Work together to answer all the following questions. It is not necessarily an exhaustive review of what we have covered so far, but it can certainly help.

1. Magnesium is essential to life on earth. I'm curious to know more about this element:
    a. What is the **atomic number** of magnesium? 12
    b. How many **electrons** does magnesium have? 12
    c. How many **protons** does magnesium have? 12
    d. What is the name of the group of elements magnesium belongs to? Alkaline Earth Metals
    e. What period is magnesium in? 3^{rd} period
    f. How many electron shells does magnesium possess? 3 electron shells
    g. How many valence electrons does magnesium possess? 2 valence electrons

2. Wow! Though I'd still like to know more about magnesium:
    a. Magnesium-24 is the most common isotope of magnesium. How many **neutrons** does it have? 12 neutrons
    b. What is the purpose of having neutrons in an atom? To stabilize the nucleus by reducing proton-proton repulsion
    c. What is the mass number for the magnesium-24 isotope? 24
    d. Which noble gas is closest to magnesium? Neon (Ne)
    e. How many electrons does magnesium gain or lose to achieve the same electron configuration as a noble gas? It would lose two electrons

3. Isn't magnesium such an interesting element? With winter quickly coming towards us we can even use it to help salt our roads. If an ionic compound of magnesium and chlorine is made:
    a. What is the **chemical formula** of this salt? $MgCl_2$
    b. What is the **name** of this salt? Magnesium chloride
    c. What is the **charge** of the magnesium ion? 2+
    d. What is the **charge** of the chlorine ion? 1-
    e. If you have **1.00 g of this salt**, how many atoms of chlorine do you have?

$$1.00\ g\ MgCl_2 \left(\frac{1\ mol\ MgCl_2}{95.211\ g\ MgCl_2}\right)\left(\frac{2\ mol\ Cl}{1\ mol\ MgCl_2}\right)\left(\frac{6.022\times 10^{23}\ atoms\ Cl}{1\ mol\ Cl}\right)$$
$$= 1.26\times 10^{22}\ atoms\ Cl$$

4. The salt described above can even be used as an aqueous solution to help in the purification of water to remove lead ions by precipitation. If you combine the above salt with a solution of lead(II) nitrate, you will find a precipitate will form.
   a. What is the **complete molecular equation** of this reaction?

$$MgCl_{2(aq)} + Pb(NO_3)_{2(aq)} \rightarrow Mg(NO_3)_{2(aq)} + PbCl_{2(s)}$$

   b. What is the **complete ionic equation** of this reaction?

$$Mg^{2+}{}_{(aq)} + 2Cl^{-}{}_{(aq)} + Pb^{2+}{}_{(aq)} + 2NO_3{}^{-}{}_{(aq)} \rightarrow Mg^{2+}{}_{(aq)} + 2NO_3{}^{-}{}_{(aq)} + PbCl_{2(s)}$$

   c. What is the **net ionic equation** of this reaction?

$$2Cl^{-}{}_{(aq)} + Pb^{2+}{}_{(aq)} \rightarrow PbCl_{2(s)}$$

   d. Does this reaction have spectator ions? If yes, what are they? Yes, the spectator ions are the magnesium and nitrate ions.

5. If you have a 500.0 mL sample of water with a $[Pb(NO_3)_2] = 0.025\ M$, and you add 100.0 mL of the salt solution with a concentration of 0.250 M:
   a. What is the **chemical formula** of the precipitate that forms? $PbCl_2$
   b. What is the **name of the precipitate** that forms? Lead(II) chloride
   c. What is the mass of the precipitate **in mg**?

$$100.0\ mL\ MgCl_2 \left(\frac{1\ L}{1000\ mL}\right)\left(\frac{0.250\ mol\ MgCl_2}{1\ L}\right) = 0.0250\ mol\ MgCl_2$$

$$500.0\ mL\ Pb(NO_3)_2 \left(\frac{1\ L}{1000\ mL}\right)\left(\frac{0.025\ mol\ Pb(NO_3)_2}{1\ L}\right)$$
$$= 0.0125\ mol\ Pb(NO_3)_2\ (should\ be\ 2\ s.f.)$$

Determine limiting reagent.

Assuming $MgCl_2$ is limiting:

$$0.0250\ mol\ MgCl_2 \left(\frac{1\ mol\ PbCl_2}{1\ mol\ MgCl_2}\right) = 0.0250\ mol\ PbCl_2$$

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$$0.0125\ mol\ Pb(NO_3)_2 \left(\frac{1\ mol\ PbCl_2}{1\ mol\ Pb(NO_3)_2}\right) = 0.0125\ mol\ PbCl_2\ (should\ be\ 2\ s.f.)$$

Lead(II) nitrate leads to less product and is the limiting reagent.

To determine mass of the precipitate:

$$0.0125\ mol\ PbCl_2 \left(\frac{278.106\ g\ PbCl_2}{1\ mol\ PbCl_2}\right)\left(\frac{1000\ mg\ PbCl_2}{1\ g\ PbCl_2}\right) = 3476.325\ mg\ PbCl_2$$

$$\therefore 3500\ mg\ PbCl_2\ (with\ sig\ fig\ rounding)$$

   d. What is the concentration of magnesium ions left in solution after reaction?

Since Mg^{2+} is a spectator ion, the moles do not change from before the reaction. The total volume of solution has changed since the two solutions were combined.

$$0.0250\ mol\ MgCl_2 \left(\frac{1\ mol\ Mg^{2+}}{1\ mol\ MgCl_2}\right) = 0.0250\ mol\ Mg^{2+}$$

$$Total\ volume = 500.0\ mL + 100.0\ mL = 600.0\ mL = 0.6000\ L$$

$$[Mg^{2+}] = \frac{0.0250\ mol\ Mg^{2+}}{0.6000\ L} = 0.0416667\ M$$

$$\therefore [Mg^{2+}] = 0.0417\ M\ (with\ sig\ figs)$$

e. What is the concentration of chlorine ions left in solution after reaction?

Chloride is involved in the chemical reaction and becomes part of the precipitate. Any chloride in the precipitate is no longer in solution so we must determine how much chloride is leftover. Since it comes from the excess reagent there should be some left. To determine this we need to calculate how many moles of chloride are reacted with the lead(II).

$$0.0125\ mol\ Pb(NO_3)_2 \left(\frac{1\ mol\ MgCl_2}{1\ mol\ Pb(NO_3)_2}\right)\left(\frac{2\ mol\ Cl^-}{1\ mol\ MgCl_2}\right) = 0.0250\ mol\ Cl^-\ reacted$$

$$0.0250\ mol\ MgCl_2 \left(\frac{2\ mol\ Cl^-}{1\ mol\ MgCl_2}\right) = 0.0500\ mol\ Cl^-\ initially$$

$$n_{Cl^-{}_{leftover}} = n_{Cl^-{}_{initial}} - n_{Cl^-{}_{reacted}} = 0.0500\ mol\ Cl^- - 0.0250\ mol\ Cl^- = 0.0250\ mol\ Cl^-$$

$$[Cl^-] = \frac{0.0250\ mol\ Cl^-}{0.6000\ L} = 0.041667\ M$$

$$\therefore [Cl^-] = 0.0417\ M\ (with\ sig\ figs)$$

f. What is the concentration of nitrate ions left in solution after reaction?

By the same logic as with the magnesium ion:

$$[NO_3{}^-] = 0.0417\ M\ (with\ sig\ figs)$$

g. What is the concentration of lead ions left in solution after reaction?

The lead ions are part of the precipitate (removed from solution), and originally came from the limiting reagent so there would be no moles of that ion still present in solution:

$$[Pb^{2+}] = 0\ M$$

I think we've talked enough about magnesium for now, but there's so much more that can be said. Now we'll look at something separate. Let's look at combustion analysis!

6. First, let's look at a combustion reaction. We'll use the example of burning acetylene (C_2H_2) gas in the presence of oxygen gas, and it makes carbon dioxide gas and water vapour as products.
   a. What is the balanced chemical equation?

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(g)}$$

   b. Assign the oxidation states of every element in this reaction.

$$C_2H_2 : C = -1, H = +1$$

$$O_2 : O = 0$$

$$CO_2 : C = +4, O = -2$$

$$H_2O : H = +1, O = -2$$

   c. What are the reducing and oxidizing agents?

$$Oxidizing\ agent: O_2$$

$$Reducing\ agent: C_2H_2$$

   d. Is this an example of a precipitation, acid-base, or redox reaction? Redox reaction
   e. How can oxidation states be used to help determine the type of reaction? If the oxidation states change before and after the chemical reaction, then it must be a redox reaction. If the states do not change, then it cannot be a redox reaction.

7. Now, lets take a look at an actual combustion analysis. If we have a 300.0 mg sample of a compound that contains carbon, hydrogen, and oxygen that we want to analyze, we will have to combust it. After combustion we recovered 694 mg of carbon dioxide and 142 mg of water.
   a. What is the empirical formula of this compound?

First, we need to use the data from CO_2 and H_2O to determine the amount of carbon and hydrogen in the original sample, respectively.

$$694\ mg\ CO_2 \left(\frac{1\ g\ CO_2}{1000\ mg\ CO_2}\right)\left(\frac{1\ mol\ CO_2}{44.011\ g\ CO_2}\right)\left(\frac{1\ mol\ C}{1\ mol\ CO_2}\right) = 0.015775\ mol\ C$$

$$142\ mg\ H_2O \left(\frac{1\ g\ H_2O}{1000\ mg\ H_2O}\right)\left(\frac{1\ mol\ H_2O}{18.016\ g\ H_2O}\right)\left(\frac{2\ mol\ H}{1\ mol\ H_2O}\right) = 0.015775\ mol\ H$$

Next, to determine the amount of oxygen we cannot use the information from CO_2 and H_2O since we have no way to know if the oxygen in those compounds came from the sample or the oxygen gas pumped in to allow combustion. However, we can use the mass of the sample, as well as determine the mass of carbon and hydrogen to determine the mass of oxygen.

$$0.015775\ mol\ C \left(\frac{12.011\ g\ C}{1\ mol\ C}\right) = 0.189458\ g\ C$$

$$0.015775\ mol\ H \left(\frac{1.008\ g\ H}{1\ mol\ H}\right) = 0.015900\ g\ H$$

$$m_{sample} = m_C + m_H + m_O$$

$$m_O = m_{sample} - m_C - m_H = 0.3000\ g - 0.189458\ g - 0.015901\ g = 0.094642\ g\ O$$

Now we will need the number of moles of oxygen.

$$0.094642\ g\ O \left(\frac{1\ mol\ O}{15.999\ g\ O}\right) = 0.005915\ mol\ O$$

With the number of moles of each element we have to normalize them to create a ratio. To do this we divide each one by whichever is the smallest value, in this case oxygen.

$$\frac{0.015774\ mol\ C}{0.005915\ mol\ O} = 2.666667\ (not\ within\ 0.15\ of\ a\ whole\ number, cannot\ round)$$

$$\frac{0.015774\ mol\ H}{0.005915\ mol\ O} = 2.666667\ (not\ within\ 0.15\ of\ a\ whole\ number, cannot\ round)$$

$$\frac{0.005915\ mol\ O}{0.005915\ mol\ O} = 1$$

Since the values have to be whole numbers, and we cannot round, we will need to multiply all three by the same whole number. In this case the value of 3 would work, and that gives use the formula $C_8H_8O_3$

b. If the actual compound has a molecular mass of 152.144 g/mol, what is its molecular formula?

To determine the molecular formula we have to compare the molecular mass with the empirical formula mass (the molar mass of the empirical formula). This gives us the ratio that we will need to multiply our empirical formula by.

$$ratio = \frac{molecular\ mass}{empirical\ formula\ mass} = \frac{152.144\frac{g}{mol}}{152.144\frac{g}{mol}} = 1$$

This means that the molecular formula is the same as the empirical formula.

$$C_8H_8O_3$$