Appendix A
Stoichiometry of Elements and Compounds

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

Divide both sides by 1 ft: $\boxed{\dfrac{\textbf{12 in}}{\textbf{1 ft}}} = \dfrac{\textbf{1 ft}}{\textbf{1 ft}} = 1$

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:

$$\textbf{6 ft} \times \frac{\textbf{12 in}}{\textbf{1 ft}} = \textbf{72 in}$$

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.

$$525\ \text{g} \times \frac{1\ \text{mol}}{63.55\ \text{g}} = 8.26\ \text{mol}$$

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 $6.02\text{x}10^{23}\ \text{mol}^{-1}$, 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.

$$\frac{6.02 \times 10^{23}\ \text{atoms}}{1\ \text{mole of atoms}} \qquad \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \text{mole of molecules}} \qquad \frac{6.02 \times 10^{23}\ \text{trees}}{1\ \text{mole of trees}}$$

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.

$$\mathbf{2.25\ g\ Ca \times \frac{1\ mol\ Ca}{40.08\ g\ Ca} \times \frac{6.02 \times 10^{23}\ Ca\ atoms}{1\ mol\ Ca} = 3.38 \times 10^{22}\ Ca\ atoms}$$

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 1.00x10^{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.

$$\mathbf{1.00 \times 10^{22}\ Br\ atoms \times \frac{1\ mol\ Br}{6.02 \times 10^{23}\ Br\ atoms} \times \frac{79.90\ g\ Br}{1\ mol\ Br} = 1.33\ g\ Br}$$

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 CO_2, contains one carbon atom and two oxygen atoms. This information is contained in the subscripts after each element. A molecule of sucrose ($C_{12}H_{22}O_{11}$) has 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms. The subscripts also indicate the ratios of the elements. A dozen CO_2 molecules have one dozen carbon atoms and two dozen oxygen atoms. A million CO_2 molecules have one million carbon atoms and two million oxygen atoms. A mole of CO_2 molecules (we usually just say "a mole of CO_2") 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^{1+} ions and one mole of Cl^{1-} ions. Ionic compounds with polyatomic ions are somewhat more complicated. One mole of sodium sulfate, Na_2SO_4, contains two moles of Na^{1+} ions and one mole of SO_4^{2-} 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 $Fe(NO_3)_3$. This tells us that for every iron(III) ion (Fe^{3+}), there are three nitrate ions (NO_3^{1-}). 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, CO_2, 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, $C_{12}H_{22}O_{11}$.

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)	=	<u>176.00 g</u>
	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, $Co_3(PO_4)_2$.

Solution:

Determine the contribution from each element and sum.

cobalt(II) phosphate :

Co:	(3 mol)(58.93 g/mol)	=	176.79 g
P:	(2 mol)(30.97 g/mol)	=	61.94 g
O:	(8 mol)(16.00 g/mol)	=	<u>128.00 g</u>
	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 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.

$$\mathbf{0.125\ mol \times \frac{342.30\ g}{1\ mol} = 42.8\ g}$$

Example 9

What is the mass of 2.50 moles of $Mg(ClO_4)_2$?

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.

Magnesium perchlorate : $Mg(ClO_4)_2$

Molar mass :

Mg:	(1 mol)(24.31 g/mol)	=	24.31 g
Cl:	(2 mol)(35.45 g/mol)	=	70.90 g
O:	(8 mol)(16.00 g/mol)	=	128.00 g
	Total	=	223.21 g/mol

$$\mathbf{2.50\ mol \times \frac{223.21\ g}{1\ mol} = 558\ g}$$

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 NH_4NO_3 and its molar mass is:

N:	(2 mol)(14.01 g/mol)	= 28.02 g
H:	(4 mol)(1.008 g/mol)	= 4.03 g
O:	(3 mol)(16.00 g/mol)	= 48.00 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:

$$\mathbf{\frac{28.02\ g}{80.05\ g} = 0.3500}$$

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 :

$$\mathbf{275\ g\ NH_4NO_3 \times \frac{28.02\ g\ N}{80.05\ g\ NH_4NO_3} = 96.3\ g\ N}$$

Notice that grams of NH_4NO_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 : KNO3

Molar mass:

K:	(1 mol)(39.10 g/mol)	=	39.10 g
N:	(1 mol)(14.01 g/mol)	=	14.01 g
O:	(3 mol)(16.00 g/mol)	=	48.00 g
	Total	=	101.11 g/mol

Percent nitrogen:

$$\frac{\textbf{14.01 g N}}{\textbf{101.11 g total}} \times \textbf{100\% = 13.86 \% N}$$

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.

$$137\text{ g } KNO_3 \times \frac{39.10\text{ g K}}{101.11\text{ g } KNO_3} = 53.0\text{ g K}$$

c) **What mass of KNO_3 contains 125 g of potassium?**

Solution:

The given quantity is 125 g of K. Clearly, we need more than 125 g of compound to give us 125 g of K, since the compound is only about 39% K. We can apply the same ratio as in part b, but turning it "upside down" to give us a result in units of grams of KNO_3.

$$125\text{ g K} \times \frac{101.11\text{ g } KNO_3}{39.10\text{ g K}} = 323\text{ g } KNO_3$$

Comment:

We said that multiplication by a conversion factor is multiplication by 1. In this case, 101.11 g of KNO_3 is the same as "the amount of KNO_3 that contains 39.10 g of K". In other words, the numerator and denominator represent the same amount of compound. This gives the effect of multiplying by 1.

A.10 COUNTING ATOMS, IONS AND MOLECULES IN COMPOUNDS

We use Avogadro's number, $6.02\text{x}10^{23}$ 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_2O?**

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 g/mol), and then apply Avogadro's number.

$$10.0\text{ g } H_2O \times \frac{1\text{ mol}}{18.02\text{ g}} \times \frac{6.02 \times 10^{23}\text{ molecules}}{1\text{ mol}} = 3.34 \times 10^{23}\text{ molecules}$$

b) How many hydrogen atoms are contained in 10.0 g of H_2O?

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.

$$\mathbf{10.0\ g\ H_2O \times \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} \times \frac{2\ mol\ H}{1\ mol\ H_2O} \times \frac{6.02 \times 10^{23}\ H\ atom}{1\ mol\ H} = 6.68 \times 10^{23}\ H\ atoms}$$

Comment:

Note that the two calculations are almost identical, the only difference being the term that converts moles of H_2O 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 $Fe_2(SO_4)_3$?

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 mol)(55.85 g/mol)	=	111.70 g
S:	(3 mol)(32.07 g/mol)	=	96.21 g
O:	(12 mol)(16.00 g/mol)	=	<u>192.00 g</u>
	Total	=	399.91 g/mol

$$\mathbf{68.4\ g\ Fe_2(SO_4)_3 \times \frac{1\ mol\ Fe_2(SO_4)_3}{399.91\ g\ Fe_2(SO_4)_3} \times \frac{2\ mol\ Fe^{3+}}{1\ mol\ Fe_2(SO_4)_3} \times \frac{6.02 \times 10^{23}\ Fe^{3+}\ ions}{mol\ Fe^{3+}\ ions} = 2.06 \times 10^{23}\ Fe^{3+}\ ions}$$

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 4.75×10^{25} carbon atoms?

Solution:

We are given the number of C 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 g/mol from Example 6). Moles of sucrose can be found from the moles of carbon atoms and the subscripts in the chemical formula. Moles of carbon atoms can be found from the number of individual atoms and Avogadro’s number.

$$\mathbf{4.75 \times 10^{25}\ C\ atoms \times \frac{1\ mol\ C}{6.02\ x\ 10^{23}\ C\ atoms} \times \frac{1\ mol\ C_{12}H_{22}O_{11}}{12\ mol\ C} \times \frac{342.30\ g\ C_{12}H_{22}O_{11}}{1\ mol\ C_{12}H_{22}O_{11}} = 2.25 \times 10^{3}\ g = 2.25\ kg\ C_{12}H_{22}O_{11}}$$

Comment: As always, each conversion factor cancels 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 $6.02 x 10^{22}$ 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 $7.5 x 10^{21}$ 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 $2.5 x 10^{-4}$ 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 $8.8 x 10^{24}$ 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 $2.5 x 10^{21}$ argon atoms?
18. How many aluminum atoms are contained in 1.5 kg of aluminum.
19. What is the mass of $3.5 x 10^{25}$ iron atoms?
20. What is the mass, in grams, of $8.25 x 10^{23}$ 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_2H_4
    - b) acetic acid, $HC_2H_3O_2$
    - c) pyridine, C_5H_5N
    - d) succinic acid, $C_4H_6O_4$
    - e) TNT, $C_7H_5N_3O_6$
    - f) calcium nitrate, $Ca(NO_3)_2$
    - g) potassium chromate, K_2CrO_4
    - h) cobalt(II)citrate, $Co_3(C_6H_8O_7)_2$
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 $2.0 x 10^{-5}$ 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)** $C_6H_{12}O_6$ **c)** C_7H_8

Combined skills / compounds

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_4H_4S (M_m = 84.14 g/mol)?
28. How many sulfur atoms are contained in 3.55 g of C_4H_4S?
29. How many carbon atoms are contained in 3.55 g of C_4H_4S?
30. If you wanted to obtain $1.00x10^5$ g of nitrogen, what mass of NH_3 would you need?
31. If you wanted to obtain $1.00x10^5$ g of nitrogen, what mass of NH_4NO_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, $C_{20}H_{24}N_2O_2$ (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. $3.3x10^{25}$ 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)** 79.1 **d)** 118.1
    e) 227.1 **f)** 164.1 **g)** 194.2 **h)** 560.9
23. **a)** 0.0320 mol CoI_2 **b)** 0.0057 mol morphine
    c) $1.2x10^{-4}$ g morphine **d)** 3.75 kg CoI_2
24. **a)** 75.0% **b)** 40.0 % **c)** 91.3 %
25. 4.19 g K
26. $6.46x10^{22}$ Cl^{1-} ions
27. 381 g S
28. $2.54x10^{22}$ atoms of S
29. $1.02x10^{23}$ 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