BIG IDEA 3: CHEMICAL REACTIONS

Big Idea 3

Changes in matter involve the rearrangement and/or reorganization of atoms and/or the transfer of electrons.

The two main kinds of changes studied in chemistry are phase changes (melting, evaporating, freezing) and rearrangement of atoms. Both processes are accompanied either by the release or absorption of energy. In Chapter 7 you learn about the types of reactions, balancing equations, predicting products of chemical reactions, and writing complete and net ionic equations. In Chapter 8 you learn about stoichiometry—calculations relating amounts of reactants and products. Chapter 9 focuses on electrochemistry, a specific application of oxidation–reduction processes. You need to have a qualitative understanding of the energy changes associated with a reaction or process, but specific calculations are covered in Big Idea 5.

You should be able to

- Classify reactions by type.
- Write balanced molecular equations, complete ionic equations, and net ionic equations.
- Predict if a precipitate will form (the only solubility rule you need to memorize is that all Na^+, K^+, NH_4^+, and NO_3^- salts are soluble).
- Identify compounds as Brønsted–Lowry acids and bases based on evidence of proton transfer in the balanced reaction equation.

- Identify a reaction as oxidation-reduction based on evidence of electron transfer, and balance the reaction equation.
- Predict products of reactions, given the chemical names of the reactants.

> **AP Tip**
>
> Be sure to know the most common charges of transition metal ions and the physical states of elements at room temperature (i.e., if calcium is placed into water, write as Ca not Ca^{2+}). Write solids, liquids, and gases in molecular form (i.e., write hydrogen chloride gas as HCl, not as separate ions).

EVIDENCE OF A CHEMICAL REACTION

You should be able to recognize macroscopic evidence that a chemical reaction has occurred, including:

Releasing heat or light (exothermic) or cooling off (endothermic)
(*Chemistry* 8th ed. page 238/9th ed. page 248)

Formation of a solid when two aqueous solutions are combined (precipitation)
(*Chemistry* 8th ed. page 145/9th ed. page 153)

Color change, usually because of an indicator in an acid–base reaction or of the formation of products differently colored than the reactants
(*Chemistry* 8th ed. pages 157–158, 728–733/9th ed. pages 166–167, 742–747)

Gas formation, usually in acid–base reactions and in oxidation–reduction reactions

EXAMPLE: $2\ HCl(aq) + CaCO_3(s) \rightarrow H_2O(l) + CaCl_2(aq) + CO_2(g)$

EXAMPLE: $2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$

DESCRIBING REACTIONS IN AQUEOUS SOLUTION

(*Chemistry* 8th ed. pages 150–151/9th ed. pages 158–160)

MOLECULAR EQUATION

A molecular equation gives the overall reaction. It gives information on stoichiometry, and you can deduce whether or not the compounds exist as ions in solution by looking at the phases of matter. For example, in the following equation, both reactants [$KCl(aq)$, $Pb(NO_3)_2(aq)$] are ions, and one product $KNO_3(aq)$ exists as ions: $2\ KCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\ KNO_3(aq)$.

COMPLETE IONIC EQUATION

The complete ionic equation gives the equation including all the ions in solution. For example, $2\ K^+ + 2\ Cl^- + Pb^{2+} + 2\ NO_3^- \rightarrow PbCl_2(s) + 2\ K^+ + 2\ NO_3^-$. The states are aqueous, unless otherwise indicated. Writing species as ions implies that they are in aqueous solution.

NET IONIC EQUATION

When writing a net ionic equation, eliminate the ions that are the same on both reactant and product side; these are called "spectator ions." Only those species that undergo a chemical change are included. The equation must still be balanced!

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

TYPES OF REACTIONS

Classification of chemical reactions by type allows the prediction of products if given the reactants.

Main Reaction Types
Precipitation
Acid-base
Redox (oxidation-reduction), including Combustion
Synthesis/Decomposition (may also be Redox)

PRECIPITATION REACTIONS

(Chemistry 8th ed. pages 154–165/9th ed. pages 163–174)

Precipitation reactions involve the formation of a solid when two solutions are mixed.

PREDICTING PRECIPITATES

When two solutions of ionic compounds are mixed, exchange anions to form products, remembering that each product formed must be neutral. If no solid is formed, there is no chemical reaction! The only solubility rule you must memorize is that all Na^+, K^+, NH_4^+, and NO_3^- salts are soluble. You can deduce other rules from the context of the question.

EXAMPLE: Write the molecular, complete ionic, and net ionic equations when solutions of potassium chloride and lead(II) nitrate are mixed and a solid product is formed.

SOLUTION:

MOLECULAR EQUATION:

$$2\ KCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\ KNO_3(aq)$$

Lead(II) chloride must be the solid because all nitrates are soluble.

COMPLETE IONIC EQUATION:

$$\cancel{2\,K^+} + 2\,Cl^- + Pb^{2+} + \cancel{2\,NO_3^-} \rightarrow PbCl_2(s) + \cancel{2\,K^+} + \cancel{2\,NO_3^-}$$

NET IONIC EQUATION: $Pb^{2+} + 2\,Cl^- \rightarrow PbCl_2(s)$

Be sure to write the equation in whatever form is requested by the AP exam. When in doubt, write the molecular equation.

ACID–BASE THEORIES: ARRHENIUS AND BRØNSTED–LOWRY

(*Chemistry* 8th ed. pages 639–642/9th ed. pages 653–656)

To work with acid–base reactions, you must be able to identify acids and bases.

The first definition of acids and bases was the Arrhenius theory which states that, in aqueous solution (water), acids produce hydrogen ions and bases produce hydroxide ions.

The Brønsted–Lowry theory says that an acid is a proton (H^+) donor and a base is a proton acceptor. It is more general than the Arrhenius theory, in which only hydroxide bases are considered.

ACID–BASE REACTIONS

(*Chemistry* 8th ed. pages 154–161/9th ed. pages 163–170)

Acid + Base → Salt + H_2O

Memorize the list of strong acids and strong bases below. All others are weak acids or bases.

Strong Acids	Strong Bases
HCl, HBr, HI	Group I Hydroxides, (LiOH, NaOH . . .)
H_2SO_4 (HSO_4^- = weak)	
HNO_3	
$HClO_4$	

STRONG ACID–STRONG BASE REACTIONS

(*Chemistry* 8th ed. page 154/9th ed. page 163)

Strong acids and bases are assumed to be 100% ionized in aqueous solution; write all reactants in ionic form. After combining the proton and the proton acceptor, the remaining cation(s) and anion(s) make up the salt, which in most cases will be soluble.

For the examples that follow, all states are aqueous, unless otherwise indicated.

EXAMPLE: Equimolar solutions of hydrochloric acid and sodium hydroxide are mixed.

SOLUTION: Starting with the complete ionic equation:

$$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O(l)$$

Omit spectator ions to get the net ionic equation:

$$H^+ + OH^- \rightarrow H_2O(l)$$

Note: For most strong acid–strong base reactions, this will be the net ionic equation, but make sure that the salt formed is soluble.

WEAK ACID–STRONG BASE REACTIONS (OR STRONG ACID–WEAK BASE)

(*Chemistry* 8th ed. page 155/9th ed. page 164)

Weak acids and bases ionize only slightly. Write the formula for the weak acid or weak base in molecular form (do not separate into ions!).

EXAMPLE: Write the net ionic equation for the reaction when equimolar solutions of acetic acid and sodium hydroxide are mixed.

SOLUTION: $HC_2H_3O_2(aq) + Na^+ + OH^- \rightleftharpoons Na^+ + C_2H_3O_2^- + H_2O(l)$

Omit spectator ions to get: $HC_2H_3O_2(aq) + OH^- \rightleftharpoons C_2H_3O_2^- + H_2O(l)$

All states are aqueous, unless otherwise indicated.

EXAMPLE: Equimolar solutions of barium hydroxide and sulfuric acid are mixed and a precipitate is formed.

$$Ba(OH)_2 + H_2SO_4 \rightarrow BaSO_4(s) + H_2O(l)$$

Barium hydroxide is only slightly soluble, so it should not be written as separate ions. The first hydrogen ion separates easily from the conjugate base HSO_4^-, but the second does not. Therefore, the net ionic equation for the reaction above is

$$Ba(OH)_2 + H^+ + HSO_4^- \rightarrow BaSO_4(s) + H_2O(l)$$

OXIDATION–REDUCTION REACTIONS

(*Chemistry* 8th ed. pages 161–168/9th ed. pages 170–177)

Oxidation–reduction reactions involve the transfer of electrons from the atom(s) being oxidized to the atom(s) being reduced. Skills required for this section include the assignment of oxidation numbers and the identification of oxidation–reduction reactions. This category includes combustion and some synthesis and decomposition reactions, so check for electron transfer when classifying these kinds of reactions.

In an oxidation–reduction reaction, you need to identify the atoms that are oxidized and the atoms that are reduced. One way to keep track is “OIL RIG”: Oxidation Is Loss of electrons and Reduction Is Gain of electrons.

Rules for Assigning Oxidation States (*Chemistry* 8th ed. page 163/9th ed. page 171)

The oxidation state of . . .	Examples
An atom in element is zero.	Na(*s*), $O_2(g)$
A monatomic ion is the same as its charge.	Na^+
Oxygen is usually –2 in its compounds. Exception: peroxides (containing $O_2{}^{2-}$, ex. Na_2O_2) in which oxygen is –1.	H_2O, CO_2
Hydrogen is +1 in its covalent compounds. (Hydrogen is –1 in binary hydrides, ex. NaH).	H_2O, NH_3
For an electrically neutral compound, the sum of the oxidation states must be zero.	$KMnO_4$ Solution: K = +1, O = –2 +1 + Mn + 4(–2) = 0 –7 + Mn = 0 ; Mn = +7
For an ionic species, the sum of the oxidation states must equal the overall charge.	Ex: $Cr_2O_7{}^{2-}$ Solution: O = –2, 2 Cr + 7(–2) = –2; 2 Cr = +12; Cr = +6

EXAMPLE: Is the following reaction an oxidation–reduction reaction? If it is an oxidation–reduction reaction, identify which atoms are oxidized and which atoms are reduced (*Chemistry* 8th ed. pages 163–165/9th ed. pages 172–174).

$$Zn + Cu(NO_3)_2 \rightarrow Cu + Zn(NO_3)_2$$

SOLUTION: The nitrate ion appears unchanged on both sides of the reaction, so neither element in the ion is oxidized or reduced. *Copper* goes from Cu^{+2} to Cu. (It is reduced, its oxidation number goes down, and it gains 2 electrons.) *Zinc* goes from Zn to Zn^{2+}. (It is oxidized, its oxidation number goes up, and it loses 2 electrons.)

NOTE: If there is no change in oxidation numbers in the reaction, then the reaction is not an oxidation–reduction reaction.

EXAMPLE: Is the following reaction an oxidation–reduction reaction? If it is an oxidation–reduction reaction, identify which atoms are oxidized and which atoms are reduced (*Chemistry* 8th ed. pages 163–165/9th ed. pages 172–174).

$$CaCO_3 \rightarrow CaO + CO_2.$$

SOLUTION: Decomposition is usually oxidation-reduction, but this one is not. *Calcium* is Ca^{2+} on both sides (paired with $CO_3{}^{2-}$ on the left and O^{2-} on the right). *Oxygen's* oxidation number does not change. The oxidation number of *carbon* is +4 in $CO_3{}^{2-}$ and in CO_2.

STEPS FOR BALANCING OXIDATION–REDUCTION REACTIONS

BALANCE OXIDATION–REDUCTION REACTIONS IN ACIDIC SOLUTION USING THE HALF-REACTION METHOD

(*Chemistry* 8th ed. pages 166–168/9th ed. pages 175–177)

EXAMPLE: $Cr_2O_7^{2-} + Cl^- \rightarrow Cr^{3+} + Cl_2$

STEP 1: *Write separate half-reactions*:

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$ $2\ Cl^- \rightarrow Cl_2$

STEP 2: *Balance all atoms except H and O*:

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$ $2\ Cl^- \rightarrow Cl_2$

STEP 3: *Balance oxygen using water*:

$Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ $2\ Cl^- \rightarrow Cl_2$

STEP 4: *Balance hydrogen with* H^+:

$14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ $2\ Cl^- \rightarrow Cl_2$

STEP 5: *Balance charge using electrons:*

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ $2\ Cl^- \rightarrow Cl_2 + 2\ e^-$

STEP 6: *Equalize electron transfer.* Multiply each reaction by numbers that will allow both reactions to have the same number of electrons exchanged:

$6\ e^- + 14\ H^+ + Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+} + 7\ H_2O$ $6\ Cl^- \rightarrow 3\ Cl_2 + 6\ e^-$

Note: 2nd reaction is multiplied by 3.

Add the two half-reactions canceling out all the electrons and the formulas which appear on both sides of the equation.

$14\ H^+ + Cr_2O_7^{2-} + 6\ Cl^- \rightarrow 2\ Cr^{3+} + 7\ H_2O + 3\ Cl_2$

Sum of Charges: $+14 - 2 - 6 = +6 \rightarrow +6 + 0 + 0 = +6$

STEP 7: *Double check that* there is the same number of each kind of atom on both sides and that the sums of all charges are the same on both sides.

BALANCE OXIDATION–REDUCTION REACTIONS IN BASIC SOLUTION USING THE HALF-REACTION METHOD

(*Chemistry* 8th ed. page 168/9th ed. page 177)

Repeat steps 1–5 above. After Step 5: Add OH^- to both sides of the equation (equal to H^+). Form H_2O on the side containing H^+ and OH^- ions. Eliminate number of H_2O appearing on both sides. Continue with steps 6 and 7.

EXAMPLE: $NO_2^- + Al \rightarrow NH_3 + AlO_2^-$

STEP 1: Write separate half-reactions.

$Al \rightarrow AlO_2^-$ $\qquad$ $NO_2^- \rightarrow NH_3$

STEP 2: All atoms except H and O are already balanced.

STEPS 3 AND 4: Balance O with H_2O and H with H^+.

$2\ H_2O + Al \rightarrow AlO_2^- + 4\ H^+$ $\qquad$ $7\ H^+ + NO_2^- \rightarrow NH_3 + 2\ H_2O$

STEP 5: Balance charge using electrons.

$6\ e^- + 7\ H^+ + NO_2^- \rightarrow NH_3 + 2\ H_2O$

$\underline{(2\ H_2O + Al \rightarrow AlO_2^- + 4\ H^+ + 3\ e^-) \times 2}$

$7\ H^+ + NO_2^- + 4\ H_2O + 2\ Al \rightarrow 2\ AlO_2^- + 8\ H^+ + NH_3 + 2\ H_2O$

CANCEL: $NO_2^- + 2\ H_2O + 2\ Al \rightarrow 2\ AlO_2^- + H^+ + NH_3$

New step for basic solution: Add OH^- to both sides equal to the number of H^+ ions:

$OH^- + NO_2^- + 2\ H_2O + 2\ Al \rightarrow 2\ AlO_2^- + H^+ + NH_3 + OH^-$

COMBINE THE PROTON AND HYDROXIDE TO FORM WATER:

$OH^- + NO_2^- + 2\ H_2O + 2\ Al \rightarrow 2\ AlO_2^- + NH_3 + H_2O$

CANCEL: $OH^- + NO_2^- + H_2O + 2\ Al \rightarrow 2\ AlO_2^- + NH_3$

ELECTRON TRANSFER TO OXYGEN (INCLUDING COMBUSTION)

(*Chemistry* 8th ed. pages 101–102/9th ed. pages 107–108)

Combustion is the name given to the complete oxidation of any organic compound containing C, H, and/or O to yield CO_2 and H_2O and heat. Other elements and compounds undergo similar reactions with oxygen, and all are oxidation–reduction reactions.

$C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$

$SiH_4(g) + 2\ O_2(g) \rightarrow SiO_2(s) + 2\ H_2O(g)$

$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

SYNTHESIS/DECOMPOSITION

SYNTHESIS (ALSO KNOWN AS "COMBINATION")

In synthesis, two atoms and/or molecules combine to form a new molecule. These may also be redox reactions if electron transfer occurs.

EXAMPLES:

Two or more elements → 1 compound
$3\ Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

Two or more compounds → 1 compound
$Li_2O(s) + H_2O(l) \rightarrow 2\ Li^+(aq) + 2\ OH^-(aq)$

$SO_3(g) + H_2O(l) \rightarrow H^+(aq) + HSO_4^-(aq)$

$6\ CaO(s) + P_4O_{10}(s) \rightarrow 2\ Ca_3(PO_4)_2(s)$

DECOMPOSITION

Decomposition is the opposite of synthesis, with one reactant forming two or more products. One of the examples below is oxidation-reduction.
1 compound → 2 or more elements or compounds (reverse of combination)

Magnesium carbonate is heated strongly in a crucible.
$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

Nitrogen triiodide decomposes into two gases.
$2\ NI_3(s) \rightarrow 3\ I_2(g) + N_2(g)$

MULTIPLE-CHOICE QUESTIONS

No calculators are to be used in this section.

1. Consider the equation: $Cl_2(g) + 2\ KI(aq) \rightarrow I_2(s) + 2\ KCl(aq)$. Which species is oxidized?
(A) Cl_2
(B) K^+
(C) I^-
(D) I_2

2. Solutions of $CaCl_2$ and Na_2CO_3 are mixed and a precipitate is formed. What is the net ionic equation for the reaction?
(A) $CO_3^{2-}(aq) + Ca^{2+}(aq) \rightarrow CaCO_3(s)$
(B) $2\ Cl^-(aq) + 2\ Na^+(aq) \rightarrow 2\ NaCl(s)$
(C) $CO_3^-(aq) + Ca^{2+}(aq) + 2\ Cl^-(aq) + 2\ Na^+(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$
(D) $CO_3^-(aq) + Ca^{2+}(aq) \rightarrow Ca(CO_3)_2(s)$

3. Which of the following solutions contains the largest number of ions?
(A) 500. mL of 0.100 *M* $FeCl_3$
(B) 700. mL of 0.200 *M* NaOH
(C) 400. mL of 0.100 *M* $Al(NO_3)_3$
(D) 600. mL of 0.200 *M* $AlCl_3$

4. There are six strong acids. Which of the following is <u>not</u> a strong acid?
(A) HCl
(B) HF
(C) HBr
(D) HI

5. Spectator ions are those which are in solution but do not react. Identify any spectator ions for the reaction of sodium phosphate with calcium nitrate.
(A) only $PO_4^{3-}(aq)$
(B) $Na^+(aq)$ and $PO_4^{3-}(aq)$
(C) $Na^+(aq)$ and $NO_3^-(aq)$
(D) $Ca^{2+}(aq)$ and $PO_4^{3-}(aq)$

6. Consider the following three equations for chemical reactions:

   $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$
   $2\ NaCl(aq) + Pb(NO_3)_2(aq) \rightarrow PbCl_2(s) + 2\ NaNO_3(aq)$
   $NaOH(aq) + HBr(aq) \rightarrow H_2O(l) + NaBr(aq)$

   These are examples of:
   (A) three acid–base reactions
   (B) a redox reaction, a precipitation reaction, then an acid–base reaction
   (C) three redox reactions
   (D) a neutralization reaction, then two precipitation reactions

7. Which of the following pairs of ions would not form a solid in aqueous solution?
   (A) Ba^{2+} and SO_4^{2-}
   (B) Pb^{2+} and Br^-
   (C) Na^+ and SO_4^{2-}
   (D) Pb^{2+} and S^{2-}

8. Which of the following ions are likely to form a soluble sulfate in aqueous solution?
   (A) Ba^{2+}
   (B) Pb^{2+}
   (C) Ca^{2+}
   (D) NH_4^+

9. If a solution of sodium bicarbonate is mixed with a solution containing an equal number of moles of nitric acid, then sodium nitrate, water, and carbon dioxide are produced. What is the net ionic equation representing the reaction?
   (A) $NaHCO_3(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$
   (B) $HCO_3^-(aq) + HNO_3(aq) \rightarrow NO_3^-(aq) + H_2O(l) + CO_2(g)$
   (C) $Na^+(aq) + HCO_3^-(aq) + HNO_3(aq) \rightarrow Na^+(aq) + NO_3^-(aq) + H_2O(l) + CO_2(g)$
   (D) $HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

10. Classify each reaction below as one of the following: precipitation, acid-base, or oxidation-reduction (redox).

    Reaction 1 $Ca(OH)_2(aq) + 2\ HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + 2\ H_2O(l)$
    Reaction 2 $Fe_2O_3(s) + 3\ CO(g) \rightarrow 2\ Fe(s) + 3\ CO_2(g)$
    Reaction 3 $CuBr_2(aq) + 2\ NaOH(aq) \rightarrow Cu(OH)_2(s) + 2\ NaBr(aq)$
    Reaction 4 $Cl_2(g) + 2\ NaI(aq) \rightarrow I_2(aq) + 2\ NaCl(aq)$

	Reaction 1	Reaction 2	Reaction 3	Reaction 4
(A)	acid-base	redox	precipitation	redox
(B)	precipitation	redox	acid-base	redox
(C)	redox	precipitation	acid-base	acid-base
(D)	precipitation	precipitation	redox	acid-base

11. The oxidation number of N in $Ca(NO_3)_2$ is
    (A) +2
    (B) +3
    (C) +4
    (D) +5

12. What is the difference between a strong acid and a weak acid?
    (A) A strong acid is more concentrated than a weak acid.
    (B) A weak acid is more soluble in water than a strong acid.
    (C) Strong acids have more hydrogen per molecule than weak acids.
    (D) Strong acids are completely dissociated in solution while weak acids are not.

13. How many moles of electrons are transferred between the substance being oxidized and the substance being reduced in the reaction given below?

$$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(l)$$

    (A) 5
    (B) 10
    (C) 16
    (D) 20

14. Which of the following is a chemical reaction?
    (A) solid carbon dioxide vaporizing
    (B) a seashell dissolving in acid
    (C) ethanol combining with water
    (D) water freezing

15. The half-reaction written correctly is
    (A) $Cl_2 + 2\ e^- \rightarrow Cl_2^-$
    (B) $Na - e^- \rightarrow Na^+$
    (C) $Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$
    (D) $Ca \rightarrow Ca^+ + 2\ e^-$

FREE-RESPONSE QUESTIONS

1. Blood alcohol (C_2H_5OH) level can be determined by titrating a sample of blood plasma with an acidic potassium dichromate solution. The *unbalanced* equation for the reaction is:

$$Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) \rightarrow Cr^{3+}(aq) + CO_2(g)$$

   (a) Identify which species is oxidized and which is reduced.
   (b) Balance the equation, using smallest whole number coefficients.
   (c) How many electrons are transferred in the balanced equation?
   (d) What visible evidence is there that a reaction has occurred?

2. Two solutions are prepared, one of $Cu(NO_3)_2$ and one of KOH.
   (a) Draw molecular representations of the two solutions, assuming that one beaker contains four formula units of $Cu(NO_3)_2$ and the other beaker contains six formula units of KOH.
   (b) Draw a molecular representation of the solution that results when the contents of the beakers are mixed. Include the correct number of formula units or ions, and the correct amounts and kinds of ions remaining.
   (c) Write the balanced net ionic equation for the reaction.

Answers

MULTIPLE-CHOICE QUESTIONS

1. C We need only look at reactants when asked to identify the species oxidized or reduced. KI is an ionic compound and we can observe that the oxidation number of K^+ is the same in both KI and KCl. Whenever an element is found in a compound on one side of the equation and as a free element on the other side of the equation, there has to be a change in oxidation number for that species.

   By definition, oxidation is the loss of electrons. Writing the two half-reactions [$Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$ and $2\ I^- - 2\ e^- \rightarrow I_2$] shows that it is the two iodide ions that lose two electrons to become diatomic iodine (*Chemistry* 8th ed. pages 161–166/9th ed. pages 170–175). LO 3.8

2. A In a precipitation reaction, the cations exchange anions. All compounds of Na^+ are soluble, so the solid must be formed from Ca^{2+} and CO_3^{2-} (*Chemistry* 8th ed. pages 145–150/9th ed. pages 145–147). LO 3.2

3. D There are three factors which must be considered in this problem, the number of ions per formula unit, the concentration of the solution, and the volume of the solution used. For example, in the first solution there are 0.200 moles of ions present: 0.500 L × 0.100 mol of formula units of $FeCl_3$ / L × 4 ions/ $FeCl_3$ = 0.200 moles of ions. In response D: 0.600 L × 0.200 mol of formula units of $AlCl_3$ × 4 ions/ $AlCl_3$ = 0.480 mol. of ions (*Chemistry* 8th ed. pages 136–140/9th ed. pages 145–149). LO 3.4

4. B Even though hydrofluoric acid will dissolve glass (!), it does not ionize significantly. It is important for you to know the six strong acids (then you know that all others are weak!). The six strong acids are $HClO_4$, HCl, HBr, HI, H_2SO_4, and HNO_3. These are the acids that are 100% ionized in water. You can assume that all other acids are weak (*Chemistry* 8th ed. pages 132–134, 642–644, Appendix 5.1/9th ed. pages 141–143, 656–658, Appendix 5.1). LO 3.7

5. C The unbalanced complete ionic equation for this reaction is: $Na^+ + PO_4^{3-} + Ca^{2+} + NO_3^- \rightarrow Ca_3(PO_4)_2 + Na^+ + NO_3^-$. Because you have memorized that all compounds containing sodium cations and nitrate anions are soluble (i.e., $NaNO_3$ is soluble in water), the only possible product is calcium phosphate (*Chemistry* 8th ed. pages 149–151, 154/9th ed. pages 157–160, 162). LO 3.2

6. B Being able to classify reactions will help in predicting products. In the first reaction, Na is oxidized and Cl_2 is reduced (a metal–nonmetal reaction can always be assumed to be a redox reaction); in the second reaction, insoluble lead(II) chloride forms; in the third reaction, a base and an acid form water and a salt (*Chemistry*

8th ed. pages 152–156, 164–166/9th ed. pages 161–165, 173–175). LO 3.2, LO 3.8

7. **C** The answer is reached by process of elimination. All salts containing sodium cation are soluble (*Chemistry* 8th ed. pages 147–148/9th ed. pages 155–156). LO 3.2

8. **D** Salts containing the ammonium ion are soluble (*Chemistry* 8th ed. pages 147–148/9th ed. pages 155–156). LO 3.2

9. **D** Sodium bicarbonate and sodium nitrate are both strong electrolytes (salts) and therefore are written as ions in aqueous solution. Nitric acid is a strong acid and is completely ionized in aqueous solution. Water and carbon dioxide are molecular species and have very little ionization occurring, so they stay written in the molecular form. The two spectator ions, Na^+ and NO_3^- are canceled from both sides of the equation (*Chemistry* 8th ed. pages 148–151/9th ed. pages 157–160). LO 3.2

10. **A** Reaction 1 is acid(HNO_3)-base($Ca(OH)_2$). Reaction 2 is oxidation-reduction; Fe^{3+} in Fe_2O_3 is reduced to Fe. Reaction 3 is precipitation because two aqueous solutions are producing a solid product. In Reaction 4, Cl_2 is reduced to Cl^- and I^- is oxidized to I_2 (*Chemistry* 8th ed. pages 144–150, 154–155, 161–166/9th ed. pages 153–158, 163–164, 170–175). LO 3.2, LO 3.7, LO 3.8

11. **D** Ca ion is always +2 and oxygen is –2. If you remember that the nitrate ion is –1, then 3(–2) + N = –1, and when solving for N, N = +5 (*Chemistry* 8th ed. pages 162–164/9th ed. pages 171–173). LO 3.8

12. **D** This is the definition of a strong electrolyte, which is why strong acids are labeled as "strong" (*Chemistry* 8th ed. pages 132–136/9th ed. pages 141–145). LO 3.7

13. **D** The change in oxidation number of the nitrogen is from –3 in the NH_3 to +2 in the nitrogen monoxide molecule. This requires 5 electrons/nitrogen atom, or 5 moles electrons/mol of nitrogen. Since there are 4 moles of nitrogen atoms, the total moles of electrons needed is 4 × 5 or 20 mol electrons (*Chemistry* 8th ed. pages 166–168/9th ed. pages 175–177). LO 3.8

14. **B** All other examples are phase changes or solution formation (*Chemistry* 8th ed. pages 482–483/9th ed. pages 495–496). LO 3.10

15. **C** The first option is not correct because when Cl_2 gains electrons, it forms two Cl^- anions. The second is not because by convention, when electrons are lost they are included among the products, and when they are gained, are included among the reactants. The last option does not represent the typical charge of a calcium cation (*Chemistry* 8th ed. pages 817–823/9th ed. pages 833–839). LO 3.8

FREE-RESPONSE QUESTIONS

1. Start by assigning oxidation numbers to determine which species are oxidized and which are reduced.

+6 −2	+2 +1 −2	+3	+4 −2	individual O.N.
$Cr_2O_7^{2-}(aq)$ +	$C_2H_6O(aq)$ $\longrightarrow$	$Cr^{3+}(aq)$ +	$CO_2(g)$	
+12 −14	−4 +6 −2	+3	+4 −4	group O.N.

(a) Chromium's oxidation number has gone from +6 to +3, so chromium in dichromate has been reduced. Carbon's oxidation number has gone from +2 to +4, so the carbon in ethanol has been oxidized.

(b) Use the half-reaction method.

STEP 1: *Write separate half-reactions:*

$Cr_2O_7^{2-} \rightarrow Cr^{3+}$ $\qquad$ $C_2H_6O \rightarrow CO_2$

STEP 2: *Balance all atoms except H and O:*

$Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+}$ $\qquad$ $C_2H_6O \rightarrow 2\,CO_2$

STEP 3: *Balance oxygen using water:*

$Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$ $\qquad$ $3\,H_2O + C_2H_6O \rightarrow 2\,CO_2$

STEP 4: *Balance hydrogen with H^+:*

$14\,H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$ $\qquad$ $3\,H_2O + C_2H_6O \rightarrow 2\,CO_2 + 12\,H^+$

STEP 5: *Balance charge using electrons:*

$6\,e^- + 14\,H^+ + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$ $\qquad$ $3\,H_2O + C_2H_6O \rightarrow 2\,CO_2 + 12\,H^+ + 12\,e^-$

STEP 6: *Equalize electron transfer.* Multiply each reaction by numbers that will allow both reactions to have the same number of electrons exchanged:

$12\,e^- + 28\,H^+ + 2\,Cr_2O_7^{2-} \rightarrow 4\,Cr^{3+} + 14\,H_2O$

$3\,H_2O + C_2H_6O \rightarrow 2\,CO_2 + 12\,H^+ + 12\,e^-$

Note: 1st reaction is multiplied by 2.

Add the two half-reactions canceling out all the electrons and the formulas which appear on both sides of the equation.

$$16\,H^+ + 2\,Cr_2O_7^{2-} + C_2H_6O \rightarrow 4\,Cr^{3+} + 11\,H_2O + 2\,CO_2$$

Sum of Charges: $+16 - 4 + 0 = +12 \rightarrow +12 + 0 + 0 = +12$

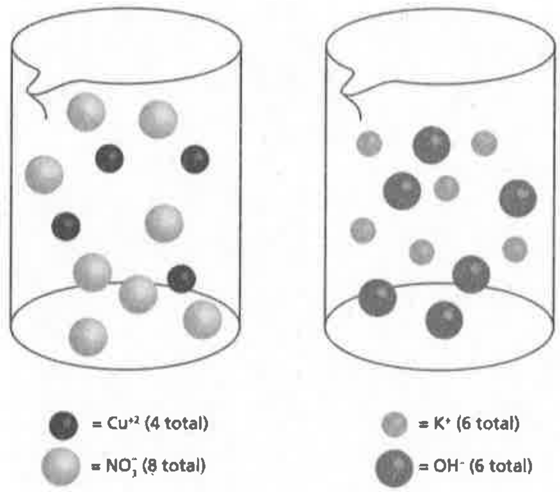
STEP 7: *Double check that* there is the same number of each kind of atom on both sides and that the sums of all charges are the same on both sides.

(c) Twelve electrons are transferred

(*Chemistry* 8th ed. pages 166–168/9th ed. pages 175–177). LO 3.8

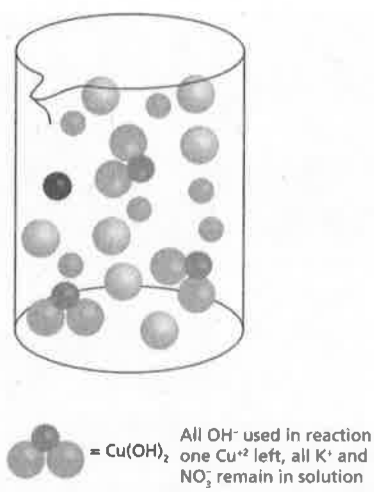
(d) There are two ways to monitor the reaction. First, a gas (CO_2) is formed. Second, the dichromate ion is orange, and the Cr(III) ion is blue-violet, so as long as there is ethanol available for reaction, the solution will be blue-violet. Once all the ethanol is consumed, the solution will stay orange, the color of unreduced dichromate (*Chemistry* 8th ed. pages 820–821/9th ed. pages 836–837). LO 3.10

2. (a)



Both substances are soluble and will ionize completely.

(b)



There are not enough OH^- ions to react with all the Cu^{2+} ions, so there will be one Cu^{2+} ion left in solution, along with the spectators K^+ and NO_3^-.

(c) $Cu^{2+} + 2\ OH^- \rightarrow Cu(OH)_2(s)$ (Chemistry 8th ed. pages 132–136, 145–154/9th ed. pages 141–145, 153–162). LO 3.1