The same procedure is followed, whether the compound is covalent or ionic, as shown in the following examples.

Example 6

Calculate the molar mass of sucrose, C₁₂H₂₂O₁₁.

Solution:

Find the molar mass of each element on the periodic table and sum up the contributions.

- Contribution from C: (12 mol)(12.01 g/mol) = 144.12 g
- Contribution from H: (22 mol)(1.008 g/mol) = 22.18 g
- Contribution from O: (11 mol)(16.00 g/mol) = 176.00 g
  
  Total = 342.30 g/mol

Comment:

Notice that the molar mass of each element was multiplied by the number of times that element appeared in the chemical formula (as indicated by the subscripts in the formula).
Example 7

Calculate the molar mass of cobalt(II) phosphate, \( \text{Co}_3(\text{PO}_4)_2 \).

Solution:
Determine the contribution from each element and sum.

- Cobalt(II) phosphate:
  - Co: \((3 \text{ mol})(58.93 \text{ g/mol}) = 176.79 \text{ g}\)
  - P: \((2 \text{ mol})(30.97 \text{ g/mol}) = 61.94 \text{ g}\)
  - O: \((8 \text{ mol})(16.00 \text{ g/mol}) = 128.00 \text{ g}\)
  - Total = 366.73 g/mol

Comment:
Even though the compound contains cobalt(II) ions, we use the molar mass of cobalt atoms. The difference in mass between a cobalt atom and a cobalt(II) ion is negligible, because the mass of an electron is so small compared to the mass of an atom.

A.8 RELATING GRAMS, MOLES AND MOLAR MASS OF COMPOUNDS

The molar mass of a compound can be used as a conversion factor in the same way as the molar mass of an element.

Example 8

How many grams of sucrose (\( M_m = 342.30 \text{ g/mol} \)) are present in 0.125 moles of sucrose?

Solution:
We start with the given information and apply the molar mass as a conversion factor.

\[
0.125 \text{ mol} \times \frac{342.30 \text{ g}}{1 \text{ mol}} = 42.8 \text{ g}
\]

Example 9

What is the mass of 2.50 moles of \( \text{Mg(ClO}_4)_2 \)?

Solution:
In order to convert moles to grams, we need a molar mass from the chemical formula and the periodic table. Then we apply the molar mass as a conversion factor.

Molar mass:
- Mg: \((1 \text{ mol})(24.31 \text{ g/mol}) = 24.31 \text{ g}\)
- Cl: \((2 \text{ mol})(35.45 \text{ g/mol}) = 70.90 \text{ g}\)
- O: \((8 \text{ mol})(16.00 \text{ g/mol}) = 128.00 \text{ g}\)
  - Total = 223.21 g/mol

\[
2.50 \text{ mol} \times \frac{223.21 \text{ g}}{1 \text{ mol}} = 558 \text{ g}
\]

Comment:
This example involves putting together several individual skills that you have learned to solve a problem. Rarely does an experiment or problem in science require only one skill. Often new discoveries and new applications result from putting known information together in new ways!

A.9 PERCENT COMPOSITION OF COMPOUNDS

The method we have used to calculate molar masses gives us a simple way to figure out the percentage of each element in a compound. Let’s use ammonium nitrate as an example. This compound has the formula \( \text{NH}_4\text{NO}_3 \) and its molar mass is:

- N: \((2 \text{ mol})(14.01 \text{ g/mol}) = 28.02 \text{ g}\)
- H: \((4 \text{ mol})(1.008 \text{ g/mol}) = 4.03 \text{ g}\)
- O: \((3 \text{ mol})(16.00 \text{ g/mol}) = 48.00 \text{ g}\)
  - Total = 80.05 g/mol

This calculation not only tells us the mass of a mole of the compound, but it also tells us how many grams of each element are in a mole of the compound. For instance, in every 80.05 grams of the compound, there are 28.02 grams of nitrogen. Taking the ratio of grams of nitrogen to total grams of compound, we find:

\[
\frac{28.02 \text{ g}}{80.05 \text{ g}} = 0.3500
\]

This number, 0.3500, is the mass fraction of nitrogen in the compound. This can easily be converted into a percent by multiplying by 100:

\[
(0.3500)(100) = 35.00 \%
\]

Ammonium nitrate is 35.00% nitrogen by mass.
The ratio of grams of nitrogen to total grams of compound can be used as a conversion factor, too. If we want to know how many grams of nitrogen there are in 275 grams of ammonium nitrate, we would do the following:

\[
275 \text{ g } \text{NH}_4\text{NO}_3 \times \frac{28.02 \text{ g } \text{N}}{80.05 \text{ g } \text{NH}_4\text{NO}_3} = 96.3 \text{ g } \text{N}
\]

Notice that grams of NH\(_4\)NO\(_3\) cancel out in this calculation, leaving grams of N in the product. Also notice that the ratio 28.02/80.05 is identical to the ratio 96.3/275 (which is identical to 35/100 from the percent composition calculation). This technique is summed up in the following example.

**Example 10**

a) What is the percent nitrogen in potassium nitrate (KNO\(_3\))?  

Solution: 

Percent nitrogen can be found from the data in a molar mass calculation. First, we need the chemical formula of potassium nitrate.

Potassium nitrate: KNO\(_3\)  

Molar mass:

- K: \((1 \text{ mol})(39.10 \text{ g/mol}) = 39.10 \text{ g}\)
- N: \((1 \text{ mol})(14.01 \text{ g/mol}) = 14.01 \text{ g}\)
- O: \((3 \text{ mol})(16.00 \text{ g/mol}) = 48.00 \text{ g}\)

Total = 101.11 g/mol

Percent nitrogen:

\[
\frac{14.01 \text{ g } \text{N}}{101.11 \text{ g } \text{total}} \times 100\% = 13.86 \% \text{ N}
\]

b) How many grams of potassium are present in 137 g of KNO\(_3\)?

Solution: 

The molar mass calculation tells us that, for every 101.11 g of compound, there are 39.10 g of K. Start with the given information and apply the ratio of grams of K to grams of compound as a conversion factor.

\[
137 \text{ g } \text{KNO}_3 \times \frac{39.10 \text{ g } \text{K}}{101.11 \text{ g } \text{KNO}_3} = 53.0 \text{ g } \text{K}
\]

A.10 COUNTING ATOMS, IONS AND MOLECULES IN COMPOUNDS

We use Avogadro’s number, \(6.02 \times 10^{23} \text{ mol}^{-1}\), to “count” the number of individual particles in a sample. The key in our calculations is to first find the number of moles of whatever item it is we wish to count, and then apply Avogadro’s number to convert from moles to individual items. The following examples all involve multiple steps, each of which we have discussed separately.

**Example 11**

a) How many water molecules are contained in 10.0 g of H\(_2\)O?

Solution: 

In order to “count” water molecules, we first find how many moles of water are present in 10.0 g (\(M_m = 18.02 \text{ g/mol}\)), and then apply Avogadro’s number.

\[
10.0 \text{ g } \text{H}_2\text{O} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{6.02 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 3.34 \times 10^{23} \text{ molecules}
\]
b) How many hydrogen atoms are contained in 10.0 g of H₂O?

Solution to part b:

In order to “count” hydrogen atoms, we first find how many moles of hydrogen atoms are present in 10.0 g of water, and then apply Avogadro’s number.

\[
\frac{10.0 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{6.02 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} = 6.68 \times 10^{23} \text{ H atoms}
\]

Comment:

Note that the two calculations are almost identical, the only difference being the term that converts moles of H₂O to moles of H in part b. This extra step was necessary because we are “counting” hydrogen atoms, not water molecules. The conversion factor comes directly from the chemical formula, where the subscript 2 (after the H) indicates that there are two hydrogen atoms per water molecule and two moles of hydrogen atoms per mole of water molecules. Note also that each calculation simply strings together a series of conversion factors. Each conversion factor is applied such that the term in the denominator “cancels out” the unit from the previous step. For example, the first conversion factor in each part is the molar mass of water simply turned “upside down” to put grams of water in the denominator.

Example 12

How many iron(III) ions are contained in 68.4 g of Fe₂(SO₄)₃?

Solution:

Many chemistry students find that the hardest part of a problem like this is figuring out where to start. Let’s use some stepwise logic to figure out what we need for each step, working backwards to see where we should start. In order to count iron(III) ions, we need moles of iron(III). We can get moles of iron(III) if we know how many moles of iron(III) sulfate we have. We can get the moles of iron(III) sulfate from the 68.4 g and the molar mass. The molar mass is obtained from the chemical formula.

Molar mass:

- Fe: \((2 \text{ mol})(55.85 \text{ g/mol}) = 111.70 \text{ g})
- S: \((3 \text{ mol})(32.07 \text{ g/mol}) = 96.21 \text{ g})
- O: \((12 \text{ mol})(16.00 \text{ g/mol}) = 192.00 \text{ g})

Total = 399.91 g/mol

\[
68.4 \text{ g Fe}_2(\text{SO}_4)_3 \times \frac{1 \text{ mol Fe}_2(\text{SO}_4)_3}{399.91 \text{ g Fe}_2(\text{SO}_4)_3} \times \frac{2 \text{ mol Fe}^{3+}}{1 \text{ mol Fe}_2(\text{SO}_4)_3} \times \frac{6.02 \times 10^{23} \text{ Fe}^{3+} \text{ ions}}{1 \text{ mol Fe}^{3+} \text{ ions}} = 2.06 \times 10^{23} \text{ Fe}^{3+} \text{ ions}
\]

Comment:

As is often the case, there are several steps required to solve the problem. Each individual step is not that hard; it is putting them together in the proper order that is the stumbling block for some students. Our approach was to use some logic, starting at the end and working back to the beginning, keeping track of everything needed along the way to solve the problem.

Example 13

How many grams of sucrose \((C_{12}H_{22}O_{11})\) contain \(4.75 \times 10^{25}\) carbon atoms?

Solution:

We are given the number of carbon atoms and are asked to find the mass of sucrose that contains them. (Notice that this problem works in the reverse direction to Examples 11 and 12, where grams were given and individual atoms were sought.) In order to calculate grams of sucrose, we need moles of sucrose \((342.30 \text{ g/mol from Example 6})\). Moles of sucrose can be found from the moles of carbon atoms and Avogadro’s number.

\[
4.75 \times 10^{25} \text{ C atoms} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23} \text{ C atoms}} \times \frac{1 \text{ mol C}_{12}H_{22}O_{11}}{12 \text{ mol C}} \times \frac{342.30 \text{ g C}_{12}H_{22}O_{11}}{1 \text{ mol C}_{12}H_{22}O_{11}} = 2.25 \times 10^3 \text{ g} = 2.25 \text{ kg C}_{12}H_{22}O_{11}
\]

Comment:

As always, each conversion factor we apply has the effect of “cancelling out” the units from the previous step.
A.11 EXERCISES

Basic skills / elements
1. Arrange each of the following sets of elements in order of increasing atomic mass.
   a) phosphorus, sodium, iron, carbon
   b) manganese, potassium, fluorine, copper
   c) selenium, beryllium, arsenic, iron
   d) chlorine, zinc, scandium, helium
2. How many moles of titanium are contained in 15.5 g of titanium?
3. How many grams of sodium are contained in 1.25 moles of sodium?
4. How many moles of vanadium does $6.02 \times 10^{22}$ vanadium atoms represent?
5. How many grams of magnesium are contained in 0.52 moles of magnesium?
6. How many moles of cobalt are contained in 66 kg of cobalt?
7. How many nickel atoms are contained in 0.50 moles of nickel?
8. How many moles of carbon are contained in 2.85 g of carbon?
9. How many moles of nitrogen does $7.5 \times 10^{21}$ nitrogen atoms represent?
10. How many moles of germanium are contained in 25 mg of germanium?
11. How many oxygen atoms are contained in 6.25 moles of oxygen atoms?
12. How many grams of chromium are contained in $2.5 \times 10^{-4}$ moles of chromium?
13. How many krypton atoms are contained in 1.22 moles of krypton?
14. How many kilograms of phosphorus atoms are contained in 38 moles of phosphorus atoms?
15. How many moles of silicon does $8.8 \times 10^{24}$ silicon atoms represent?

Combined skills / elements
16. How many lithium atoms are contained in 1.0 g of lithium?
17. What is the mass of $2.5 \times 10^{21}$ argon atoms?
18. How many aluminum atoms are contained in 1.5 kg of aluminum.
19. What is the mass of $3.5 \times 10^{25}$ iron atoms?
20. What is the mass, in grams, of $8.25 \times 10^{23}$ silver atoms?
21. How many barium atoms are contained in 0.050 g of barium?

Basic skills / compounds
22. Calculate the molar mass of the following compounds:
   a) hydrazine, N$_2$H$_4$
   b) acetic acid, HC$_2$H$_3$O$_2$
   c) pyridine, C$_5$H$_5$N
   d) succinic acid, C$_4$H$_6$O$_4$
   e) TNT, C$_7$H$_5$N$_3$O$_6$
   f) calcium nitrate, Ca(NO$_3$)$_2$
   g) potassium chromate, K$_2$CrO$_4$
   h) cobalt(II)citrate, Co$_3$(C$_6$H$_8$O$_7$)$_2$
23. Use the following molar masses to answer these questions:
   cobalt(II) iodide, CoI$_2$ : 312.74 g/mol
   morphine, C$_{17}$H$_{19}$NO$_3$ : 285.35 g/mol
   a) How many moles of CoI$_2$ are contained in 10.0 g of CoI$_2$?
   b) How many grams of morphine are contained in $2.0 \times 10^{-5}$ moles of morphine?
   c) How many moles of morphine are contained in 35 mg of morphine?
   d) How many kilograms of CoI$_2$ are contained in 12.0 moles of CoI$_2$?
24. What is the % carbon in each of the following compounds:
   a) CH$_4$
   b) C$_3$H$_4$O$_6$
   c) C$_3$H$_8$
25. How many grams of potassium are contained in 8.00 g of KCl?
26. How many chloride ions are contained in 8.00 g of KCl?
27. How many grams of sulfur are contained in 1.00 kg of thiophene, C$_4$H$_4$S (M$_m$ = 84.14 g/mol)?
28. How many sulfur atoms are contained in 3.55 g of C$_4$H$_4$S?
29. How many carbon atoms are contained in 3.55 g of C$_4$H$_4$S?
30. If you wanted to obtain $1.00 \times 10^5$ g of nitrogen, what mass of NH$_3$ would you need?
31. If you wanted to obtain $1.00 \times 10^5$ g of nitrogen, what mass of NH$_4$NO$_3$ would you need?
32. How many oxygen atoms are contained in 5.25 g of Fe(NO$_3$)$_3$?
33. How many carbon atoms are contained in 65 g of quinine, C$_{20}$H$_{24}$N$_2$O$_2$ (M$_m$ = 324.41 g/mol)?
34. How many grams of carbon are contained in 65 g of quinine?
Answers:

1. a) C(12) < Na(23) < P(31) < Fe(56)
   b) F(19) < K(39) < Mn(55) < Cu(64)
   c) Be(9) < Fe(56) < As(75) < Se(79)
   d) He(4) < Cl(35) < Sc(45) < Zn(65)

2. 0.324 mol Ti
3. 28.8 g Na
4. 0.100 mol V
5. 13 g Mg
6. 1.1x10^3 mol Co
7. 3.0x10^{23} atoms of Ni
8. 0.237 mol C
9. 0.012 mol N (or 0.0062 mol N_2)
10. 3.4x10^{-4} mol Ge = 0.34 mmol Ge
11. 3.76x10^{24} atoms of O
12. 0.013 g Cr
13. 7.35x10^{23} atoms of Kr
14. 1.2 kg of P atoms
15. 15 mol Si
16. 8.7x10^{22} atoms of Li
17. 0.17 g Ar
18. 3.3x10^{23} atoms of Al
19. 3.2 kg Fe
20. 148 g Ag
21. 2.2x10^{20} atoms of Ba
22. a) 32.05 
   b) 60.34
   
c) 1.2x10^{-4} g morphine
   
d) 3.0x10^{23} atoms of Ni
23. a) 0.0320 mol CoI_2
   
b) 0.0057 mol morphine
   
c) 79.1 
   
d) 560.9
24. a) 75.0% 
   b) 40.0% 
25. 4.19 g K
26. 6.46x10^{22} Cl^- ions
27. 381 g S
28. 2.54x10^{22} atoms of S
29. 1.02x10^{21} atoms of C
30. 122 kg NH_3
31. 286 kg NH_4NO_3
32. 1.18x10^{23} atoms of O
33. 2.4x10^{24} atoms of C
34. 48 g C

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Appendix B
Gases

B.1 GAS MOLECULES ACTING COLLECTIVELY

According to the Kinetic Molecular Theory (CAMS Section 7.2), gases are in constant random motion and the average kinetic energy is proportional to the absolute temperature. The kinetic energy of a molecule is \( \frac{1}{2}mv^2 \), so the average speed of the molecules also depends on the absolute temperature. The average O\(_2\) molecule moves at about 1,000 mph on a nice day.

However, it is the collective action of large numbers of molecules that we sense or measure as gases not the individual molecules. When you fan your face, you feel some wind, which is the effect of molecules in the air hitting your face. You cannot sense the individual molecules hitting your skin, for they are much too small, but you can feel their collective action.

The fact that a balloon expands when it is filled with a gas also shows how gas molecules act collectively. The molecules in the balloon are moving around with an average kinetic energy dictated by the temperature. When a molecule strikes the inside wall of the balloon, it exerts a force on the balloon and pushes it outward. The collective forces of all of the molecules inside the balloon pushing outward cause it to stay inflated. At the same time, the gas molecules in the outside air are striking the outer surface of the balloon exerting a force pushing inward. The size of the balloon adjusts until the force from the “strikes” on the outside balances the force from the “strikes” on the inside.

The collective force of all of the molecules pushing on the inside wall of the balloon results in pressure. The collective force per unit area or pressure of the gas depends on the number of collisions with the walls per second and the force of each collision. The common units are pounds per square inch, atmosphere (atm), the millimeter of mercury (mm Hg or torr) and the pascal, the SI unit of pressure (N/m\(^2\)).

B.2 RELATIONSHIP OF PRESSURE TO OTHER GAS PROPERTIES

Let’s analyze what happens to the pressure of a gas as the temperature, the number of molecules and volume of the gas change. Imagine a cylinder with a movable piston. The gas molecules in the piston have kinetic energy (are moving) and are hitting the walls of the cylinder, the piston and each other.

If the temperature of the gas is increased, the molecules will move faster and will strike the piston more frequently and with more force. Consequently the pressure increases. If more molecules are added to the cylinder (moles of gas increase), the frequency of collisions and therefore the pressure increases. Finally, if we push the piston down and compress the gas to a smaller volume, the gas molecules have less distance to travel before they hit the piston, and they collide with the piston more frequently. Thus, a decrease in volume will result in an increase in pressure.

The relationships among the pressure, volume, number of moles, and temperature of a gas are summed up quantitatively in the ideal gas law:

\[
P V = n RT
\]

Where \( P \) is the pressure in atmospheres (atm), \( V \) the volume in liters (L), \( n \) the number of moles, \( T \) the absolute temperature in kelvins, and \( R \) is a constant called the ideal gas constant, which is 0.0821 L-atm-K\(^{-1}\)-mol\(^{-1}\). However, when using SI units, \( P \) is expressed in pascals, \( V \) in m\(^3\), and \( R = 8.314 \text{ J.mol}^{-1}.\text{K}^{-1} \).

B.3 USING THE IDEAL GAS LAW

The ideal gas law contains four experimental quantities: pressure, volume, temperature, and number of moles. If we know three of the quantities, we can solve for the fourth. The first step is always to ensure that the units on the known quantities are consistent with our value of \( R \). The following examples show how this can be done.
Example 1

What is the volume of 1.00 mole of gas at 1.00 atm and 0 °C?

Solution:
n is in mol and P in atm, but T is in °C, not K. To convert from °C to K, we add 273: 0 °C + 273 = 273 K = T. Next, rearrange the ideal gas law to solve for the unknown, which in this case is the volume.

\[ V = \frac{nRT}{P} = \frac{(1.00 \text{ mol}) (0.0821 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}) (273 \text{ K})}{1.00 \text{ atm}} = 22.4 \text{ L} \]

Comment:
The conditions 0 °C and 1.00 atm are often referred to as the Standard Temperature and Pressure (STP) for a gas. The volume at STP is 22.4 L, which is an experimental, not theoretical, number that students often remember from high school chemistry.

Example 2

An experiment yields 5.67 mL of CO₂(g) at 26 °C and 782 mm Hg. How many grams of CO₂ is this?

Solution:
We are given V, T, and P, but none have the correct units for our value of R, so we convert each into the proper units.

\[ V = 5.67 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00567 \text{ L} \]
\[ T = 26 \text{ °C} + 273 = 299 \text{ K} \]
\[ P = 782 \text{ mm Hg} \times \frac{1.00 \text{ atm}}{760 \text{ mm Hg}} = 1.03 \text{ atm} \]

Next, rearrange the ideal gas law to solve for n.

\[ n = \frac{PV}{RT} = \frac{(1.03 \text{ atm}) (0.00567 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}) (299 \text{ K})} = 2.38 \times 10^{-4} \text{ mol CO}_2 \]

To find grams, apply the molar mass (M_m = 44.01 g/mol) as a conversion factor as done in Appendix A,

\[ 2.38 \times 10^{-4} \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{\text{mol CO}_2} = 0.0105 \text{ g CO}_2 \]

Comment:
Notice that five and two-thirds mL of a gas seems like a very small amount in terms of mass! That is because we are used to weighing out solids and liquids, which are much denser than gases.

B.4 EXERCISES

1. What volume does 0.50 moles of CO₂ occupy at 725 mm Hg and 25 °C?
2. How many moles of He occupy a 2.50-L flask whose pressure is 945 mm Hg at 75 °C?
3. What is the pressure exerted by 28.8 g of N₂ contained in a 4.25 L-flask at 0 °C?
4. What volume does 5.8 moles of O₂ occupy at 285 mm Hg and -78 °C?
5. What is the temperature of 5.0 moles of N₂ contained in a 20.0 L-tank at a pressure of 7.5 atm?
6. What volume does 6.32 g of NH₃ occupy at 745 mm Hg and 25 °C?
7. How many moles of CH₄ occupy a 10.0-L tank whose pressure is 3.5 atm at 30 °C?
8. What volume does 0.45 g of Ar occupy at 1.25 atm and 27 °C?
9. What is the pressure exerted by 3.5 moles of H₂ contained in a 2.0-L tank at 27 °C?
10. What volume does 0.75 moles of N₂ occupy at 760 mm Hg and 0 °C?
11. What is the temperature of 7.65 g of He contained in a 6.25 L flask at a pressure of 1.75 atm?
12. How many moles of HCl gas occupy a 4.5 L tank whose pressure is 1875 mm Hg at 27 °C?
13. For this question, note that M_m = g/mol and density, d = mass/volume.
   a) What is the density of helium in g/L at 1.00 atm and 27 °C?
   b) What is the density of nitrogen in g/L at 1.00 atm and 27 °C?

ANSWERS:

1. 13 L  
2. 0.109 mol  
3. 5.42 atm  
4. 2.5 × 10² L  
5. 365 K = 92 °C  
6. 9.26 L  
7. 1.4 mol  
8. 0.22 L  
9. 43 atm  
10. 17 L  
11. 69.7 K = -203 °C  
12. 0.45 mol  
13. a) 0.163 g/L  
13. b) 1.14 g/L
Appendix C

Molarity

C.1 MOLARITY AND THE MOLE

The molar mass is the mass of a mole of a pure substance while the molarity, M, is the number of moles of a pure substance contained in a liter of a solution.

\[ \text{molarity} = \frac{\text{moles}}{\text{liter}} = \frac{n}{V} \]

One liter of a solution that contains 0.1 moles of sugar (C\(_{12}\)H\(_{22}\)O\(_{11}\)) is 0.1 M, or the solution is 0.1 molar in sugar. It can also be represented as: \([\text{C}_{12}\text{H}_{22}\text{O}_{11}] = 0.1 \text{ M}\), which is read as “the molar concentration of sugar is 0.1 molar.”

C.2 MOLARITY AS A CONVERSION FACTOR

Molarity is used to convert between moles of substance and liters of solution.

Example 1

How many moles of NaCl are in 325 mL of 0.25 M NaCl solution?

Solution:

We first convert mL to L, and then apply molarity as a conversion factor.

\[ 325 \text{ mL} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{0.25 \text{ moles NaCl}}{1 \text{ L}} = 0.081 \text{ mol NaCl} \]

Example 2

How many mL of 5.0 M HCl contains 0.15 moles of HCl?

Solution:

Our known quantities are moles of HCl and molarity. We start with moles and apply molarity as a conversion factor. The final step is to convert liters to milliliters.

\[ 0.15 \text{ mol HCl} \times \frac{1 \text{ L}}{5.0 \text{ mol HCl}} \times \frac{10^3 \text{ mL}}{1 \text{ L}} = 30 \text{ mL} \]

Comment:

Note that our definition of molarity is turned upside down, and we were careful to write the units L and moles HCl in the numerator and denominator. Do not use M as the units of the conversion factor.

C.3 CONCENTRATIONS OF IONS

When ionic compounds dissolve, individual solvated ions are formed. (Recall that there are no molecules in ionic compounds.) When we refer to a 0.1 M NaCl solution, we mean that the solution has 0.1 moles of NaCl units in every liter. We can also determine the concentrations of the individual ions from the chemical formula. As discussed in Appendix A, the chemical formula relates moles of compound to moles of each element in the compound. In one mole of Na\(_2\)SO\(_4\) there are two moles of Na\(^{1+}\) ions and one mole of SO\(_{4}^{2-}\) ions.

Example 3

a) What is the concentration of Na\(^{1+}\) ions in 0.25 M NaCl?

b) What is the total concentration of all ions in the above solution?

Solution:

Part a), knowing the molarity of the compound and the formula, we can easily see that for every mole of NaCl, there is one mole of Na\(^{1+}\).

\[ \frac{0.25 \text{ mol NaCl}}{1 \text{ L}} \times \frac{1 \text{ mol Na}^{1+}}{1 \text{ mol NaCl}} = 0.25 \text{ M Na}^{1+} \]

Part b), for every mole of NaCl, there is one mole of Na\(^{1+}\) ions and one mole of Cl\(^{-}\) ions, which adds up to two moles total of ions. Again, start with the concentration of the compound and find the concentration of ions.

\[ \frac{0.25 \text{ mol NaCl}}{1 \text{ L}} \times \frac{2 \text{ mol ions}}{1 \text{ mol NaCl}} = \frac{0.50 \text{ mol ions}}{1 \text{ L}} = 0.50 \text{ M ions} \]
Comment:
In each step, we have used the moles of ions per mole of compound as a conversion factor much like we did in Examples 11-13 in Appendix A. Notice that this conversion gets us directly to the molarity of the ions (moles of ions per liter).

Example 4
a) What is the concentration of chloride ions in a 0.1 M CaCl₂?

b) What is the total concentration of ions in a 0.1 M solution of CaCl₂?

Solution:
As in Example 3, we start with the solution concentration and apply a conversion factor that converts moles of compound to moles of individual ions. In CaCl₂, the chemical formula tells us that there are two moles of Cl⁻ in every mole of compound, and three moles of total ions (1 mole Ca²⁺, 2 moles Cl⁻) in every mole of compound.

\[
\frac{0.1 \text{ mol CaCl}_2}{1 \text{ L}} \times \frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} = 0.2 \text{ mol Cl}^- = 0.2 \text{ M Cl}^-
\]

\[
\frac{0.1 \text{ mol CaCl}_2}{1 \text{ L}} \times \frac{3 \text{ mol ions}}{1 \text{ mol CaCl}_2} = 0.3 \text{ mol ions} = 0.3 \text{ M ions}
\]

Comment:
We have taken some care to write down the units in detail for each conversion, but once you understand chemical formulas and the fact that ionic compounds dissolve to form individual ions, you will be able to do these calculations in your head!

C.4 MAKING SOLUTIONS

One of the most common tasks in the chemistry laboratory is making solutions of desired concentrations. In this section, we will explore how to make solutions starting with a solid solute.

Example 5
How many grams of Na₂CO₃ (Mₘ = 105.99 g/mol) are required to make 0.500 L of a 0.10 M Na₂CO₃ solution?

Solution:
We have a target volume and molarity for our solution, and so we can calculate the necessary moles of Na₂CO₃. We can then use the molar mass to calculate the necessary grams of Na₂CO₃.

\[
0.500 \text{ L} \times \frac{0.10 \text{ mol}}{1 \text{ L}} \times \frac{106 \text{ g}}{1 \text{ mol}} = 5.3 \text{ g}
\]

Example 6
What is the molar concentration of a 2.5 L of solution that contains 254 g of Na₂CO₃?

Solution:
In this problem, we have a known mass of solid and a molar mass, enough information to calculate moles of Na₂CO₃. We can then use the relationship between moles and volume to calculate molarity.

\[
\frac{254 \text{ g}}{106 \text{ g}} = 2.40 \text{ mol}
\]

\[
\text{concentration} = \frac{2.40 \text{ mol}}{2.5 \text{ L}} = 0.96 \text{ M}
\]

Comment:
Note that Examples 5 and 6 start from opposite ends of the same type of calculation. In each case, we have enough information to calculate moles of Na₂CO₃. In Example 5, we had a target volume and molarity; in Example 6, we had a mass and a molar mass. Determining the amount of solute that is needed to make a desired solution or the concentration of a particular solution by knowing how it was made are two types of calculations that are performed routinely in the chemistry laboratory.

C.5 DILUTION OF SOLUTIONS

In the previous section, solution concentrations were related to the mass of the solute. That type of calculation is appropriate when the solutes come from a chemical supply house in solid form. Some compounds are supplied as concentrated solutions. HCl is a good example. Most HCl in the laboratory is
purchased as ‘concentrated hydrochloric acid’, which is often called a “stock solution”. The dilution of stock solutions to give new solutions of desired concentrations is another very common laboratory procedure. The quantitative aspects will be detailed here.

When calculating molarity and volume of diluted solutions, we can take a shortcut if we are simply diluting with pure solvent. (Be careful, this shortcut does not work for experiments where you dilute with another solution, see Example 9, or for reaction stoichiometry, see Appendix D.) The shortcut is based on the idea that in diluting a concentrated solution with pure solvent, you are not changing the number of moles of solute. The molarity changes of course, because the volume changes. Since the number of moles of solute in the concentrated stock solution (n_c) equals the number of moles of solute in the diluted solution (n_d), we can write that n_c = n_d.

Rearranging the relationship M = n/V, we find that n = MV, so:

\[ M_c V_c = M_d V_d \]

The only restriction on the units of the volumes is that they must be the same.

### Example 7

**What is the concentration of the solution prepared by diluting 25 mL of 12 M HCl to 1.0 L with pure water?**

**Solution:**

The volume of the concentrated solution (before dilution) is 25 mL, and its concentration is 12.0 M solution. The volume of the diluted solution is 1.0 L, but its concentration is unknown. Remember that the volumes must have the same units. Using our shortcut, we write:

\[ (12 \text{ M}) (25 \text{ mL}) = M_d (1000 \text{ mL}) \]

\[ M_d = 0.30 \text{ M} \]

**Comment:**

Notice that, upon rearranging the equation to solve for M_d, the mL units cancel out. Whether we use 25 and 1000 mL or 0.025 and 1.0 L for the two volumes, the results are the same. Also, we write the final result as 0.30 M. We could have written it out the long way, 0.30 moles/liter.

### Example 9

15 mL of a 12 M solution of HCl was diluted with 100 mL of a 0.50 M solution of HCl. What is the concentration of the resulting solution? Assume that the volumes are additive.

**Solution:**

Note here that we are not diluting a stock solution with pure solvent as was done in Examples 7 and 8. We will not be able to use our shortcut because both the 12.0 M and the 0.50 M solutions contribute some moles of HCl to the final solution. Instead, we add up the total number of moles of HCl and divide by the total volume.

The first solution contributes:

\[ 15 \text{ mL} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{12.0 \text{ mol HCl}}{1 \text{ L}} = 0.18 \text{ mol HCl} \]

The second solution contributes:

\[ 100 \text{ mL} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{0.50 \text{ mol HCl}}{1 \text{ L}} = 0.050 \text{ mol HCl} \]

The total number of moles of HCl in the final solution is:

\[ 0.18 + 0.050 = 0.23 \text{ moles} \]

The total volume is: 15 mL + 100. mL = 115 mL = 0.115 L

The concentration of the final solution is:

\[ \frac{0.23 \text{ moles}}{0.115 \text{ L}} = 2.0 \text{ M} \]
Comment:
We found the total number of moles contributed by both initial solutions by first finding the number of moles of HCl in each solution separately and adding them together. Because we are not using the shortcut, volumes had to be in liters, and the appropriate conversions were applied. This problem also points out that a small volume of a more concentrated solution (the first one) often contains more of the solute than a larger volume of a less concentrated solution.

C.6 EXERCISES

1. What is the concentration of NaCl when 25.0 g of NaCl is dissolved in water to make 450. mL of solution?
2. How many mL of a 5.0 M solution of HCl contains 0.10 moles of HCl?
3. How many moles of K₂SO₄ are contained 100. mL of a 1.35 M solution?
4. What is the concentration of K¹⁺ ions in 500. mL of a 0.125 M solution of K₂SO₄?
5. How many mL of a 0.10 M solution of NaCl contains 6.2 x 10⁻³ moles of NaCl?
6. How many grams of CaCl₂ are required to make 10.0 mL of 1.00 M CaCl₂ solution?
7. How many moles of Cl¹⁻ ions are contained in 250. mL of a 0.552 M solution of MgCl₂?
8. How many mL of a 0.80 M solution of Na₂CO₃ contains 0.20 moles of Na¹⁺ ions?
9. How many moles of Li₂CO₃ are contained in 25.0 mL of a 1.15 M solution?
10. An experiment calls for 1.00 L of a 0.150 M KCl solution. How many mL of a 4.00 M stock solution of KCl must be used to prepare this solution?
11. How many moles of Cl¹⁻ ions are contained in 18.5 mL of a 1.28 M solution of NaCl?
12. What is the concentration of K¹⁺ ions in 25.0 mL of a 1.00 M solution of KCl?
13. What is the concentration of NaCl when 5.75 g of NaCl is dissolved in water to make 1.86 L of solution?
14. How many grams of LiCl are required to make 125 mL of 0.100 M LiCl solution?
15. How many mL of a 1.25 M solution of KCl contains 2.35 g of KCl?
16. How many grams of LiCl are required to make 625 mL of 2.87 M LiCl solution?
17. 10.0 mL of a 3.25 M solution of HCl is diluted with 200 mL of a 0.100 M solution of HCl. What is the concentration of the resulting solution? Assume that the volumes are additive.
18. What is the concentration of the solution prepared by diluting 25 mL of a 0.50 M solution of HCl to 125 mL with pure water?
19. What is the concentration of the solution prepared by diluting 5.0 mL of a 6.25 M solution of HCl to 65 mL with pure water?
20. How many moles of Li¹⁺ ions are contained in 0.500 L of a 2.25 M solution of Li₂CO₃?
21. An experiment calls for 125 mL of a 0.625 M HCl solution. How many mL of a 12.0 M stock solution of HCl must be used to prepare this solution?
22. 12.5 mL of a 12.0 M stock solution of HCl is diluted with 85.0 mL of a 0.200 M solution of HCl. What is the concentration of the resulting solution? Assume that the volumes are additive.

ANSWERS:

1. 0.950 M  6. 1.11 g  11. 0.0237 mol  16. 76.1 g  21. 6.51 mL
2. 20 mL  7. 0.276 mol  12. 1.00 M  17. 0.250 M  22. 1.71 M
3. 0.135 mol  8. 125 mL  13. 0.0528 M  18. 0.10 M
4. 0.250 M  9. 0.0288 mol  14. 0.530 g  19. 0.48 M
5. 62 mL  10. 37.5 mL  15. 25.2 mL  20. 2.25 mol
D.1 INTRODUCTION

In Appendix A, the stoichiometry of elements and compounds was presented. There, the relationships among grams, moles and number of atoms and molecules were reviewed. A similar relationship exists for chemical reactions, and we will now extend this concept of stoichiometry to reactions. In reaction stoichiometry, we are interested in the quantitative relationships between the amounts of reactants and products in a reaction. We will find, as we did in Appendix A, that the mole is the central character in these calculations.

D.2 QUANTITATIVE RELATIONSHIPS IN REACTIONS

Chemical reactions surround us. Chemists use shorthand notation to describe them in a sentence called a chemical equation. The chemical equation that describes the combustion of benzene is

$$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(g)$$

This equation implies that for every two benzene molecules that react, 15 dioxygen molecules must also react and 12 carbon dioxide molecules and six water molecules will be produced. In other words, it tells us about the “stoichiometry”, or the amounts of reactants and products involved. The coefficients, or numbers in front of each chemical formula, tell us the relative number of molecules involved in the reaction. They also tell us the relative number of moles involved in the reaction. Two moles of benzene will react with 15 moles of dioxygen to form twelve moles of carbon dioxide and six moles of water. It is important to note that the equation does not give us any direct information about the number of grams of each reactant or product, only the moles. If we want to know about a measurable quantity like grams, we will have to do some conversions.

The first requirement for any stoichiometric calculation is a balanced equation. Once we have a balanced equation, the calculations will be performed by following three steps.

1. Convert the given quantitative information to moles. Experiments are always set up (thus chemistry problems are always written) such that number of moles of at least one reactant or product can be determined.
2. Use the balanced equation to convert from the moles of given substance to the moles of desired substance. Remember, we’re given quantitative information on one reactant or product, we desire quantitative information on another. This is the heart of all stoichiometry problems.
3. Convert from moles of the desired compound to the appropriate quantity.

These three steps are purposely vague. In the three previous appendices, we have discussed the conversion of moles to grams for a solid, moles to pressure, volume and temperature for a gas, and moles to volume and molarity for a solution. As you might guess, there are several variations on this three-step theme. The following examples will show some of the variety.

D.3 REACTION STOICHIOMETRY INVOLVING GRAMS

For calculations involving grams of reactants and products, our three-step scheme looks like this:

In the preceding diagram, A and B are products and/or reactants. The gram to mole conversion is achieved through application of the molar mass, the mole to mole conversion comes from the coefficients in the balanced equation.
Example 1

Consider the combustion of benzene:
\[ 2\text{C}_6\text{H}_6(\ell) + 15\text{O}_2(\text{g}) \rightarrow 12\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g}) \]

What is the maximum mass of CO\(_2\) (M\(_m\) = 44.0 g/mol) that can be produced from the combustion of 10.0 g of C\(_6\)H\(_6\) (M\(_m\) = 78.1 g/mol)?

Solution:
The given information is grams of benzene. The desired information is grams of CO\(_2\). The road map above tells us we must convert grams of benzene to moles, use the mole ratio from the equation to give moles of carbon dioxide, and then convert back to grams of carbon dioxide.

\[
10.0\ \text{g C}_6\text{H}_6 \times \frac{1\ \text{mol C}_6\text{H}_6}{78.1\ \text{g C}_6\text{H}_6} = 0.128\ \text{mol C}_6\text{H}_6
\]
\[
0.128\ \text{mol C}_6\text{H}_6 \times \frac{12\ \text{mol CO}_2}{2\ \text{mol C}_6\text{H}_6} = 0.768\ \text{mol CO}_2
\]
\[
0.768\ \text{mol CO}_2 \times \frac{44.0\ \text{g CO}_2}{1\ \text{mol CO}_2} = 33.8\ \text{g CO}_2
\]

Comment:
As we have seen several times in the previous appendices, each individual step is not difficult. Putting the steps together in a logical manner is the challenge. In this example, we have done each of the three steps separately. However, we could have strung together the conversion factors to save ourselves some writing. We will do that in the next example.

Example 2

How many grams of O\(_2\) (M\(_m\) = 32.0 g/mol) are required to completely react with 10.0 g of C\(_6\)H\(_6\)?

Solution:
The given information is the mass of benzene, the desired information is the mass of dioxygen. Start with the given information and apply conversion factors following the road map.

\[
10.0\ \text{g C}_6\text{H}_6 \times \frac{1\ \text{mol C}_6\text{H}_6}{78.1\ \text{g C}_6\text{H}_6} \times \frac{15\ \text{mol O}_2}{2\ \text{mol C}_6\text{H}_6} \times \frac{32.0\ \text{g O}_2}{1\ \text{mol O}_2} = 30.7\ \text{g O}_2
\]

D.4 LIMITING REACTANTS

We know from experience that the amount of product that is formed depends on the amount of reactant that is consumed. You can drive a car only as long as it has gasoline. The gasoline is the limiting reactant because it dictates how much product (miles) can be achieved. The amount of gasoline determines not only how far you can go, but it also determines how much CO\(_2\) and H\(_2\)O (the reaction products of the combustion of gasoline) can be made.

In any chemical reaction, the amount of products that are made is limited by the amount of reactants. When any one reactant runs out, the reaction stops. The reactant that runs out is called the limiting reactant or limiting reagent. Any reactants that do not run out are said to be in excess. In most chemical reactions, one or more of the reactants is in excess. In the gasoline combustion reaction, there is certainly more oxygen available than there is gasoline in the gas tank, and so the oxygen is in excess.

In calculating the amount of product formed in a reaction, we always have to identify the limiting reactant. In some cases it is obvious. In Example 1 above, we read that 10.0 g of C\(_6\)H\(_6\) reacts with excess O\(_2\). Clearly, C\(_6\)H\(_6\) is the limiting reactant, and O\(_2\) is the excess reactant. But consider the following example.

Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas, via the following reaction

\[ \text{CaH}_2(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Ca(OH)}_2(s) + 2\text{H}_2(g) \]

If 10.0 g of CaH\(_2\) reacts with 9.00 g of H\(_2\)O, what mass of Ca(OH)\(_2\) can be formed? We start by determining the limiting reactant, but we cannot tell which
reactant will limit the amount of product just by comparing the grams of each reactant, we must determine how much Ca(OH)\(_2\) we can make from each reactant. Based on the amount of CaH\(_2\) we start with, we can make:

\[
10.0 \text{ g CaH}_2 \times \frac{1 \text{ mol CaH}_2}{42.10 \text{ g CaH}_2} \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 17.6 \text{ g Ca(OH)}_2
\]

Based on the amount of H\(_2\)O available, we can make:

\[
9.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \times \frac{1 \text{ mol Ca(OH)}_2}{2 \text{ mol H}_2\text{O}} \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 18.5 \text{ g Ca(OH)}_2
\]

Even though we have enough water to make 18.5 g of Ca(OH)\(_2\), there is only enough calcium hydride to make 17.6 g. In this case, CaH\(_2\) is the limiting reactant, H\(_2\)O is in excess, and 17.6 g of Ca(OH)\(_2\) would be produced.

Let’s determine how much of the excess reactant remains. CaH\(_2\) is the limiting reactant and all amounts are calculated from it, and so we must now determine how much water reacts with the CaH\(_2\).

\[
10.0 \text{ g CaH}_2 \times \frac{1 \text{ mol CaH}_2}{42.10 \text{ g CaH}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol CaH}_2} \times \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 8.56 \text{ g H}_2\text{O}
\]

The above is how much water reacts, we now determine how much remains by subtracting the amount that reacts from the initial amount.

\[9.00 - 8.56 = 0.44 \text{ g H}_2\text{O remains}\]

Let’s summarize limiting reactants with another example.

**Example 3**

25.0 g of Na\(_2\)SO\(_4\) is added to 7.00 g of carbon and allowed to react according to the following equation:

\[\text{Na}_2\text{SO}_4(s) + 4\text{C}(s) \rightarrow \text{Na}_2\text{S}(s) + 4\text{CO}(g)\]

a) What is the limiting reactant?

b) How many grams of Na\(_2\)S can be formed?

c) How many grams of the excess reactant will be leftover?

**Solution:**

Perform the three-step calculation twice, starting from the information given for each reactant.

\[
\frac{25.0 \text{ g Na}_2\text{SO}_4}{142.05 \text{ g Na}_2\text{SO}_4} \times \frac{1 \text{ mol Na}_2\text{S}}{1 \text{ mol Na}_2\text{SO}_4} \times \frac{78.05 \text{ g Na}_2\text{S}}{1 \text{ mol Na}_2\text{S}} = 13.7 \text{ g Na}_2\text{S}
\]

\[
\frac{7.00 \text{ g C}}{120.11 \text{ g C}} \times \frac{1 \text{ mol Na}_2\text{S}}{4 \text{ mol C}} \times \frac{78.05 \text{ g Na}_2\text{S}}{1 \text{ mol Na}_2\text{S}} = 11.4 \text{ g Na}_2\text{S}
\]

Comparing the two calculations leads us to conclude that C is the limiting reactant and Na\(_2\)SO\(_4\) is in excess. 11.4 g of Na\(_2\)S can be produced. Next, find the amount of excess reactant that reacts.

\[
\frac{7.00 \text{ g C}}{120.11 \text{ g C}} \times \frac{1 \text{ mol Na}_2\text{SO}_4}{4 \text{ mol C}} \times \frac{142.05 \text{ g Na}_2\text{SO}_4}{1 \text{ mol Na}_2\text{SO}_4} = 20.8 \text{ g Na}_2\text{SO}_4
\]

The amount remaining is given by the difference

\[25.0 \text{ g initially} - 20.8 \text{ g consumed} = 4.2 \text{ g of Na}_2\text{SO}_4 \text{ remain}\]

### D.5 REACTIONS INVOLVING GASES

Examples 1 and 2 dealt with the combustion of benzene. In each, a known amount of benzene was burned, and we calculated the mass of CO\(_2\) or O\(_2\) produced. However, the measurable quantities of gases are pressure, volume, and temperature, not mass. In this section, we use the treatment presented in Appendix B to introduce these quantities into our stoichiometric calculations.

If there is known quantitative information about the pressure, volume, and temperature of a gas, we can calculate the moles of that gas, using \(PV = nRT\). We can use this relationship in the first or last step of our calculation. We summarize this through the following road map or flowchart:

---

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Example 4

Octane combusts via the following chemical reaction:

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$$

How many liters of CO₂, collected at a pressure of 1.00 atm and a temperature of 25 °C, can be produced by the combustion of 35.0 kg of octane (the amount held by a typical car gasoline tank)?

Solution:

We first must recognize this as a reaction stoichiometry problem. We know the mass and therefore the number of moles of octane. We desire the volume of the carbon dioxide, so PV = nRT will be our third step. It is somewhat complicated to use the ideal gas law as a conversion factor, so we will string together the first two steps, and then do the third step separately, solving for volume.

$$35 \text{ kg } C_8H_{18} \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mol } C_8H_{18}}{114 \text{ g } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}} = 2.45 \times 10^3 \text{ mol } CO_2$$

Next apply the ideal gas law, PV = nRT.

$$V = \frac{nRT}{P} = \frac{(2.45 \times 10^3 \text{ mol})(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{1.00 \text{ atm}} = 5.99 \times 10^4 \text{ L}$$

Comment:

Burning a tank of gasoline generates enough carbon dioxide at 1.00 atm and 25 °C to fill a 16 ft × 16 ft room with an 8 ft ceiling.

Example 5

The following reaction is used to quickly inflate some car airbags:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(s)$$

How many grams of NaN₃ must be used if you wish to fill a 20.0 L airbag with dinitrogen to a pressure of 1.25 atm at 25 °C?

Solution:

In this problem the known information involves a gaseous product (a pressure, volume and temperature are all given) and information on the reactant is desired. This time, the first of the three steps involves the ideal gas law, with the second and third steps being simple application of conversion factors.

$$35 \text{ kg } C_8H_{18} \times \frac{10^3 \text{ g}}{\text{kg}} \times \frac{1 \text{ mol } C_8H_{18}}{114 \text{ g } C_8H_{18}} \times \frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}} = 2.45 \times 10^3 \text{ mol } CO_2$$

Next apply the ideal gas law, PV = nRT.

$$V = \frac{nRT}{P} = \frac{(2.45 \times 10^3 \text{ mol})(0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{1.00 \text{ atm}} = 5.99 \times 10^4 \text{ L}$$

Comment:

Burning a tank of gasoline generates enough carbon dioxide at 1.00 atm and 25 °C to fill a 16 ft × 16 ft room with an 8 ft ceiling.

Example 6

$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

If 25.0 mL of 0.375 M Pb(NO₃)₂ is reacted with excess KI, how many grams of PbI₂ will be produced?

Solution:

The moles of Pb(NO₃)₂ can be found since the volume and molarity of the solution are known. We convert the volume to liters then apply...
molarity to determine moles. The final two steps of the calculation are
the mole to mole conversion from the balanced equation, and the
conversion from moles of PbI₂ to grams, using the molar mass of
461.0 g/mol.

\[
25.0 \text{ mL} \times \frac{0.375 \text{ mol Pb(NO₃)₂}}{1000 \text{ mL}} \times \frac{1 \text{ mol PbI₂}}{1 \text{ mol Pb(NO₃)₂}} \times \frac{461.0 \text{ g PbI₂}}{1 \text{ mol PbI₂}} = 4.32 \text{ g PbI₂}
\]

Example 7

The silver ions in AgNO₃ solution can be precipitated by the
addition of aqueous NaCl:

\[
\text{AgNO₃(aq)} + \text{NaCl(aq)} \rightarrow \text{AgCl(s)} + \text{NaNO₃(aq)}
\]

When 35.0 mL of a AgNO₃ solution of unknown concentration is
reacted with excess NaCl solution, 8.53 g of AgCl is formed.
What is the concentration of the AgNO₃ solution?

Solution:

In this problem, there is some information about the AgNO₃ solution
and some information about the solid AgCl. After rereading the
problem, it should become clear that the desired quantity is molarity of
the AgNO₃ solution. Thus, we need to start at the other end, with the
mass of AgCl. You should also recognize that whenever mass data is
presented along with either a molar mass or a chemical formula (from
which we can get a molar mass), we have an entry into our road map.
Here, we will do the first two steps of the calculation in the usual
manner, and then use the \(M = \frac{n}{V}\) relationship as our third step.

\[
8.53 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.35 \text{ g AgCl}} \times \frac{1 \text{ mol AgNO₃}}{1 \text{ mol AgCl}} = 0.0595 \text{ mol AgNO₃}
\]

\[
M = \frac{n}{V} = \frac{0.0595 \text{ mol}}{0.0350 \text{ L}} = 1.70 \text{ mol/L} = 1.70 \text{ M}
\]

Comment:

The first two steps are as we have done many times now. The third
step, using the molarity relationship, may at first seem a little unusual.
Molarity is always the ratio of moles of a substance to the volume in
liters. The first two steps tell us that there are 0.0595 moles of AgNO₃
contained in the 35.0 mL of solution. We simply take the ratio of these
two numbers, first converting 35.0 mL to 0.0350 L, since molarity is
moles per liter. Note that the experiment did not have to be done on a
1 L scale in order to calculate the molarity!

Example 8

Potassium permanganate solutions can react with acidic
hydrogen peroxide solutions via the following balanced equation in water:

\[
2\text{KMnO₄} + 5\text{H₂O₂} + 6\text{HCl} \rightarrow 2\text{MnCl₂} + 5\text{O₂(g)} + 2\text{KCl} + 8\text{H₂O(l)}
\]

When 15.0 mL of 0.0200 M KMnO₄ reacts with excess H₂O₂ and
HCl, how many liters of O₂, collected at a total pressure of 1.00
atm and a temperature of 27 °C, will be formed?

Solution:

The volume and molarity data on KMnO₄ allow us to enter into the
road map. The desired information is volume of O₂. We start by
converting mL to L, and then proceed through the first two of the three
steps in the calculation to find moles of O₂. We will do the third step,
manipulation of \(PV = nRT\), separately.

\[
15.0 \text{ mL} \times \frac{0.0200 \text{ mol KMnO₄}}{1000 \text{ mL}} \times \frac{5 \text{ mol O₂}}{2 \text{ mol KMnO₄}} = 7.50 \times 10^{-4} \text{ mol O₂}
\]

\[
V = \frac{nRT}{P} = \frac{(7.50 \times 10^{-4} \text{ mol}) (0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}) (300 \text{ K})}{1.00 \text{ atm}} = 0.0185 \text{ L}
\]

Comment:

As you can see, there are many possible routes along our road map
for solving stoichiometry problems.

Example 9

How many mL of a 0.250 M NaOH solution are required to
completely react with 40.0 mL of a 0.150 M H₂SO₄ solution. The
overall reaction is:

\[
\text{H₂SO₄(aq)} + 2\text{NaOH(aq)} \rightarrow 2\text{H₂O(l)} + \text{Na₂SO₄(aq)}
\]

Solution:

We see that there is enough information to calculate moles of H₂SO₄.
We desire information on NaOH. Start with the volume of 0.0400 L of
H₂SO₄ solution and proceed as usual.

\[
0.0400 \text{ L} \times \frac{0.150 \text{ mol H₂SO₄}}{1 \text{ L}} \times \frac{2 \text{ mol NaOH}}{1 \text{ mol H₂SO₄}} \times \frac{1 \text{ L}}{0.250 \text{ mol NaOH}} \times \frac{10³ \text{ L}}{1 \text{ mL}} = 48.0 \text{ mL}
\]

Comment:

The mL to L conversion was done at the beginning and the end of the
problem since the volume information is in mL, but the concentration (molarity) is, of course, in moles per liter. Both molarities were used as conversion factors; 0.150 M converted volume of solution to moles of H₂SO₄, 0.250 M was “turned upside down” to convert moles of NaOH to liters of solution.

Example 10

Aqueous HCl and NaOH react in the following manner:

HCl(aq) + NaOH(aq) → H₂O(l) + NaCl(aq)

When 30.0 mL of 0.100 M HCl are mixed with 20.0 mL of 0.125 M NaOH, what is concentration of the excess reagent?

Solution:

This is a limiting reactant problem. At first glance, it would appear quite different than the limiting reactant problems we saw earlier. However, upon careful reading of the experiment, we see that we have quantitative information on both reactants, enough to calculate moles of both. The desired quantity the concentration of the excess reactant. In essence, we are reacting an acid and a base, and need to determine which reactant is limiting, and find how much of the excess reactant is leftover, as was done in Example 3 above. In order to determine the limiting reactant, we calculate how much product can be made from each reactant. It doesn’t matter which product we choose. Let’s pick water.

$$30 \text{ mL} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{0.100 \text{ mol HCl}}{1 \text{ L}} \times \frac{1 \text{ mol H₂O}}{1 \text{ mol HCl}} = 0.0030 \text{ mol H₂O}$$

$$20 \text{ mL} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{0.125 \text{ mol NaOH}}{1 \text{ L}} \times \frac{1 \text{ mol H₂O}}{1 \text{ mol NaOH}} = 0.0025 \text{ mol H₂O}$$

Fewer moles of water can be made from the NaOH, so NaOH is the limiting reactant, HCl is the excess reactant. Notice that it was not necessary to go all the way through and calculate the grams of water. Clearly, if we multiply each result by 18.02 g/mol (the molar mass of H₂O), the conclusion is the same, NaOH is limiting. In order to calculate molarity of HCl, we need the number of moles of HCl which were leftover, and the total solution volume.

Moles of HCl at the start:

$$30 \text{ mL} \times \frac{1 \text{ L}}{10^3 \text{ mL}} \times \frac{0.100 \text{ mol HCl}}{1 \text{ L}} = 0.0030 \text{ mol HCl}$$

Moles of HCl remaining:

$$0.0030 - 0.0025 = 0.0005 \text{ mol HCl remain}$$

Concentration of HCl at the end

$$\frac{0.0005 \text{ mol HCl}}{0.050 \text{ L}} = 0.010 \text{ M}$$

Comment:

This problem is actually very similar to the limiting reactant problems we did before. The difference is that, instead of finding the grams of the leftover reactant, we had to find the concentration, which involved a calculation of the number of moles of the leftover reactant.

D.7 EXERCISES

Use the following molar masses to do the following problems:

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Molar Mass (g/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₄H₈</td>
<td>56.10</td>
</tr>
<tr>
<td>C₄H₉OH</td>
<td>74.12</td>
</tr>
<tr>
<td>Fe₂O₃</td>
<td>159.70</td>
</tr>
<tr>
<td>Al₂O₃</td>
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<tr>
<td>V₂O₅</td>
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<tr>
<td>NH₄VO₃</td>
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</tr>
<tr>
<td>CuS</td>
<td>80.55</td>
</tr>
<tr>
<td>Cu₂O</td>
<td>143.4</td>
</tr>
<tr>
<td>Cu₂S</td>
<td>159.17</td>
</tr>
<tr>
<td>CuO</td>
<td>79.55</td>
</tr>
<tr>
<td>AgCl</td>
<td>143.4</td>
</tr>
</tbody>
</table>

1. In the presence of acids, water can react with alkenes to form alcohols:

$$C₄H₈ + H₂O → C₄H₉OH$$

If 250 g of C₄H₈ reacts with excess H₂O, how many grams of C₄H₉OH can be produced?

2. Aluminum reacts with iron(III) oxide in the “thermite reaction”:

$$2Al(s) + Fe₂O₃(s) → 2Fe(s) + Al₂O₃(s)$$

a) If 10.0 g of Al reacts with excess Fe₂O₃, how many grams of Al₂O₃ can be produced?

b) If 25.0 g of Al reacts with 10.0 g of Fe₂O₃, how many grams of Al₂O₃ can be produced?

c) In the experiment in part b, what is the mass of the excess reactant remaining after complete reaction?
3. Vanadium(V) oxide reacts with ammonia and water as follows:
\[ \text{V}_2\text{O}_5 + 2\text{NH}_3 + \text{H}_2\text{O} \rightarrow 2\text{NH}_4\text{VO}_3 \]
   a) If 50.0 g of V\(_2\)O\(_5\) is reacted with excess ammonia and water, how many grams of NH\(_4\)VO\(_3\) can be produced?
   b) How many grams of NH\(_3\) are required to completely react with 50.0 g of V\(_2\)O\(_5\)?

4. Vanadium(III) oxide can be made by reduction of vanadium(V) oxide with hydrogen:
\[ \text{V}_2\text{O}_5(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{V}_2\text{O}_3(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \]
   a) How many liters of H\(_2\), measured at 1.00 atm and 30 °C, are required to completely react with 75.0 g of V\(_2\)O\(_5\)?
   b) If 10.0 g of V\(_2\)O\(_5\) reacts with 1.65 L of H\(_2\), measured at 1.00 atm and 30 °C, how many grams of V\(_2\)O\(_3\) can be produced?

5. Copper(I) sulfide is prepared by heating copper and sulfur in the absence of air:
\[ 2\text{Cu}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{Cu}_2\text{S}(\text{s}) \]
   a) How many grams of Cu\(_2\)S can be produced from the reaction of 25.0 g of Cu with excess S?
   b) How many grams of sulfur are required to form 75.0 g of Cu\(_2\)S?
   c) If a mixture of 135 g of Cu and 45 g of S is allowed to react, how many grams of Cu\(_2\)S could be produced?
   d) How many grams of the excess reactant remain in the experiment in part c?

6. Copper(I) oxide can be prepared by thermal decomposition of copper(II) oxide:
\[ 4\text{CuO}(\text{s}) \rightarrow 2\text{Cu}_2\text{O}(\text{s}) + \text{O}_2(\text{g}) \]
   a) How many grams of Cu\(_2\)O can be produced upon the decomposition of 450 g of CuO?
   b) How many liters of O\(_2\), collected at 1.00 atm and 27 °C, can be produced by the decomposition of 450 g of CuO?

7. The silver ions in aqueous silver sulfate can be precipitated by addition of excess chloride:
\[ \text{Ag}_2\text{SO}_4(\text{aq}) + 2\text{NaCl}(\text{aq}) \rightarrow 2\text{AgCl}(\text{s}) + \text{Na}_2\text{SO}_4(\text{aq}) \]
   a) How many grams of silver chloride can be formed when 35.0 mL of a 0.100 M Ag\(_2\)SO\(_4\) solution is reacted with excess sodium chloride solution?
   b) If 22.7 mL of a silver sulfate solution of unknown concentration yields 0.985 g of AgCl upon reaction with excess sodium chloride solution, what is the concentration of the silver sulfate solution?

8. Zn metal reacts with hydrochloric acid to produce hydrogen gas and zinc(II) chloride:
\[ \text{Zn}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g}) \]
   a) If 15.0 g of Zn are added to excess HCl(aq), how many liters of H\(_2\)(g), collected at 27 °C and 725 mm Hg, are produced?
   b) If excess Zn is added to 25.0 mL of 0.025 M HCl(aq), how many liters H\(_2\)(g), collected at 27 °C and 725 mm Hg, can be produced?

9. Potassium permanganate and iron(II) chloride undergo an electron transfer reaction in acid solution:
\[ \text{KMnO}_4(\text{aq}) + 5\text{FeCl}_2(\text{aq}) + 8\text{HCl}(\text{aq}) \rightarrow \text{MnCl}_2(\text{aq}) + 5\text{FeCl}_3(\text{aq}) + \text{KCl}(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) \]
   How many mL of 0.150 M FeCl\(_2\)(aq) are needed to completely react with 13.7 mL of 0.110 M KMnO\(_4\)?

10. Citric acid reacts with sodium hydroxide in a proton transfer reaction:
\[ \text{H}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq}) + 3\text{NaOH}(\text{aq}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + \text{Na}_3\text{C}_6\text{H}_5\text{O}_7(\text{aq}) \]
   a) How many mL of 0.125 M NaOH(aq) are required to completely react with 25.0 mL of 0.0695 M citric acid?
   b) If 37.5 mL of 1.25 M NaOH(aq) is needed to completely react with 22.5 mL of a citric acid solution, what is the concentration of the citric acid solution?
ANSWERS:

1. 330 g
2. a) 18.9 g  b) 6.38 g  c) 21.6 g
3. a) 64.3 g  b) 9.34 g
4. a) 20.5 L  b) 4.97 g
5. a) 31.3 g  b) 15.1 g  c) 169 g  d) 11 g
6. a) 405 g  b) 34.8 L
7. a) 1.00 g  b) 0.151 M
8. a) 5.92 L  b) 8.07 mL
9. 50.2 mL
10. a) 41.7 mL  b) 0.694 M
Appendix E
Fundamental Constants

ENERGY

1 joule (J) = 1 kg \cdot m^2 \cdot s^{-2}

1 calorie (cal) = 4.184 J

1 V = 96.485 kJ/mol

FORCE

1 newton (N) = 1 kg \cdot m/s^2

LENGTH

1 meter (m) = 39.37 inches (in)

1 inch = 2.54 centimeters (cm) - exact

1Å = 1x10^{-10} m

MASS

1 kilogram (kg) = 2.205 pounds (lb)

1 lb = 453.6 grams (g)

1 amu = 1.661x10^{-24} g

PRESSURE

1 atm = 760 mm Hg (torr)

= 1.01325x10^5 Pa

VOLUME

1 liter (L) = 1000 mL = 1000 cm^3

PHYSICAL CONSTANTS

<table>
<thead>
<tr>
<th>Physical Constant</th>
<th>Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>Avogadro’s number NA</td>
<td>(6.0221 \times 10^{23}) mol⁻¹</td>
</tr>
<tr>
<td>Electronic Charge e</td>
<td>(1.6022 \times 10^{-19}) C</td>
</tr>
<tr>
<td>Electron rest mass me</td>
<td>(9.1094 \times 10^{-31}) kg</td>
</tr>
<tr>
<td>Faraday constant (F)</td>
<td>(9.6485 \times 10^{4}) C \cdot mol⁻¹</td>
</tr>
<tr>
<td>Gas constant R</td>
<td>(0.08206) L \cdot atm \cdot K⁻¹ \cdot mol⁻¹</td>
</tr>
<tr>
<td>Neutron rest mass m_n</td>
<td>(1.675 \times 10^{-27}) kg</td>
</tr>
<tr>
<td>Planck’s constant h</td>
<td>(6.6261 \times 10^{-34}) J \cdot s</td>
</tr>
<tr>
<td>Proton rest mass m_p</td>
<td>(1.6726 \times 10^{-27}) kg</td>
</tr>
<tr>
<td>Speed of light (vacuum) c</td>
<td>(2.9979 \times 10^{8}) m \cdot s⁻¹</td>
</tr>
</tbody>
</table>

TEMPERATURE

\[ 0 \text{ K} = -273.15 \degree \text{Celsius} = -459.67 \degree \text{Fahrenheit} \]

\[ \degree \text{F} = \left(\frac{9}{5}\right) \degree \text{C} + 32 \]

\[ \degree \text{C} = \left(\frac{5}{9}\right)(\degree \text{F} - 32) \]

\[ K = \degree \text{C} + 273.15 \]

SI PREFIXES

\[ 10^9 \text{ giga (G)} = 10^6 \text{ mega (M)} = 10^3 \text{ kilo (k)} = 10^1 \text{ deci (d)} = 10^2 \text{ centi (c)} = 10^3 \text{ milli (m)} = 10^6 \text{ micro (µ)} = 10^9 \text{ nano (n)} = 10^{-12} \text{ pico (p)} \]
Absolute (or Kelvin) temperature scale is used for the temperature in all calculations involving T. The unit is the kelvin (K). The average kinetic energy of the molecules in a system is directly proportional to its temperature in kelvins.

Absolute zero is 0 K, which is -273.16 °C. It is the temperature at which molecules have no kinetic energy.

Absorbance is a measure of the amount of light absorbed by a substance. The absorbance of a solution depends upon both the concentration and the molar absorptivity of the absorbing substance at the wavelength of the light, and the distance through the solution that the light travels. See Beer’s Law.

Absorption of a photon increases the energy of an atom or a molecule by the energy of the photon (hv). A photon can be absorbed only if its energy matches the energy difference between two energy levels in the atom or molecule.

An absorption spectrum presents the absorbance of a substance as a function of the wavelength or frequency of light.

The acceptor orbital is the orbital on an oxidizing agent that accepts the transferred electrons in a redox reaction.

The acid dissociation or ionization constant is the equilibrium constant for the reaction of an acid with water: HA + H₂O → H₃O⁺ + A⁻.

An acidic salt is a salt in which the acidity of the cation is greater than the basicity of the anion.

An acidic solution is one with [H₃O⁺] > [OH⁻]. As a result, pH < 7.0 at 25 °C for acidic solutions.

The activation energy is the energy of the transition state relative to that of the reactants or products. It is the minimum energy that the reactants must have in order for a reaction to occur.

An active electrode is one that is a participant in a reaction. For example, a copper electrode in a Cu²⁺ + 2e⁻ → Cu half-cell is active because copper metal participates in the reaction.

The activity is the ratio of the concentration of a substance to its concentration in the standard state. It is unitless. The activities of pure solids and liquids are unity. The activity of a gas equals the partial pressure of the gas in atmospheres divided by 1 atm, while the activity of a solute equals its molar concentration divided by 1 M.

Addition polymers are formed by addition reactions.

An addition reaction is a reaction in which two reactants combine to form a single product.

Adhesive forces are forces between different molecules (compare with cohesive force).

An alcohol is a compound with the general formula R-OH, where R is a generic group of atoms and OH is the hydroxyl group.

An alkali metal is an element that belongs to Group 1A.

An alkaline earth metal is an element that belongs to Group 2A.

An alkane is a saturated hydrocarbon, i.e., a hydrocarbon that contains no multiple bonds.

An alkene is a hydrocarbon that contains carbon-carbon double bonds.

An alkyl group is an organic group formed by removing one hydrogen atom from an alkane.

Allotropes are different crystalline forms of the same element that have different properties. Graphite and diamond are allotropes of carbon.

Alpha decay is the emission of an alpha particle. It is common among the heavy isotopes because it is the best way to reduce mass.

An alpha particle is a helium nucleus.

An amide is an amine attached to a carbonyl.

An amine is an ammonia molecule in which one or more of the hydrogen atoms have been replaced with other groups.

An amino acid is a compound that contains both amine and carboxylic acid functional groups.

Amorphous solids have ordered arrangements of particles over short distances only. This is referred to as local order.

The ampere (A) is the SI unit for electrical current. 1 A = 1C/s.

An amphiprotic substance is able to function as either an acid or a base.

An analyte is a substance that is being analyzed.

The angstrom (Å) is 10⁻¹⁰ m. It is commonly used for bond lengths because most bond lengths are between 1 and 2 Å.
Angular momentum ($\mathbf{L}$) is a property of a rotating object. It is equal to the mass of the object times its velocity times its distance from the center of rotation; i.e., $L = mv\mathbf{r}$.

The angular momentum quantum number ($\ell$) is an integer between 0 and $n-1$ that defines the shape of an atomic orbital.

An anion is a negatively charged ion.

The anode compartment or electrode is where oxidation occurs in an electrochemical cell.

Antibonding interactions occur in molecular orbitals when the atomic orbitals on adjacent atoms used to construct the molecular orbital have opposite phases.

An antibonding MO is one in which the number of antibonding interactions exceeds the number of bonding interactions.

An antiferromagnetic substance is not magnetic because all of its electron spins are paired.

An Arrhenius acid is a substance that contains $\text{H}^+$ ions in water.

An Arrhenius base is a substance that contains $\text{OH}^-$ and produces $\text{H}^+$ ions in water.

The Arrhenius equation relates a rate constant to the temperature and activation energy of the reaction: $k = Ae^{-E_a/RT}$ or $\ln k = \ln A - E_a/RT$.

An Arrhenius plot is a plot of $\ln k$ (rate constant) versus $1/T$. The slope is $-E_a/R$ and the intercept is $\ln A$ (the pre-exponential).

Atoms are the building blocks of matter. Elements consist of only one type of atom.

Atomic mass or atomic weight is the average mass of the atoms of an element relative to that of carbon-12, which is assigned a relative mass of exactly 12.

The atomic mass unit (amu) is a unit of mass that is $\frac{1}{12}$ the mass of a single atom of carbon-12.

The atomic number ($Z$) is the number of protons in the nucleus. It identifies the atom.

The atomization energy ($\Delta H_{\text{atom}}$) is the energy required to break all of the bonds in a molecule in the gas phase to produce the atoms.

Autoionization of water is the reaction of water with itself: $2\text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{OH}^-$. Avogadro’s law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

Avogadro’s number is $6.02 \times 10^{23}$. It is the number of items in a mole.

The band gap is the energy separation between the valence and conduction bands of a metallic or covalent solid.

The band or belt of stability is the region of a plot of the number of neutrons versus the number of protons in a nucleus in which the stable nuclei fall.

Band theory is an extension of mo theory to metals. A very large number of atomic orbitals in a metal combine to form a very large number of molecular orbitals. The resulting molecular orbitals are so close in energy that they form an energy band.

A barometer is a device used to determine atmospheric (or barometric) pressure.

A base pair consists of two complementary, N-containing bases whose structures maximize H-bonding between them. Guanine and cytosine are base pairs as are adenine and thymine. Base pairs hold the two strands of DNA together.

A basic salt is one in which the basicity of the anion exceeds the acidity of the cation.

Basic solutions are solutions in which $[\text{H}^+] < [\text{OH}^-]$. A basic solution has a pH $> 7.0$ at $25^\circ C$.

A battery is a galvanic cell or a collection of galvanic cells. Batteries harness the free energy changes in redox reactions.

Beer’s Law: The absorbance ($A$) of a solution equals the product of its molar absorptivity ($\varepsilon$), its molar concentration ($c$), and the path length ($l$) of the cell in which its absorbance is measured; $A = \varepsilon cl$.

Belt of stability See band of stability.

Beta decay is the ejection from the nucleus of an electron produced by the decay of a neutron. Beta decay reduces the neutron/proton ratio, so it is common among nuclei that lie above the band of stability.

A beta particle ($\beta$) is a high energy electron.

A bimolecular process is one that involves two molecules.

Binary compounds are composed of only two elements; $\text{Al}_2\text{O}_3$ is a binary compound because it contains only Al and O.

The binding energy is the energy that holds the nucleus together. It is related to the mass defect by $\Delta E = \Delta mc^2$.

Blackbody radiation is the light emitted by a solid when it is heated.

A body-centered cubic (bcc) unit cell is one in which the particles that lie on the corners are also in the body center.

The boiling point is the temperature at which the vapor pressure equals the external pressure. If the external pressure is 1 atm, then the temperature is called the normal boiling point.
Boiling point elevation (ΔT_b) is the increase in the boiling point caused by the addition of a non-volatile solute to a solvent.

The boiling point elevation constant (k_b) is the proportionality constant that relates the boiling point elevation of a solution to its colligative molality. ΔT_b = k_b m_c.

The bond angle is the angle formed by two bonds to an atom.

The bond dipole is a measure of bond polarity. It is represented by an arrow pointing from the less electronegative atom toward the more electronegative atom.

The bond energy or bond dissociation energy is the amount of energy required to break one mole of bonds in the gas phase.

The bond length is the distance between two bound nuclei.

The bond order is the number of shared pairs in a bond. As the bond order increases, the length of the bond decreases and its strength increases. The bond order in a diatomic molecule is also equal to 1/2 the difference between the number of its bonding and antibonding electrons.

The bonding electrons are the shared electrons in a covalent bond.

A bonding interaction occurs in a molecular orbital when the phases of the atomic orbitals of two adjacent atoms are the same.

A bonding MO is a molecular orbital in which the number of bonding interactions exceeds the number of antibonding interactions.

Boyle’s law states that the pressure-volume product of a fixed amount of gas at constant temperature is constant. PV = k(n,T).

A branched chain hydrocarbon contains a chain of carbons atoms in which at least one carbon is bound to three or four other carbon atoms.

A Brønsted acid is a proton donor.

A Brønsted base is a proton acceptor.

A buffer is a solution that contains a weak acid and its conjugate base in appreciable and comparable amounts. Buffers reduce pH changes brought about by the addition of strong acids and bases.

The buffer capacity is the amount of strong acid or base on which a buffer can act.

The buffer range is the pH range over which a buffer can function.

A bulk property is a property of a material (such as a pure solid or liquid) as opposed to individual atoms or molecules. Bulk properties are different than the atomic or molecular properties of its components due to the interactions between the components.

A cathode compartment or electrode is where reduction occurs in an electrochemical cell.

A cathode ray is light emitted from the cathode (negative electrode) of a gas discharge tube.

A cation is a positively charged ion.

The cell potential is the potential difference between the cathode and anode of an electrochemical cell. \( E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}} \)

The Celsius (or centigrade) scale is the temperature scale based on the freezing (0 °C) and boiling points (100 °C) of water.

A chain reaction is a reaction in which a product initiates more reaction.

Charles’ law states that the volume of a fixed amount of gas at constant pressure is proportional to its absolute temperature. \( V = k(n,P)T \)

A chemical property is a property of a substance that requires the substance to change into another substance. Hydrogen and oxygen react to produce water is a chemical property of hydrogen.

Chemistry is that branch of science that deals with matter and the changes it undergoes.

A cis configuration is one in which two groups are on the same side of a bond or atom.

Cohesive forces are forces between like molecules (compare with adhesive force).

The colligative concentration is the concentration of all solute particles in a solution. The colligative concentration of a solute equals its concentration times its van’t Hoff factor.

Colligative properties are those properties of a solution that depend upon the concentration, but not the identity, of the solute particles.
The **collision frequency** is the number of collisions per unit volume per unit time, which normally has units of (moles of collisions)/(liter-s).

**Combustion** is a reaction with oxygen.

A **common ion** is an ion that appears in an equilibrium but has at least two sources.

A **complex ion** is an ion in which a central metal is surrounded by molecular or anionic ligands.

A **compound** is a pure substance that consists of more than one element.

The **concentration** of a solute is the amount of solute divided by the volume in which it is contained.

A **concentration cell** is an electrochemical cell in which the two compartments differ only in their concentrations. The cell potential depends upon the concentration difference.

**Condensation** is the process of converting a vapor into its liquid.

**Condensation polymers** are formed by condensation reactions.

A **condensation reaction** is a reaction in which two reactants combine to form two products (one of which is often a small molecule such as water or an alcohol).

The **conduction band** is the lowest energy unfilled band in a solid that has no partially filled bands. Electrons in a conduction band are free to move throughout the metal due to the presence of unfilled orbitals. Thus, electrons can conduct electricity only when they are in the conduction band.

A **conductor** is a substance that conducts electricity at all temperatures. Its conduction decreases slightly with increasing temperature.

A **conjugate acid-base pair** is a Bronsted acid and base that differ by one proton only.

**Connectivity** is the manner in which the atoms in a molecule are connected.

**Constitutional isomers** are compounds with the same formula but different connectivities.

A **continuous chain hydrocarbon** is a chain of carbon atoms in which no carbon is bound to more than two other carbon atoms.

A **continuous spectrum** is a spectrum in which all wavelengths of light in the region are present. Thus, they merge into one another continuously. A rainbow is a continuous spectrum of visible light.

A **coordinate covalent bond** is a bond in which both bonding electrons are contributed by the same atom. The bonds formed in Lewis acid-base reactions are coordinate covalent because both bonding electrons always come from the base.

The **coordination number** of a particle is the number of its nearest neighbors in a crystal or in a compound.

**Core electrons** are the tightly bound electrons that are unaffected by chemical reactions. They reside in filled sublevels and form a spherical shell of negative charge around the nucleus that affects the amount of nuclear charge that the outermost electrons experience.

**Corrosion** is the natural oxidation of a metal.

The **coulomb** (C) is the SI unit of electrical charge. The charge on one electron is 1.602x10\(^{-19}\) C.

**Coulomb’s law** states that two charged particles experience a force that is proportional to the product of their charges and varies inversely with the dielectric of the medium and the square of the distance that separates them. Negative forces are attractive, while positive forces are repulsive.

A **counter ion** is an ion that accompanies a desired ion in order to maintain the electrical neutrality of the compound that contains the desired ion. Counter ions are spectator ions in net chemical equations.

A **covalent bond** results when electrons are shared. It can be viewed as the attraction of the bonding electrons for the bound nuclei.

The **covalent radius** of an atom X is equal to one-half of the distance between the X atoms in X\(_2\).

The **critical mass** is the minimum mass of a radioactive material required to maintain a chain reaction.

The **critical point** is the temperature and pressure beyond which the liquid cannot exist. Substances beyond their critical point are supercritical fluids.

The **critical pressure** is the pressure required to liquefy a gas at its critical temperature.

The **critical temperature** is the highest temperature at which a gas can be liquefied.

A **crystal orbital** is to a crystal what a molecular orbital is to a molecule.

**Crystalline solids** are solids with well defined and ordered repeat units. The order, which exists throughout the crystal, is said to be long range order.

**Degrees of freedom** are the basic set of motions (translations, rotations, and vibrations) that a molecule undergoes. The kinetic energy of a molecule is distributed amongst its degrees of freedom. A molecule with N atoms has 3N degrees of freedom.

**Delocalized** electrons or bonds are spread over several atoms.
Density is the mass to volume ratio of a substance or solution. \( d = \frac{m}{V} \)

The density of states is the number of allowed energy states in a region of energy.

Deposition is the process in which a vapor is converted into its solid.

A detergent is a substance that has both a hydrophobic region that interacts well with nonpolar molecules such as grease, and a hydrophilic region that interacts well with polar molecules such as water.

Diamagnetism is the tendency of certain atoms not to be attracted by a magnetic field. It is an atomic property associated with atoms that have no unpaired electrons.

Diatomic molecules contain two and only two atoms.

The dielectric constant \( (\varepsilon) \) is a number that relates the ability of a medium to shield two charged particles from one another. A medium with a high dielectric constant shields the charges better than one with a low constant.

A dipole consists of two electrical poles, one positive and one negative. Bonds dipoles arise between atoms of different electronegativities. A molecular dipole is the vector sum of its bond dipoles.

Dipole-dipole or dipolar forces are the inter-molecular forces that result from the interaction of the oppositely charged poles of two polar molecules.

Dispersion forces are forces between molecules that result from the interaction of temporary or induced dipoles. Dispersion forces increase approximately with molecular size.

The dissociation constant is the equilibrium constant for the dissociation of a complex ion into its component ions and/or molecules. Also see acid dissociation constant.

The dissociation or bond energy is the energy required to break one mole of bonds in the gas phase.

Dissolution is the process in which an ionic substance dissolves in water to produce ions.

A donor orbital is the orbital on the reducing agent that contains the electrons to be transferred in a redox reaction.

The double helix is the structure adopted by DNA. It consists of a pair of intertwined polynucleotide strands held together by hydrogen bonding between base pairs.

Dynamic equilibria are attained when two competing processes occur at equal rates. Contrast to a static equilibrium where the competing processes stop.

Effective nuclear charge \( (Z_{\text{eff}}) \) is the nuclear charge experienced by an electron in an atom. It is less than the nuclear charge due to shielding by the other electrons.

An electrical current is the rate at which charge flows through a circuit. A current of one ampere is a rate of one Coulomb of charge per second.

An electrochemical cell is a device used to extract the free energy change of a spontaneous redox reaction (see Galvanic cells) or to inject the energy required to drive a redox reaction that is not spontaneous (see electrolytic cells).

Electrochemistry is the combination of electrical conduction through a circuit and electron transfer reactions.

An electrode is a metal that provides a surface at which electrons can be transferred between an electrical circuit and a reactant in a redox reaction.

Electrodes are active if they participate in the reaction and passive if they do not.

Electrolysis is a non-spontaneous redox reaction that is driven uphill in free energy by the application of an external electrical potential.

An electrolyte is a material that produces ions when dissolved in water. Electrolytes can be weak or strong depending upon the extent to which they produce ions. Substances that dissolve in water as molecules rather than ions are called non-electrolytes.

An electrolytic cell is an electrochemical cell that converts electrical potential energy into chemical potential energy. See electrolysis.

Electrolytic conduction is conduction of electricity through a solution as a result of the migration of ions in the solution.

Electromagnetic radiation is an electric and a magnetic field oscillating perpendicular to one another that travels through space in the form of radio waves, microwaves, infrared waves, visible light, ultraviolet light, etc.

An electron is the basic quantity of negative charge. It carries a charge of \(-1.602\times10^{-19}\) C and has a mass of \(5\times10^{-4}\) amu.

Electron capture is the capture of a core electron by the nucleus. It converts a proton into a neutron.

The electron configuration of an atom is a listing of the sublevels that are occupied and the number of electrons in them.

Electron density is the probability of finding an electron in a particular region of space. The electron density is high in regions where the probability of finding an electron is high.
Electronegativity \((\chi)\) is a relative measure of the ability of an atom to attract bonding electrons to itself. Atoms with high electronegativities have unfilled orbitals that are low in energy.

An **electronic transition** is the changing of the energy of an electron from one quantum state to another.

An **element** is a pure substance that cannot be broken down into a simpler substance by chemical means.

The **elemental composition** of a compound is a listing of the relative masses, usually expressed as percents, of the elements in the compound.

An **elementary reaction** is a reaction that occurs in one step.

**Emission** is the ejection of a photon by an atom or a molecule. The energy of the atom or molecule decreases by the energy of the photon \((h\nu)\).

An **empirical or simplest formula** is a chemical formula that indicates only the smallest whole number ratio of the atoms present in the compound.

**Enantiomers** are two molecules that are non-superimposable mirror images of one another.

An **endothermic** process absorbs heat.

The **end point** is the point at which an indicator changes color. The end point should be nearly the same as the equivalence point.

**Energetics** is a combination of thermodynamics and kinetics.

**Energy** is the capacity to do work or to transfer heat.

An **energy band** is a region of allowed energy in a metal in which there is no separation between adjacent energy levels.

The **energy of interaction** is the energy of two interacting particles relative to the energy of the two particles when they are not interacting. Energies of interaction in chemistry result from the electrostatic interactions.

The **enthalpy or heat of combustion** is the heat absorbed when one mole of a substance reacts with oxygen. Heats of combustion are negative because they are exothermic.

The **enthalpy or heat of reaction** is the heat absorbed by a reaction run at constant temperature and pressure. A negative heat of reaction simply means that the heat is given off not absorbed.

**Entropy** is the thermodynamic measure the number of ways in which a system can distribute its energy. It is often related to the disorder in the system.

An **enzyme** is a biological compound (usually a protein) that acts as a catalyst.

The **equilibrium constant** \((K)\) is the value of the reaction quotient \((Q)\) when equilibrium activities are used.

The **equivalence point** is the point in a titration at which stoichiometric amounts of reactants are present.

**Esters** are compounds with the general formula \(\text{RCOOR}'\), *i.e.*, two groups connected by a carboxyl group.

**Esterification** is a condensation reaction between a carboxylic acid and an alcohol to produce an ester and water.

**Evaporation** is the conversion of a liquid to its vapor.

An **excited state** is an allowed state that is not the lowest energy state.

An **exothermic** process gives off heat.

**Exponential decay** is a decrease in concentration that goes as \(e^{-t}\). First order reactions undergo exponential decay: \([A] = [A]_oe^{-kt}\).

An **extensive property** is one that depends upon the amount of material. Mass and volume are extensive properties. Also see intensive property.

An **extensive reaction** is one with a large equilibrium constant. If a reaction is extensive, then the equilibrium concentration of least one of the reactants will be very small.

**F**

A **face centered cubic** (fcc) unit cell is one in which the atoms that are located in the corners are also found in the centers of the faces.

The **factor label method** is a method that uses the labels (units) of the factors to determine the order and manner in which the factors should be used to convert one quantity into another.

**Family** See group.

The **Faraday** \((F)\) is the charge of one mole of electrons. \(1F = 96,485\) C/mol.

A **fatty acid** is a carboxylic acid with a long hydrocarbon chain.

The **Fermi level** the highest occupied energy level in a band.

A **ferrimagnet** is a magnetic material whose particles have opposing but unequal spins.

A **ferromagnet** is a magnetic material whose particles have aligned spins.

**Ferromagnetism** is a bulk magnetism in a material (such as iron) resulting from the alignment of the spins of adjacent atoms in the same direction.

The **first law of thermodynamics** states that energy is neither created nor destroyed in any process.

**Fission** is the process in which a heavy nucleus splits into lighter nuclei.
A **group (or family)** is a vertical column in the periodic table. The elements in a group have similar properties.

**Heat of vaporization** ($\Delta H_{\text{vap}}$) is the amount of heat required to convert one mole of a liquid into its gas.

The **Henderson-Hasselbalch equation** is used to calculate the pH of a buffer solution:

$$pH = pK_a + \log\left(\frac{[b]}{[a]}\right)$$

**Hess’ law of heat summation** states that if a process can be expressed as the sum of several steps, then the enthalpy change of the process is the sum of the enthalpy changes of the steps.

A **heterogeneous catalyst** is in a different phase than the reactants. Typically it is a solid for gas or solution reactions.

A **heterogeneous mixture** is one whose composition varies as in a mixture of water and oil.

A **high spin metal** is one in which the splitting of the d orbitals is small enough that the d electrons remain unpaired in the higher energy set rather than pairing in the lower energy set.

**Homo** is the abbreviation for the highest occupied molecular orbital.

A **homogeneous catalyst** is in the same phase as the reactants.

A **homogeneous mixture** is a mixture whose composition is the same throughout, *i.e.*, one in which the concentration of each component is the same regardless of the volume that is sampled. Homogeneous mixtures are called solutions.

A **homonuclear diatomic molecule** is one in which the two atoms are the same.

**Hund’s rule** states that the number of electrons with identical spin is maximized when filling the orbitals of a sublevel.

A **hybrid orbital** is an orbital constructed by mixing two atomic orbitals on the same atom. They are used to explain bonding in the valence bond model.
Hybridization is the process by which hybrid orbitals are produced from atomic orbitals.

A **hydrate** is a compound with a characteristic number of water molecules associated with it.

Hydration is the process in which a solute particle interacts with the surrounding water molecules.

A **hydrocarbon** is a compound that contains only carbon and hydrogen.

**Hydrocarbon** is the addition of hydrogen to a compound.

The **hydrogen bond** is an especially strong dipolar interaction that occurs in compounds containing a hydrogen atom attached to N, O, or F.

The **hydronium ion** ($\text{H}_3\text{O}^+$) is the conjugate acid of water. Therefore, it is the strongest acid that can be present in aqueous solutions.

A **hydrophilic** molecule interacts well with water.

A **hydrophobic** molecule is excluded from water because it does not interact well with water.

The **hydrophobic effect** is the tendency of water to exclude hydrophobic molecules by establishing an ice-like structure around them.

A **hypothesis** is a proposed explanation of an observation. If a hypothesis proves successful in explaining many other experiments, it becomes a theory, but if it fails to explain a test, it is discarded or modified.

An ideal gas is a hypothetical gas composed of molecules that do not interact with one another.

The **ideal gas law** is the relationship between the pressure (P), volume (V), temperature (T) and number of moles (n) of an ideal gas. $P\text{V} = nRT$.

Ideal gases obey the ideal gas law at all T and P, while real gases deviate at high P and low T.

An **indicator** is a compound that changes color within a small pH range. The pH at which the indicator changes color is called the end point.

An **induced dipole** is a molecular dipole in one molecule caused by the asymmetric charge distribution in a neighboring molecule.

The instantaneous rate of a reaction is the rate at a specified time. It is equal to the slope of the concentration vs. time plot at the specified time.

An **insulator** is a substance that does not conduct electricity at reasonable temperatures because its band gap is too large.

**Integrated rate law** expresses the concentration of a reactant as a function of time.

An **intensive property** is independent of sample size. Color and density are intensive properties.

An **intermediate** in a chemical reaction is a substance that is formed and then consumed in the reaction. Intermediates do not appear in the net chemical equation for the reaction.

**Intermolecular interactions** are between different molecules. Dipolar and dispersion forces are intermolecular interactions.

**Intramolecular interactions** are within a molecule. Chemical bonds are intramolecular interactions.

An ion is a charged chemical species.

The ion product ($Q_\text{Ip}$) is the reaction quotient for the reaction in which a solid dissolves as its ions in solution. $Q_{\text{Ip}} = K_{\text{Ip}}$ at equilibrium.

The ion product constant of water ($K_w$) is the equilibrium constant for the reaction $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^{+} + \text{OH}^-$. $K_w = [\text{H}_3\text{O}^+]\text{[OH}^-]$, which has a value of 1.0x10^{-14} at 25°C.

An ionic bond is an electrostatic (Coulombic) force between oppositely charged ions.

The ionic radius of an ion is determined from the distances between it and adjacent ions in an ionic crystal. The distance between the two adjacent ions equals the sum of their ionic radii.

The ionization energy is the energy required to remove an electron from an atom or molecule.

Ionizing radiation is high energy radiation that can remove electrons from a substance. X-rays are ionizing radiation.

Two substances are isoelectronic if they have the same number of electrons.

Isomers are different molecules with the same formula.

Isotopes are atoms with the same atomic number but different mass numbers, i.e., isotopes have the same number of protons but different numbers of neutrons.

The joule (J) is the SI unit of energy.

$1 \text{ J} = 1 \text{ kg} \cdot \text{m}^2 \cdot \text{s}^{-2}$

Kaolinite clays are composed of silicate and aluminate layers (aluminosilicates). They are the main component of china clay.

The kelvin (K) is the SI unit of temperature. $K = ^\circ C + 273.15$.

**Kinetics** is the study of reaction rates and mechanisms.

**Kinetic energy** ($KE=\frac{1}{2}mv^2$) is energy of motion. Anything in motion has the capacity to do work on another object by simply colliding with it.
Kinetic-molecular theory is the model used to explain the ideal gas law. One of its postulates is that the average kinetic energy of the molecules in a gas is directly proportional to the absolute temperature of the gas. The kinetic region of a reaction is the period of the reaction in which concentrations are changing.

A crystal lattice is the arrangement of the particles in a crystal. Each particle lies on a lattice site. A law is a statement that summarizes many observations.

The law of combining volumes states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules. The law of conservation of energy is stated by the first law of thermodynamics; \( \Delta E_{\text{univ}} = 0 \).

The law of conservation of mass states that the total mass or reactants and products remains constant during a chemical reaction; \( i.e. \), mass is neither created nor destroyed in a chemical reaction. The law of definite or constant proportions states that the elements of a compound are always present in definite proportions by mass.

The law of multiple proportions states that the masses of one element that combine with a fixed mass of another element in different compounds of the same elements are in a ratio of small whole numbers. Le Châtelier’s principle states that a system at equilibrium will respond to a stress in such a way as to minimize the effect of the stress.

A level or shell is an allowed energy designated by the principal quantum number \( n \).

The leveling effect of a solvent requires that no acid in a solvent can be stronger than the conjugate acid of the solvent and no base can be stronger than the conjugate base of the solvent. Thus, hydronium ion is the strongest acid that can exist in water and hydroxide ion is the strongest base.

A Lewis acid is a substance with a low lying, empty orbital that can be used to form a covalent bond to a Lewis base. Lewis acidic sites are characterized by less than four electron regions. A Lewis base is a substance with a lone pair that can be shared with a Lewis acid to form a covalent bond between the acid and the base.

A Lewis acid-base reaction is the conversion of a lone pair on a Lewis base and the empty orbital on a Lewis acid into a covalent bond between the acid and the base. The kinetic region of a reaction is the period of the reaction in which concentrations are changing.

A ligand is a molecule or ion that is attached to a metal. The ligand field splitting energy (\( \Delta \)) is the energy difference between the sets of d-orbitals in an atom. It results from the electrostatic field of the ligands, \( i.e. \), the ligand field. The limiting reactant is that reactant whose amount limits the amount of product that can be obtained in a reaction, \( i.e. \), the reactant that is totally consumed.

A line spectrum is a spectrum in which only certain wavelengths (lines) are present. Atomic spectra are line spectra.

A liquid junction is a device that allows ion migration between the electrodes of an electro-chemical cell to complete the electrical circuit. A load is a device in a galvanic cell that utilizes the free energy given off by the transferred electrons. A lone pair is a pair of nonbonding valence electrons. A low-spin metal is a metal in which the d electrons pair in the lower energy set of orbitals before occupying the higher energy set. The lumo is the lowest unoccupied molecular orbital.

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A Lewis acid-base reaction is the conversion of a lone pair on a Lewis base and the empty orbital on a Lewis acid into a covalent bond between the acid and the base. The kinetic region of a reaction is the period of the reaction in which concentrations are changing.

A ligand is a molecule or ion that is attached to a metal. The ligand field splitting energy (\( \Delta \)) is the energy difference between the sets of d-orbitals in an atom. It results from the electrostatic field of the ligands, \( i.e. \), the ligand field. The limiting reactant is that reactant whose amount limits the amount of product that can be obtained in a reaction, \( i.e. \), the reactant that is totally consumed.

A line spectrum is a spectrum in which only certain wavelengths (lines) are present. Atomic spectra are line spectra.

A liquid junction is a device that allows ion migration between the electrodes of an electro-chemical cell to complete the electrical circuit. A load is a device in a galvanic cell that utilizes the free energy given off by the transferred electrons. A lone pair is a pair of nonbonding valence electrons. A low-spin metal is a metal in which the d electrons pair in the lower energy set of orbitals before occupying the higher energy set. The lumo is the lowest unoccupied molecular orbital.
The **mechanical surroundings** is that portion of the surroundings that exchanges energy with the system in the form of work.

The **melting point** is the temperature at which the solid and liquid states are in equilibrium.

A **meniscus** is the curved shape of the top of a liquid.

A **metal** is a material that is shiny, malleable, and a good conductor of electricity. Elements that are metals lie on the left side of the periodic chart and represent about 80% of the elements.

A **metallic bond** is one delocalized over the entire metal. The large number of atoms involved in a typical metallic bond is so large that the bonding electrons occupy bands of energy.

**Metalloids** have properties intermediate between the metals and nonmetals. The eight metalloids are shiny and brittle. They are not good conductors of heat or electricity (they are semiconductors).

A **micelle** is spherical arrangement of detergent molecules in which the heads form a polar outer shell and the tails form a hydrophobic liquid center.

**Micro** (μ) is the SI prefix for $10^{-6}$, a millionth.

**Milli** (m) is the SI prefix for $10^{-3}$, a thousandth.

Two liquids are **miscible** if they are soluble in one another in all proportions.

The **molality** (m) of a solute is the number of moles of solute present in 1 kg of solvent.

The **molar absorptivity** (ε) is the absorbance of a 1 M solution in a 1 cm cell.

The **molarity** (M) of a solute is the number of moles of solute present in a liter of solution.

The **molar mass** (M_m) is the mass of one mole of substance. It is equal to the atomic or molecular mass (weight) expressed in grams.

The mole (mol) is $6.02 \times 10^{23}$ items. It is the number of molecules or atoms in a sample of a compound or element that has a mass equal to its molecular or atomic mass expressed in grams.

The **mole fraction** (X) of a substance in a mixture is the number of moles of that substance divided by the number of moles of all components of the mixture.

A **molecular dipole** is equal to the product of the charges on the two poles of a polar molecule and the distance between them. It is represented by an arrow pointing from the center of positive charge toward the center of negative charge.

The **molecular formula** of a compound shows the actual numbers of atoms present in the molecule. Contrast with the simplest or empirical formula that shows only the smallest integers that are in the same ratio as in the molecular formula.

**Molecularity** is the number of reacting molecules in an elementary reaction.

The **molecular mass or weight** is the relative mass of a molecule relative to the mass of a carbon-12 atom.

**Molecular orbital theory** is a bonding theory in which bonds are formed from the combination of several atomic orbitals on several atoms.

**Molecular weight** See molecular mass

A **molecule** is an independent particle that consists of two or more chemically bound atoms.

A **monatomic ion** is derived from a single atom.

A **monomer** is a single unit building block that can be bound to other monomers to form larger molecules. Linking two monomers produces a dimer, linking three produces a trimer, and linking many produces a polymer.

A **neutral solution** is the absorbance of a 1 M solution in a 1 cm cell.

**Neutral salt** is a compound in which the acid and base strengths of the cation and anion are equal.

In **neutral solutions**, $[\text{H}_3\text{O}^+]=[\text{OH}^-]$. The pH of a neutral solution is 7.0 at 25°C.

In **neutralization reactions**, an acid reacts with a base to produce water and a salt.

A **neutron** is a subatomic particle found in the nucleus. It has no charge and a mass of ~1 amu.

A **noble gas** is an element that belongs to Group 8A. The noble gases are helium, neon, argon, krypton, xenon, and radon.

A **nodal plane** is a plane of zero electron density that lies between regions of opposite algebraic sign in an orbital. A p orbital and a π orbital each contain a single nodal plane.

A **nonbonding MO** has an equal number of bonding and antibonding interactions.

A **nonelectrolyte** is a substance whose aqueous solution does not conduct electricity. Electricity is not conducted because the electrolyte produces no ions in solution.

**Nano** (n) is the SI prefix for $10^{-9}$, a billionth.

**Nanotechnology** is science and engineering of systems on the nanoscale (1–50 nm).

The **Nernst equation** relates a cell’s potential to its standard potential and its reaction quotient.

$$E_\text{cell} = E^\circ - \frac{(RT)}{nF}\ln Q$$

A **net chemical equation** shows only those substances that are changed during the reaction.

In a **network covalent solid**, all of the atoms are bound covalently with no discernable molecules.

A **neutralization reaction** is a chemical reaction in which an acid reacts with a base to form water and a salt.

A **nonmetal** is an element that is a poor conductor of electricity. Elements that are nonmetals are located on the right side of the periodic chart and represent about 20% of the elements. They are shiny and brittle. They are not good conductors of heat or electricity (they are insulators).

**Noble gas** is an element that belongs to Group 8A. The noble gases are helium, neon, argon, krypton, xenon, and radon.

A **nucleus** is the central part of a nucleus.

**Molality** (m) is the number of moles of solute present in 1 kg of solvent.

**Molar absorptivity** (ε) is the absorbance of a 1 M solution in a 1 cm cell.

**Mole fraction** (X) of a substance in a mixture is the number of moles of that substance divided by the number of moles of all components of the mixture.
Non-ionizing radiation, such as visible light, does not have sufficient energy to ionize matter.

Nonmetals are elements on the right side of the periodic table. They can be gases, liquids, or solids and are dull, brittle, and poor conductors of electricity. Nonmetals react with one another to form covalent compounds or with metals to form ionic compounds.

The normal boiling point is the temperature at which the vapor pressure of a liquid is 1 atm.

The nuclear binding energy is the energy required to break one mole of nuclei into their constituent nucleons.

Nuclear chemistry or radiochemistry is the study of reactions that involve changes in the nucleus.

Nuclear fission is the splitting of a heavier nucleus into lighter nuclei.

Nuclear fusion is the combination of two lighter nuclei into a heavier one.

Nucleons are the particles found in the nucleus. Protons and neutrons are nucleons.

A nucleotide is a unit of a nucleic acid that consists of a phosphate, a sugar, and an N-containing base. Nucleic acids are polymers built with nucleotides.

The atomic nucleus contains all of the positive charge, virtually all of the mass, but occupies almost none of the volume of an atom.

Nylon is a polyamide produced in the reaction of a diamine and a diester.

The octet rule states that atoms in molecules strive to obtain an octet (eight) of valence electrons by sharing the bonding electrons with other atoms.

An orbital is a solution to the wave equation. Electrons reside in atomic or molecular orbitals, and bonding results from the interaction of atomic orbitals of different atoms.

An organic compound is one that is based on carbon.

Osmosis is the net movement of solvent molecules from a dilute solution into a more concentrated one through a semipermeable membrane, i.e., one that allows only solvent molecules to pass.

Osmotic pressure is the pressure caused at a semipermeable membrane bounded by solutions of different concentration. It results because solute particles cannot pass through the membrane but solvent molecules can.

Overpotential is the amount by which the applied potential for electrolysis must be increased above that predicted from half-cell potentials to carry out the electrolysis at a reasonable rate. Overpotentials are due to high activation energies.

An oxidant is an oxidizing agent.

Oxidation is the loss of electrons or increase in oxidation state that accompanies electron transfer.

The oxidation state of an atom is the charge it would have if its bonds were assumed to be ionic, i.e., if its bonding electrons were assigned to the more electronegative atom in each bond.

An oxidizing agent is a substance that promotes oxidation in other substances. The oxidizing agent is reduced by the electron transfer.

An oxoacid is a Bronsted acid in which the proton is attached to an oxygen atom.

An oxoanion has a central atom surrounded by oxygen atoms. The central atom is usually in a high oxidation state because it is surrounded by the very electronegative oxygen atoms.

P

Packing efficiency is the fraction of the volume of the unit cell that is occupied by particles.

Paramagnetism is the tendency of certain atoms to be attracted by a magnetic field. It is an atomic property that is related to the number of unpaired electrons on the atom.

Partial ionic character is a measure of the charge separation in a bond, which results from electronegativity differences between the bound atoms. A bond is considered to be ionic if it is has over 50% ionic character.
The **percent yield** is the fraction of the theoretical yield, expressed as a percent, that is actually isolated in a chemical reaction.

A **period** in the periodic table is a horizontal row. The properties of the elements in a period vary gradually across the period.

The **periodic law** states that the elements exhibit a periodicity in the chemical and physical properties when they are arranged in the order of their atomic numbers.

The **periodic table or chart** is an arrangement of the elements into rows (periods) and columns (groups) such that the elements in the same group have similar properties.

**pH** is the negative base 10 logarithm of the hydronium ion concentration in a solution. 

\[ \text{pH} = -\log[H_3O^+] \]

A **phase diagram** shows the state of a substance as a function of its temperature and pressure.

A **photon** is a quantum of energy in the form of electromagnetic radiation.

**Photosynthesis** is the process in which plants use solar energy to covert CO₂ and H₂O into carbohydrates.

A **physical property** is one that is independent of other substances. Melting point, boiling point, color, and hardness are some physical properties.

A **pi (\(\pi\)) bond** is formed from the side-on interaction of two p orbitals. Pi bonds have nodal planes that contain the internuclear axis.

The **pKₐ** of an acid is the negative base 10 logarithm of the acid dissociation constant. 

\[ \text{pK}_a = -\log K_a \]

**Planck’s constant** (\(h\)) is the proportionality constant that relates the frequency of a photon to its energy. 

\[ h = 6.626 \times 10^{-34} \text{ J/s} \]

A **polar covalent bond** is a covalent bond in which the bonding electrons are NOT shared equally. Thus, the bonds are between atoms of different electronegativities.

**Polar molecules** have asymmetric charge distributions. The result is a molecular dipole.

The **polarizability** of an atom or molecule is a measure of the ease with which its electron cloud can be deformed.

A **polyamide** is a condensation polymer that contains many amide linkages. Nylons and peptides are polyamides.

A **polyatomic ion** is an ion, such as CO₃²⁻, in which two or more atoms are covalently bound.

A **polyene** is an organic compound with many double bonds.

A **polymer** is a large molecule consisting of many single unit building blocks called mers.

A **polypeptide** is a polyamide produced from the reaction of many amino acids.

**Polyprotic acid acids** have more than one acidic proton. Examples: H₂SO₄ is a diprotic acid and H₃PO₄ is a triprotic acid.

**Polyunsaturated** organic compounds contain many C-C multiple bonds.

A **positron** is an elementary particle with the mass of an electron and a positive charge. It is the antimatter analog of the electron.

**Positron decay** is the emission of a positron from the nucleus. Positron decay increases the neutron/proton ratio, so it is common in nuclei that lie below the band of stability.

**Potential energy** is energy due to position. In chemistry, potential energy arises from the interaction of charged particles, and the closer they are, the stronger they interact.

A **precipitate** is a solid formed when two solutions are mixed, or the act of forming the solid. Thus, AgCl precipitates and is a precipitate when it does.

The **precision** of a number is given by the number of significant figures to which it is reported. 1.00 m is more precise than 1.0 m.

A **pre-exponential** precedes an exponential. Typically used in the Arrhenius equation: 

\[ k = A \exp\left\{\frac{-E}{RT}\right\}, \]

where A is the pre-exponential.

**Pressure (P)** is force per unit area: \( P = F/A \).

**Pressure-volume or PV work** is done when the volume of a gas changes against an external pressure.

The **principal quantum number** \((n)\) specifies the energy level of an electron in an atom.

A **protein** is a large polypeptide.

A **proton** is a subatomic particle found in the nucleus. It carries a +1 charge and has a mass of ~1 amu.

A **proton acceptor** is called a Brønsted base.

A **proton donor** is called a Brønsted acid.

A **purely covalent bond** is a covalent bond in which the bonds between atoms of the same electronegativity are purely covalent.

A **qualitative observation** is one that does not involve numbers.

A **quantitative observation** is one that does involve numbers.

A **quantity** in the factor label method is an amount and is characterized by a single unit. For example, 3 m is a quantity, but 3 m/s is a factor.

A **quantum** is a packet of energy.
A **quantum number** is a number (usually an integer) that designates an allowed state. All atomic and molecular states (e.g., electronic, vibrational, rotational, and nuclear) are described by quantum numbers.

A **radioactive** nucleus is unstable and disintegrates spontaneously to another nucleus by emitting or capturing particles.

**Radioactive dating** is the determination of the age of a material from the amount of material involved in the radioactive decay of one of its components.

**Radiochemistry** See nuclear chemistry.

**Radioisotopes** are radioactive nuclei.

The **rate of change** of a quantity is the rate at which it changes as a function of the change in another quantity.

A **rate constant** (k) is the proportionality constant between the concentrations of the components (usually reactants) of a reaction and the rate of reaction.

The **rate-determining step** (RDS) is the elementary reaction in a mechanism that is so much slower than the other elementary reactions that it dictates the rate of the overall reaction.

The **rate law** expresses the rate of a reaction as a function of the concentrations of the substances (usually reactants) involved in the reaction.

The **rate of disappearance** is the rate at which a reactant reacts.

The **rate of formation** or **appearance** is the rate at which a product is produced.

The **rate of reaction** is the rate at which a product is produced or a reactant reacts divided by its coefficient in the chemical equation.

A **reactant order** is the exponent of the concentration of a reactant in the rate equation for a reaction.

The **reaction coordinate** is the combination of intermolecular distance, bond length and bond angle changes required to convert reactant molecules into product molecules.

A **reaction mechanism** is a series of elementary processes that leads to the overall reaction.

The **reaction order** is the sum of all of the reactant orders in a reaction.

The **reaction quotient** (Q) is expressed as the activities of the products divided by the activities of the reactants. Each activity is raised to an exponent equal to the coefficient of the substance in the balanced equation. When the activities are equilibrium activities, the reaction quotient is called the equilibrium constant.

A **redox couple** is the oxidized and reduced forms of the species involved in a half-reaction. For example, Cu^{2+}/Cu is a redox couple.

The **redox electrons** are the electrons that are transferred in a redox reaction.

**Redox reactions** involve an electron transfer from a reductant to an oxidant.

A **reducing agent** or **reductant** is a substance that promotes reduction in another material. It is oxidized in the process.

**Reduction** is the gain of electrons, which results in a decrease in oxidation state of the species being reduced.

A **residue** in a protein is one of the amino acids making up the protein.

A **resonance structure** is a Lewis structure of a molecule that differs from another Lewis structure only in the placement of electrons.

**Respiration** is the process whereby animals extract energy from carbohydrates.

**Rotational degrees of freedom** are spinning motions about an axis through the center of mass of the molecule. Linear molecules have two rotational degrees of freedom, and nonlinear molecules have three.

**Salts** are ionic compounds formed in an Arrhenius acid-base reaction. The anion of a salt is supplied by the acid, and the cation by the base.

A **salt bridge** is a liquid junction that consists of a saturated solution of a strong electrolyte, such as KCl. Ions enter and leave the bridge so as to maintain electrical neutrality in the two half-cells of an electrochemical cell.

**Saturated carbons** are involved in four sigma bonds.

**Science** is that branch of knowledge that is gained by the application of the scientific method.

The **Schrödinger equation** relates the energy of an electron to its wavefunction.

The **scientific method** is used to further scientific knowledge. It involves observation, hypothesis formulation, prediction, and testing.

The **second law of thermodynamics** states that the entropy of the universe increases in all spontaneous processes.

A **semiconductor** is a substance whose electrical conductivity increases with temperature. Semiconductors are characterized by small but nonzero band gaps.
A semipermeable membrane allows the passage of solvent molecules but not of solute particles.

**Shell** See level.

**Shielding** is the amount by which the nuclear charge experienced by an electron is reduced by interference from other electrons. Core electrons shield valence electrons much better than do other valence electrons because most of the electron density and charge of the core electrons lies between the valence electrons and the nucleus.

**Sigma bonds** are formed from the interaction of s orbitals or the end-on interaction of p or d orbitals. The electron density in a sigma bond contains the internuclear axis.

**Significant figures** are used to express the precision of a measurement or result.

In a simple cubic (sc) unit cell, the particles are found only at the corners.

The **simplest or empirical formula** is a chemical formula whose subscripts indicate only the smallest whole numbers that are in the same ratio as the actual numbers of atoms present in the molecule.

**Smectic clays** are also called swelling clays, they consist of a layer of aluminite octahedra sandwiched between two layers of silicate tetrahedra.

**Soaps** are similar to detergents except the polar head is a COO⁻ (carboxylate) group because soaps are the salts of fatty acids.

The **solubility** of a solute is the maximum amount of the solute that can dissolve in a solvent at a given temperature.

The **solubility product constant** (Ksp) is the equilibrium constant for the dissolution of a salt in water.

A **solute** is a component of a solution that is not the solvent.

A **solution** is a homogeneous mixture.

**Solvation** is the process in which the solvent molecules interact with solute particles.

The **solvent** is the substance responsible for the phase of a solution. If one of the components of a solution is a liquid, then the liquid is considered the solvent.

An **sp hybrid orbital** is one of the two orbitals obtained by mixing one s and one p orbital on an atom. The two sp hybrids are separated by 180°.

An **sp² hybrid orbital** is one of the three orbitals obtained by mixing one s and two p orbitals on an atom. The three sp² hybrids lie in plane and are separated by 120°.

An **sp³ hybrid orbital** is one of the four orbitals obtained by mixing one s and three p orbitals on an atom. The four sp³ hybrids point toward the corners of a tetrahedron and are separated by 109°.

The **specific heat (s)** of a substance is the amount of heat required to raise the temperature of 1 g of the substance by 1 °C.

**Spectator ions** are ions in solution that do not undergo reaction. When KCl(aq) is added to AgNO₃(aq), the Ag⁺ and Cl⁻ ions react, but the K⁺ and NO₃⁻ ions are spectator ions. Spectator ions are brought into solution as counter ions to the ions that do react.

A **spectrum** is a display of radiant energy arranged in order of it frequency or wavelength.

The **spin quantum number** (m_s) of an electron is +1/2 or -1/2. It indicates the direction of the magnetic field produced by the electron.

A **spontaneous** process is one that takes place without intervention. ΔS_mis > 0 for all spontaneous processes, or ΔG < 0 for spontaneous processes at constant temperature and pressure.

The **standard cell potential** (E°) is the cell potential when all reactants and products are in their standard states.

The **standard enthalpy or heat of reaction** (ΔH°) is the enthalpy change for a reaction when it is carried out with all reactants and products in their standard states.

The **standard heat or enthalpy of formation** (ΔH_f) is the heat absorbed when one mole of a substance is formed from its elements in their standard states.

The **standard hydrogen electrode (SHE)** is a half-cell containing 1 M H⁺ and 1 atm H₂. It is used as the reference for standard reduction potentials. The standard reduction potential of the SHE is assigned a value of exactly 0 V.

The **standard reduction potential** of a redox couple is a measure of the free energy of the redox electrons relative to those in a reference couple such as the H⁺/H₂ couple. The more positive the standard reduction potential, the lower is the energy of the electrons.

A **standard solution** is a solution of known concentration that is used to determine an unknown concentration.

The **standard state** is a reference state used to compare thermodynamic quantities. It is 1 atm pressure for a gas, a concentration of 1 M for a solute, and the pure substance for a solid or a liquid.

A **state function** is a quantity that depends only upon the initial and final states.

A **stereocenter** in organic chemistry is a carbon atom that has four different groups attached to it.
Stereoisomers have the same connectivities but different spatial arrangements of their atoms.

The steric factor in kinetics represents the probability that a collision between the particles in an elementary process have the correct orientation to react.

The stoichiometric factor or link is the conversion factor in a stoichiometric calculation that converts from one substance into another. It is the ratio of subscripts in a chemical formula or the coefficients in a balanced chemical equation.

Stoichiometry is the study of the conversion from one chemical species into a chemically equivalent amount of another. The conversion is made through the use of chemical formulas or balanced chemical equations.

A straight or continuous chain is a chain of atoms in which no atom is bound to more than two other atoms in the chain.

A strong acid is an acid that reacts extensively with water, i.e., one whose acid dissociation (ionization) constant is much greater than one. Aqueous solutions of strong acids are represented by H_3O^+.

A strong base is a base that reacts extensively with water to produce OH^-.

The sublevel of an electron is specified by the n and l quantum numbers. It dictates the energy, size, and shape of its orbitals.

Sublimation is the process in which a solid is converted into its vapor.

A supercritical fluid is the phase of matter beyond the critical point. It has some properties of the liquid and the gas, but it is neither.

The surface tension of a liquid is the energy required to increase its surface area by a fixed amount.

Surroundings See thermodynamic surroundings.

System See thermodynamic system.

Temperature is a measure of the average kinetic energy of the molecules in a system.

A termolecular process involves three molecules.

The theoretical yield is the amount of product predicted from the amount of limiting reactant and the stoichiometry of the reaction.

A theory is an explanation of many observations.

Thermal energy is the kinetic energy of a molecule, ion, or atom. Thermal energy depends only upon the temperature.

The thermal surroundings is that portion of the surroundings that exchanges heat with the system.

A thermochemical equation is a chemical equation that includes a thermodynamic quantity, usually ΔH or ΔG.

Thermochemistry is that branch of thermodynamics that deals with energy change in chemical reactions.

Thermodynamics is the study of energy and its transformations.

The thermodynamic region of a reaction is after equilibrium has been established.

The thermodynamic surroundings is that part of the universe that exchanges energy with the system.

A thermodynamic system is that part of the universe that is under investigation.

The thermodynamic universe is the system and its surroundings.

A thermonuclear reaction is a nuclear reaction that requires a large input of energy for initiation. Fusion reactions are thermonuclear.

The third law of thermodynamics states that the entropy of a perfect crystal at 0K is zero.

The titrant is the solution whose volume is determined in a titration.

In a titration, the volume of one solution of known concentration (the titrant) that is required to react with another solution (the analyte) is determined in order to find the concentration of the analyte.

A titration curve is a plot of the pH of the solution versus the volume of titrant.

The torr is a unit of pressure. A pressure of 1 torr supports a column of Hg to a height of 1 mm.

A trans configuration is one in which two groups are on opposite sides of a bond or atom.

A transition element or metal is an element (metal) in the d-block (B groups) of the periodic table.

The transition state is the highest energy species through which the reactants must pass in order to make the transition to the products.

Translational degrees of freedom are the straight-line motions of a particle. All straight line motion can be expressed as a sum of x, y, and z components, so all molecules have three translational degrees of freedom.

The triple point is the temperature and pressure at which the solid, liquid, and vapor states of a substance are in equilibrium.
The **uncertainty principle** states that it is impossible to know both the position and speed of subatomic particles to high accuracy. In order to measure one more accurately, you most lose accuracy in the other.

A **unimolecular** process involves only one particle.

A **unit cell** is the simplest arrangement of particles that generates the entire lattice when translated in all three dimensions.

**Unsaturated** carbon atoms are involved in less than four sigma bonds.

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The **valence band** is the highest energy filled band containing the valence electrons of a metal.

In **valence bond theory**, bonds arise from the overlap of orbitals on adjacent atoms. The orbitals can be either atomic or hybridized.

**Valence electrons** are those outermost electrons that dictate the properties of the atom and are involved in chemical bonding. They reside in the outermost s sublevel and any unfilled sublevels.

**Valence-shell electron-pair repulsion (VSEPR)** theory is used to explain molecular shapes in terms of electron regions adopting the spatial orientation that minimizes the electron-electron repulsions between them.

The **Van’t Hoff factor** ($i$) relates the colligative concentration to the concentration of the solute. For example, $m_c = im$.

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The **van der Waals radius** is one-half of the distance between identical, nonbonded atoms in a crystal. Atoms that are closer than the sum of their van der Walls radii are assumed to be interacting.

**Vaporization** is the process by which a liquid is converted into its vapor.

The **vapor pressure** of a liquid is the pressure of its vapor in equilibrium with the liquid at a given temperature.

**Vapor pressure lowering ($\Delta P$)** is amount by which the vapor pressure of a solvent is reduced by the addition of a volatile solute.

**Vibrational degrees of freedom** of a molecule are the relative motions of its atoms that result in small oscillating changes in bond lengths and angles.

**Viscosity** is the resistance of a liquid to flow.

**Void space** is unoccupied space.

The **volt** is the SI unit of electrical potential.

$1 \text{ V} = 1 \text{ J} \cdot \text{C}^{-1}$. [Footnote 1]

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The **wave function** of an electron is a function that contains all of the information about the electron.

The **wavelength** ($\lambda$) is the distance between two corresponding points on a wave.

**Wave-particle duality** is a term used to indicate that photons (light) and very small particles, such as electrons, behave as both particles and waves.

A **weak acid** is an acid that does not react extensively with water, *i.e.*, it is an acid with a dissociation constant that is much less than one.

A **weak base** is a substance that reacts only slightly with water to produce hydroxide ions.

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A **weak electrolyte** is a substance whose aqueous solution conducts only a small current of electricity because only a small fraction of weak electrolyte molecules produce ions in water.

**Work** ($w$) is a force through a distance, $w = fd$. Thus, something must move, and there must be a resistance to the movement in order for work to be done. By definition, the symbol $w$ is the work done ON the system, and $-w$ is the work done BY the system.

**X-ray diffraction** is a technique in which x-rays are scattered from atoms in the solid to determine the distances between the atoms and ions in the crystal.

**Zeolites** are aluminosilicates built from tetrahedral $\text{AlO}_4$ and $\text{SiO}_4$ units bridged by oxygen atoms. They are filled with channels and pores, which provide many uses for the material.
Useful tables

Select a table in the bookmark panel to navigate to it.

- Acid-Base Table (Table 12.3)
- Bond Energies and Bond Lengths (Table 5.3)
- Electronegativites (Table 5.1)
- Polyatomic Ions (Table 4.1)
- Prefixes for covalent compounds (Table 5.2)
- Roots for Simple Organic Compounds (Table 13.1)
- Solubility Rules for Ionic Compounds in Water (Table 10.4)
- Standard Reduction Potentials (Table 11.1)