1. What is the mass of 5.00 × 10[20] atoms of Cr?

$$5.00 \times 10^{20} \text{ atoms Cr} \times \frac{1 \text{ mol Cr}}{6.022 \times 10^{23} \text{ atoms Cr}} \times \frac{52.00 \text{ g}}{1 \text{ mol Cr}} = 4.32 \times 10^{-2} \text{ g}$$

2. How many moles are in a 50.0 g sample of ammonium carbonate?

$$50.0 \text{ g } (NH_4)_2CO_3 \times \frac{1 \text{ mol } (NH_4)_2CO_3}{96.094 \text{ g}} = 0.520 \text{ mol } (NH_4)_2CO_3$$

3. What is the mass of one molecule of dinitrogen tetroxide?

$$1 \text{ molecule } N_2O_4 \times \frac{1 \text{ mol } N_2O_4}{6.022 \times 10^{23} \text{ molecules } N_2O_4} \times \frac{92.02 \text{ g}}{1 \text{ mol } N_2O_4} = 1.528 \times 10^{-22} \text{g}$$

4. Balance the following two reactions.

__3__ H_2 (*g*) + __1__ N_2 (*g*) → __2__ NH_3 (*g*)

__2__ $Pb(N_3)_2$ (*s*) + __1__ O_2 (*g*) → __2__ PbO (*s*) + __6__ N_2 (*g*)

5. A compound is composed of only C, H, and N atoms.

A) Determine the empirical formula if the compound is found to be 74.1% C and 8.70% H by mass.

Convert percentages to masses assuming 100 g of $C_xH_yO_z$ sample. Convert from masses to moles of each element. Then divide by the smallest mole amount (1.228 mol O) to get mole ratio.

74.1 g C	→	6.169 mol C		~5 mol C
8.70 g H	→	8.632 mol H	→	~7 mol H
17.2 g N	→	1.228 mol N		1 mol N

→ C_5H_7N (molar mass = 81.116 g/mol)

B) If the mass of 0.123 moles of the compound has a mass of 19.94 g, what is the molecular formula of the compound?

Recognize molecular mass as 19.94 g per 0.123 moles = 162.11 g/mol

Then, take the ratio between the molecular and empirical mass to extract n, $(C_5H_7N)_n$:

$$n = \frac{162.11 \text{ g/mol}}{81.116 \text{ g/mol}} \approx 2$$

Final answer: $C_{10}H_{14}N_2$

6. You react 10.0 g of hydrogen gas with 60.0 g of oxygen gas to form water vapor. How much water can be formed from this reaction?

   Balanced chemical equation: $2\ H_2\ (g) + 1\ O_2\ (g) \rightarrow 2\ H_2O\ (g)$

   Now, determine limiting reactant is O_2:

Method 1:

How much H_2O can we make from all H_2?

$$10.0\ g\ H_2 \times \frac{1\ mol\ H_2}{2.016\ g\ H_2} \times \frac{2\ mol\ H_2O}{2\ mol\ H_2} = 4.96\ mol\ H_2O$$

How much H_2O can we make from all O_2?

$$60.0\ g\ O_2 \times \frac{1\ mol\ O_2}{32.00\ g\ O_2} \times \frac{2\ mol\ H_2O}{1\ mol\ O_2} = 3.75\ mol\ H_2O$$

Therefore, O_2 is limiting (makes less H_2O)!

Method 2:

We have 4.96 mol H_2 and 1.875 mol O_2. How much H_2 do we need to react with all of the O_2?

$$60.0\ g\ O_2 \times \frac{1\ mol\ O_2}{32.00\ g\ O_2} \times \frac{2\ mol\ H_2}{1\ mol\ O_2} = 3.75\ mol\ H_2$$

We need 3.75 mol H_2, but have 4.96 mol H_2, so H_2 is in excess.

Therefore, O_2 is limiting!

From method 1 above, we can make 3.75 mol H_2O or (67.6 g).

7. Solid iron(III) oxide reacts with carbon monoxide to form elemental iron and carbon dioxide gas.

   A) Write a balanced chemical equation for the reaction described above.

   __1__ $Fe_2O_3\ (s)$ + __3__ $CO\ (g) \rightarrow$ __2__ $Fe\ (s)$ + __3__ $CO_2\ (g)$

   B) How much iron metal is obtained if 433.2 g of iron(III) oxide reacts with 250. L of carbon monoxide (density of carbon monoxide is 1.145 g/L).

   Determine limiting reactant is Fe_2O_3 through 1:3 mole ratio (2.7128 mol Fe_2O_3 and 10.22 mol CO).
   Determine how much Fe (s) can be produced from the limiting reactant:

$$433.2\ \text{g}\ Fe_2O_3 \times \frac{1\ \text{mol}\ Fe_2O_3}{159.687\ \text{g}} \times \frac{2\ \text{mol Fe}}{1\ \text{mol}\ Fe_2O_3} = 5.426\ \text{mol Fe (or 303.0 g Fe)}$$

   C) How much starting material would be left over after the reaction is complete?

   No Fe_2O_3 leftover because it is the limiting reactant.

$$433.2\ \text{g}\ Fe_2O_3 \times \frac{1\ \text{mol}\ Fe_2O_3}{159.687\ \text{g}} \times \frac{3\ \text{mol CO}}{1\ \text{mol}\ Fe_2O_3} \times \frac{28.011\ \text{g}}{1\ \text{mol CO}} \times \frac{1\ \text{L CO}}{1.145\ \text{g}} = 199.096\ \text{L CO used}$$

$$250.\ \text{L} - 199.096\ \text{L} = 51\ \text{L CO left}$$

8. If you have equal mass samples of each of the following compounds, which sample contains the greatest number of oxygen atoms?

H_2SO_4	$C_{12}H_{22}O_{11}$	$KClO_3$
$\frac{4 \times 16.00\ \text{g}}{98.076\ \text{g}} \times 100 = 62.26\ \%$	$\frac{11 \times 16.00\ \text{g}}{342.296\ \text{g}} \times 100 = 51.42\ \%$	$\frac{3 \times 16.00\ \text{g}}{122.55\ \text{g}} \times 100 = 39.17\ \%$

Since masses are the same, the mass percent of O will be an indication of O atoms in the samples. Therefore, H_2SO_4 contains the greatest number of O atoms.

Note: You can also solve this by assuming an arbitrary mass (e.g. 100 g) of each, then determining the number of O atoms through dimensional analysis.