

Chapter 4 Multiple Choice Review Key

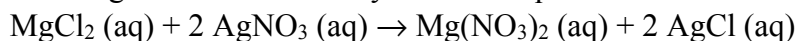
- E. Only KOH will dissociate into ions in aqueous solution.
- B. CH_3COOH is a weak acid, so dissociates slightly in aqueous solution.
- E. Na_2CO_3 is an ionic compound containing 2 Na^+ ions and one CO_3^{2-} ion.
- B. Mobile ions (charge carriers) respond to an electric field to allow electrical conduction.
- B. SO_4^{2-} is insoluble with Ba^{2+} .
- B. All compounds containing Na^+ ions are soluble.
- B. PbCl_2 is insoluble, so the reaction forming PbCl_2 from Pb^{2+} and 2 Cl^- ions is correct.
- E. Mg^{2+} requires 2 Br^- ions to form MgBr_2 .
- C. SO_4^{2-} requires 2 K^+ ions to form K_2SO_4 .
- A. In K_2SO_4 , each K has a +1 oxidation number, each oxygen has a -2 oxidation number, so algebraically, the oxidation number on S is given by $2(+1) + x + 4(-2) = 0$, or $x = +6$.
- B. N has 5 valence electrons, so losing all 5 results in a +5 oxidation number.
- C. In double-replacement reactions, ions change places but no oxidation numbers are changed.
- B. In $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$, the oxidation number of Mn changes from +7 to +2, so it is reduced; hence MnO_4^- is the oxidizing agent.
- A. In $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+}$, the oxidation number of Fe changes from +2 to +3, so it is oxidized and is the reducing agent.
- A. Elemental Fe will reduce Cu^{2+} to elemental Cu and be oxidized to Fe^{2+} , forming Cu and FeSO_4 .
- E. Elemental Al reduces H^+ to H_2 and is oxidized to Al^{3+} , replacing H^+ in the compound.
- B. Convert each solution to the number of moles of Cl^- , then add the moles and divide by the new total volume to obtain the concentration:

$$? \text{ mol Cl}^-_{\text{NH}_4\text{Cl}} = 25.0 \text{ mL} \times \frac{0.2450 \text{ mol Cl}^-}{1000 \text{ mL}} = 6.13 \times 10^{-3} \text{ mol}$$

$$? \text{ mol Cl}^-_{\text{FeCl}_3} = 55.5 \text{ mL} \times \frac{3 \text{ Cl}^- \text{ ions} \times 0.1655 \text{ mol Cl}^-}{1000 \text{ mL}} = 2.76 \times 10^{-2} \text{ mol}$$

$$? M_{\text{Cl}^-} = \frac{(6.13 \times 10^{-3} + 2.76 \times 10^{-2}) \text{ mol}}{(25.0 + 55.5) \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = \boxed{0.419 \text{ M}}$$

- B. Assume that the MgCl_2 is the limiting reagent, so all Cl^- ends up in the AgCl precipitate. Simply determine the moles of Mg^{2+} stoichiometrically from the equation



to determine the concentration:

$$? M_{\text{Mg}^{2+}} = 0.1183 \text{ g AgCl} \times \frac{1 \text{ mol AgCl}}{143.32 \text{ g AgCl}} \times \frac{1 \text{ mol MgCl}_2}{2 \text{ mol AgCl}} \times \frac{1 \text{ mol Mg}^{2+}}{1 \text{ mol MgCl}_2} \times \frac{1}{0.0500 \text{ L}} = \boxed{8.25 \times 10^{-3} \text{ M}}$$

- A. Since all of the I^- and the H_2O_2 react in the given reaction, neither is limiting. Determine the moles of H_2O_2 stoichiometrically and divide by the volume to determine the concentration:

$$? M \text{ H}_2\text{O}_2 = 37.12 \text{ mL I}^- \times \frac{0.1500 \text{ M I}^-}{1000 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}_2}{2 \text{ mol I}^-} \times \frac{1}{0.05000 \text{ L}} = \boxed{5.568 \times 10^{-2} \text{ M}}$$

- C. First determine the mol H^+ used in the given reaction, then subtract from the moles H^+ in the original solution and divide by the volume to find the concentration:

$$? \text{ mol H}^+ \text{ used} = 12.7 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \times \frac{2 \text{ mol H}^+}{1 \text{ mol Zn}} = 0.388 \text{ mol H}^+$$

$$? \text{ mol H}^+ \text{ left} = \underbrace{500. \text{ mL HCl} \times \frac{1.450 \text{ M HCl}}{1000 \text{ mL}} \times \frac{1 \text{ mol H}^+}{1 \text{ mol HCl}}}_{0.725 \text{ mol H}^+} - 0.388 \text{ mol H}^+ = 0.377 \text{ mol H}^+$$

$$? M \text{ H}^+ = \frac{0.377 \text{ mol H}^+}{0.500 \text{ L}} = \boxed{0.674 \text{ M}}$$