1) An unknown sample in the laboratory contains sodium chloride, NaCl. To determine the percentage of NaCl in the unknown, the sample was dissolved in 50.0 mL of distilled water. The sample was then added to an excess silver nitrate, $AgNO_3$, solution to form a precipitate. The precipitate was then filtered, washed, and dried to a constant mass. Using the information below, calculate the percentage of NaCl in the unknown sample.

Mass of unknown sample	1.453 grams
Mass of dry filter paper	0.862 grams
Mass of precipitate and filter paper after 1st drying	3.565 grams
Mass of precipitate and filter paper after 2nd drying	3.081 grams
Mass of precipitate and filter paper after 3rd drying	3.082 grams

(a) Write a balanced net-ionic chemical reaction for the formation of the precipitate above.
(b) Calculate the mass of the precipitate.
(c) Calculate the moles of the precipitate.
(d) Determine the moles of the sodium chloride in the unknown sample.
(e) Calculate the mass of the sodium chloride in the unknown sample.
(f) Calculate the percentage of sodium chloride in the unknown sample.

(a) $Ag^+ + Cl^- \rightarrow AgCl\ (s)$

(b) 3.082 grams AgCl and FP – 0.862 grams FP = 2.22 grams AgCl (s)

(c) 2.22 grams AgCl (s) / 143.32 g/mol = 0.0155 moles AgCl (s)

(d) 1 mole AgCl : 1 mole Cl^- : 1 mole NaCl

0.0155 moles AgCl : 0.0155 moles Cl^- : 0.0155 moles NaCl

(e) 0.0155 moles NaCl · 58.44 g/mol = 0.906 grams NaCl

(f) 0.906 grams NaCl / 1.453 grams unknown · 100% = 62.4% of NaCl

2) A teacher provided their students with an unknown alkaline earth metal ion chloride (XCl_2). The students dissolved the unknown sample in 50.0 mL of water. The students added using a buret a known solution of sodium sulfate and measured the conductivity to know when the chemical reaction was complete. At the completion of the chemical reaction, the students had added 17.2 mL of 0.50 *M* Na_2SO_4 (aq). The students then filtered, washed, and dried the precipitate accordingly and measure the mass to be 1.579 grams.

(a) Write a balanced net-ionic chemical reaction using X for the unknown.
(b) As the sodium sulfate was being added to the unknown (prior to completion of the reaction), was the conductivity increasing, decreasing, or remaining the same? Explain your answer.
(c) Calculate the number of moles of Na_2SO_4 that was added by the completion of the reaction.
(d) Determine the moles of precipitate was made at the completion of the chemical reaction?
(e) Calculate the molar mass of the precipitate.
(f) What is the identity of the unknown alkaline earth metal ion?

(a) $X^{2+} + SO_4^{2-} \rightarrow XSO_4$ (s)

(b) The conductivity was decreasing. When a precipitate is being formed, an ionic compound of zero overall charge is being formed. The aqueous ions are decreasing to form a precipitate, so the conductivity is decreasing as there are fewer mobile ions to conduct electricity.

(c) 0.50 *M* of Na_2SO_4 · 0.0172 L = 0.00860 moles of Na_2SO_4 added

(d) 1 mole Na_2SO_4 : 1 mole of SO_4^{2-} : 1 mole of XSO_4 (s)

0.00860 moles Na_2SO_4 : 0.00860 moles SO_4^{2-} : 0.00860 moles XSO_4 (s)

(e) Molar mass of XSO_4 = 1.579 grams XSO_4 / 0.00860 moles XSO_4 = 183.6 g/mol XSO_4

(f) 183.6 g/mol XSO_4 – 96.06 g/mol SO_4^{2-} = 87.5 g/mol X^{2+}

The identity of X^{2+} is Sr^{2+}.

3) A 50.0 mL unknown aqueous sample was tested in the laboratory. The unknown sample was identified to be lead (II) ions by a precipitation reaction with sodium iodide. Since the unknown was identified to have lead (II) ions in it, a student added excess sodium chloride (NaCl) to the 50.0 mL sample. A white precipitate was made and was subsequently filtered, washed, and dried to completion. The precipitate was measure to have a mass of 2.350 grams..

(a) How did the students determine through the addition of sodium iodide that the sample had lead (II) ions in it?
(b) Write the balanced net-ionic chemical reaction for the unknown sample and sodium chloride.
(c) Calculate the moles of the precipitate.
(d) Determine the moles of lead (II) ions are present in the unknown solution?
(e) Calculate the concentration in Molarity of the lead (II) ions in the unknown solution.

(a) When lead (II) ions are mixed with sodium iodide, a yellow precipitate is formed.

(b) $\mathbf{Pb^{2+} + 2\ Cl^{-} \rightarrow PbCl_2\ (s)}$

(c) 2.350 grams $\mathbf{PbCl_2}$ **/ 278.1 g/mol = 0.00845 moles** $\mathbf{PbCl_2}$ **(s)**

(d) 1 mole $\mathbf{PbCl_2}$ **: 1 mole** $\mathbf{Pb^{2+}}$

0.00845 moles $\mathbf{PbCl_2}$ **: 0.00845 moles** $\mathbf{Pb^{2+}}$

(e) 0.00845 moles $\mathbf{Pb^{2+}}$ **/ 0.050 L = 0.169 *M*** $\mathbf{Pb^{2+}}$

4) A teacher gave their students an 8.00 gram tablet that contains magnesium nitrate, $Mg(NO_3)_2$ (molar mass = 148.32 g/mol). In order to determine the percentage of magnesium nitrate in the tablet, the students dissolved the tablet in 100. mL of distilled water and add excess sodium phosphate, Na_3PO_4. After a precipitate is formed, the students filter, wash, and dry the precipitate to completion. The precipitate is measured to have a mass of 3.524 grams. (Note: the precipitate will have a molar mass of 262.84 g/mol)

(a) Write a balanced net-ionic chemical reaction for the formation of the precipitate above.
(b) Calculate the moles of the precipitate.
(c) Determine the moles of the magnesium nitrate in the tablet.
(d) Calculate the mass of the magnesium nitrate in the tablet.
(e) Calculate the percentage of magnesium nitrate in the tablet.

(a) 3 Mg^{2+} + 2 PO_4^{3-} → $Mg_3(PO_4)_2$ (s)

(b) 3.524 grams $Mg_3(PO_4)_2$ (s) / 262.84 g/mol = 0.0134 moles $Mg_3(PO_4)_2$ (s)

(c) 1 mole $Mg_3(PO_4)_2$ (s) : 3 moles Mg^{2+}

0.0134 moles $Mg_3(PO_4)_2$ (s) : 0.0402 moles Mg^{2+}

1 mole Mg^{2+} : 1 mole $Mg(NO_3)_2$

0.0402 mole Mg^{2+} : 0.0402 moles $Mg(NO_3)_2$

(d) 0.0402 moles $Mg(NO_3)_2$ ·148.32 g/mol = 5.96 grams $Mg(NO_3)_2$

(e) 5.96 grams $Mg(NO_3)_2$ / 8.00 grams tablet · 100% = 74.5% of $Mg(NO_3)_2$