Big Idea #3

Chemical Reactions
Changes in matter involve the rearrangement and/or reorganizations of atoms and/or the transfer of electrons.
LO 3.1: Students can translate among macroscopic observations of change, chemical equations, and particle views.
Types of Chemical Reactions

**Combustion**

\[ C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O \]

**Acid-Base (Neutralization)**

\[ HA + BOH \rightarrow H_2O + BA \]

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**2015 AP® CHEMISTRY FREE-RESPONSE QUESTIONS**

3. Potassium sorbate, \( KC_6H_7O_2 \) (molar mass 150. g/mol) is commonly added to diet soft drinks as a preservative. A stock solution of \( KC_6H_7O_2(aq) \) of known concentration must be prepared. A student titrates 45.00 mL of the stock solution with 1.25 \( M \) HCl(aq) using both an indicator and a pH meter. The value of \( K_a \) for sorbic acid, \( HC_6H_7O_2 \), is \( 1.7 \times 10^{-5} \).

(a) Write the net-ionic equation for the reaction between \( KC_6H_7O_2(aq) \) and HCl(aq).

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Combustion
\[ C_xH_y + O_2 \rightarrow CO_2 + H_2O \]

Acid-Base (Neutralization)
\[ HA + BOH \rightarrow H_2O + BA \]

Oxidation-Reduction
\[ A^+ + e^- \rightarrow A \]
\[ B \rightarrow B + e^- \]

Precipitation
\[ AB (aq) + CD (aq) \rightarrow AD (aq) + CB (s) \]

LO 3.1: Students can translate among macroscopic observations of change, chemical equations, and particle views.
Balanced Equations

**Complete Molecular:** \( AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq) \)

**Complete Ionic:** \( Ag^+(aq) + NO_3^-(aq) + K^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + K^+(aq) + NO_3^-(aq) \)

**Net Ionic:** \( Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s) \)

Spectator ions should not be included in your balanced equations.

Remember, the point of a Net Ionic Reaction is to show only those ions that are involved in the reaction. Chemists are able to substitute reactants containing the same species to create the intended product.

You only need to memorize that compounds with nitrate, ammonium, halides and alkali metals are soluble.

**TABLE 4.2 Solubility Rules for Common Ionic Compounds in Water at 25°C**

<table>
<thead>
<tr>
<th>Soluble Compounds</th>
<th>Insoluble Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Compounds containing alkali metal ions (Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺) and the ammonium ion (NH₄⁺)</td>
<td>Halides of Ag⁺, Hg²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>Nitrates (NO₃⁻), bicarbonates (HCO₃⁻), and chlorates (ClO₃⁻)</td>
<td>Sulfates of Ag⁺, Ca²⁺, Sr²⁺, Ba²⁺, Hg²⁺, and Pb²⁺</td>
</tr>
<tr>
<td>Halides (Cl⁻, Br⁻, I⁻)</td>
<td></td>
</tr>
<tr>
<td>Sulfates (SO₄²⁻)</td>
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<tr>
<th>Insoluble Compounds</th>
<th>Soluble Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbonates (CO₃²⁻), phosphates (PO₄³⁻), chromates (CrO₄²⁻), sulfides (S²⁻)</td>
<td>Compounds containing alkali metal ions and the ammonium ion</td>
</tr>
<tr>
<td>Hydroxides (OH⁻)</td>
<td>Compounds containing alkali metal ions and the Ba²⁺ ion</td>
</tr>
</tbody>
</table>

**LO 3.2:** The student can translate an observed chemical change into a balanced chemical equation and justify the choice of equation type (molecular, ionic, or net ionic) in terms of utility for the given circumstances.
LO 3.3: The student is able to use stoichiometric calculations to predict the results of performing a reaction in the laboratory and/or to analyze deviations from the expected results.

2) a. Solid copper carbonate is heated strongly:

b. What evidence of a chemical change would be observed with this reaction?

c. What is the percent yield of CO$_2$ if you had originally heated 10.0g CuCO$_3$ and captured 3.2g CO$_2$?

d. How could you improve your percent yield?

Click reveals answer and explanation.

Click reveals answer and explanation.

Click reveals answer and explanation.

Click reveals answer and explanation.
Making Predictions

2) 
   a. Solid copper carbonate is heated strongly:

   \[ \text{CuCO}_3 \ (s) \rightarrow \text{CuO} \ (s) + \text{CO}_2 \ (g) \]

   b. What evidence of a chemical change would be observed with this reaction?

   One would observe a color change and evolution of a gas.

   c. What is the percent yield of \( \text{CO}_2 \) if you had originally heated 10.0g \( \text{CuCO}_3 \) and captured 3.2g \( \text{CO}_2 \)?

   Step 1: Find the Theoretical Yield
   \[
   10.0g \text{CuCO}_3 \times \left( \frac{1 \text{mol}}{123.555 \text{g}} \right) \times \left( \frac{1 \text{mol CO}_2}{1 \text{mol CuCO}_3} \right) \times \frac{44.01 \text{g CO}_2}{\text{mol}} = 3.562 \text{ g CO}_2
   \]

   Step 2: Find Percent Yield
   \[
   \left( \frac{3.2 \text{ g}}{3.562 \text{ g}} \right) \times 100 = 89.8\% \rightarrow 90\% \text{ with correct sig figs}
   \]

   d. How could you improve your percent yield?
   - reheat the solid, to see if there is any further mass loss
   - make sure you have pure \( \text{CuCO}_3 \)

LO 3.3: The student is able to use stoichiometric calculations to predict the results of performing a reaction in the laboratory and/or to analyze deviations from the expected results.
Limiting Reactants – D.A.

3)

\[ \text{Al}_2\text{S}_3 + 6 \text{H}_2\text{O} \rightarrow 2\text{Al(OH)}_3 + 3 \text{H}_2\text{S} \]

15.00 g aluminum sulfide and 10.00 g water react

a) Identify the Limiting Reactant

Click reveals answer and explanation.

b) What is the maximum mass of \( \text{H}_2\text{S} \) which can be formed from these reagents?

Click reveals answer and explanation.

c) How much excess reactant is left in the container?

Click reveals answer and explanation.

***Dimensional Analysis is not the only way to solve these problems. You can also use BCA tables (modified ICE charts), which may save time on the exam.***

LO 3.4: The student is able to relate quantities (measured mass of substances, volumes of solutions, or volumes and pressures of gases) to identify stoichiometric relationships for a reaction, including situations involving limiting reactants and situations in which the reaction has not gone to completion.
Limiting Reactants – D.A.

3) \( \text{Al}_2\text{S}_3 + 6 \text{H}_2\text{O} \rightarrow 2\text{Al(OH)}_3 + 3 \text{H}_2\text{S} \)

15.00 g aluminum sulfide and 10.00 g water react

a) Identify the Limiting Reactant

\[
15.00\text{g Al}_2\text{S}_3 \times \frac{1\text{mol}}{150.158\text{g}} \times \frac{6\text{mol H}_2\text{O}}{1\text{mol Al}_2\text{S}_3} \times \frac{18\text{g/mol H}_2\text{O}}{1\text{mol H}_2\text{O}} = 10.782\text{ g H}_2\text{O needed}
\]

\[
10\text{g H}_2\text{O} \times \frac{1\text{mol}}{18.015\text{g}} \times \frac{1\text{mol Al}_2\text{S}_3}{6\text{mol H}_2\text{O}} \times \frac{150.158\text{g/mol}}{1\text{mol Al}_2\text{S}_3} = 13.892\text{g Al}_2\text{S}_3 \text{ needed}
\]

H\(_2\)O is limiting, because we need more than we were given

b) What is the maximum mass of H\(_2\)S which can be formed from these reagents?

Theoretical Yield

\[
10.00\text{g H}_2\text{O} \times \frac{1\text{mol}}{18.015\text{g}} \times \frac{3}{6} \times \frac{34.0809\text{g/mol}}{34.0809\text{g/mol}} = 9.459\text{g H}_2\text{S produced}
\]

c) How much excess reactant is left in the container?

\[
15.00\text{g} - 13.892\text{g} = 1.11\text{g Al}_2\text{S}_3
\]

**Dimensional Analysis is not the only way to solve these problems. You can also use BCA tables (modified ICE charts), which may save time on the exam**

LO 3.4: The student is able to relate quantities (measured mass of substances, volumes of solutions, or volumes and pressures of gases) to identify stoichiometric relationships for a reaction, including situations involving limiting reactants and situations in which the reaction has not gone to completion.
15.00 g aluminum sulfide and 10.00 g water react according to the following equation:

\[ \text{Al}_2\text{S}_3 + 6 \text{H}_2\text{O} \rightarrow 2\text{Al(OH)}_3 + 3 \text{H}_2\text{S} \]

### a) Identify the Limiting Reactant

\[
\begin{array}{c|c|c}
\text{mass} & \text{molar mass} & \text{mol} \\
15.00 \text{g} & 150.158 \text{g/mol} & 0.100 \text{mol} \\
10.00 \text{g} & 18.015 \text{g/mol} & 0.555 \text{mol} \\
\end{array}
\]

Water is the limiting reactant.

### b) What is the maximum mass of \( \text{H}_2\text{S} \) which can be formed from these reagents?

\[
0.2775 \text{ mol} \times 34.0809 \text{ g/mol} = 9.459 \text{ g} \]

### c) How much excess reactant is left in the container?

Before:
- \(0.0999\) mol \(\text{Al}_2\text{S}_3\)
- \(0.5551\) mol \(\text{H}_2\text{O}\)

Change:
- \(-0.0925\) mol \(\text{Al}_2\text{S}_3\)
- \(-0.5551\) mol \(\text{H}_2\text{O}\)

After:
- \(0.0074\) mol \(\text{Al}_2\text{S}_3\)
- \(0\) mol \(\text{H}_2\text{O}\)
- \(+0.1850\) mol \(\text{Al(OH)}_3\)
- \(+0.2775\) mol \(\text{H}_2\text{S}\)

LO 3.4: The student is able to relate quantities (measured mass of substances, volumes of solutions, or volumes and pressures of gases) to identify stoichiometric relationships for a reaction, including situations involving limiting reactants and situations in which the reaction has not gone to completion.
Limiting Reactants – BCA Table

3) Alternative technique
15.00 g aluminum sulfide and 10.00 g water react according to the following equation:

\[ \text{Al}_2\text{S}_3 + 6 \text{H}_2\text{O} \rightarrow 2\text{Al(OH)}_3 + 3 \text{H}_2\text{S} \]

a) Identify the Limiting Reactant

\[ 15.00 \text{g Al}_2\text{S}_3 \times (1 \text{mol/150.158 g}) = 0.100 \text{mol} \]
\[ 10 \text{g H}_2\text{O} \times (1 \text{mol/18.015 g}) = 0.555 \]

Water is the limiting reactant.

<table>
<thead>
<tr>
<th></th>
<th>Al(_2)S(_3)</th>
<th>6 H(_2)O</th>
<th>2Al(OH)(_3)</th>
<th>3 H(_2)S</th>
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b) What is the maximum mass of H\(_2\)S which can be formed from these reagents?

\[ 0.2775 \text{ mol H}_2\text{S} \times (34.0809 \text{ g/mol}) = 9.459 \text{ g H}_2\text{S} \text{ produced} \]

c) How much excess reactant is left in the container?

\[ 0.0074 \text{ mol Al}_2\text{S}_3 \times 150.158 \text{ g/mol} = 1.11 \text{ g Al}_2\text{S}_3 \]

LO 3.4: The student is able to relate quantities (measured mass of substances, volumes of solutions, or volumes and pressures of gases) to identify stoichiometric relationships for a reaction, including situations involving limiting reactants and situations in which the reaction has not gone to completion.
Experimental Design

4) Synthesis

a. A sample of pure Cu is heated in excess pure oxygen. Design an experiment to determine quantitatively whether the product is CuO or Cu$_2$O.

Find the mass of the copper. Heat in oxygen to a constant new mass. Subtract to find the mass of oxygen that combined with the copper. Compare the moles of oxygen atoms to the moles of original copper atoms to determine the formula.

Decomposition

CaCO$_3$(s) $\rightarrow$ CaO(s) + CO$_2$(g)

b. Design a plan to prove experimentally that this reaction illustrates conservation of mass.

LO3.5: The student is able to design a plan in order to collect data on the synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.
Experimental Design

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a. A sample of pure Cu is heated in excess pure oxygen. Design an experiment to determine quantitatively whether the product is CuO or Cu₂O.

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Decomposition \[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]

b. Design a plan to prove experimentally that this reaction illustrates conservation of mass.

Find the mass of calcium carbonate and seal it in a rigid container. Evacuate the container of remaining gas. Heat the container and take pressure readings (this will be the pressure exerted by the CO₂). Using PV=nRT, calculate the moles of carbon dioxide gas present in the container and compare it to the molar relationships afforded by the balanced chemical equation.

LO3.5: The student is able to design a plan in order to collect data on the synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.
**Data Analysis**

1) Determine the number of moles of tin. \( \frac{5.200}{118.7} = 0.0438 \text{ moles.} \)

2) Subtract the mass of the crucible from the mass after the third heating. \( 25.252 - 18.650 = 6.602 \text{ g SnO}_x \)

3) Subtract the mass of tin from the mass of oxide to get the mass of oxygen. \( 6.602 - 5.200 = 1.402 \text{ grams of oxygen.} \)

4) Calculate the moles of oxygen atoms, and divide by the moles of tin atoms to get the formula ratio. \( \frac{1.402 \text{ g}}{16.00 \text{ g/mol of atoms}} = 0.0876 \text{ moles.} \frac{0.0876}{0.0438} = 2.00 \)

The formula must be \( \text{SnO}_2 \).

**Video**

When tin is treated with concentrated nitric acid, and the resulting mixture is strongly heated, the only remaining product is an oxide of tin. A student wishes to find out whether it is \( \text{SnO} \) or \( \text{SnO}_2 \).

Mass of pure tin \ 5.200\text{ grams.}\)
Mass of dry crucible \ 18.650\text{ g}\)
Mass of crucible + oxide after first heating \ 25.500\text{ g}\)
Mass after second heating \ 25.253\text{ g}\)
Mass after third heating \ 25.252\text{ g}\)

**How can you use this data, and the law of conservation of mass, to determine the formula of the product?**

Click reveals answer and explanation.

LO 3.6: The student is able to use data from synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.
**Data Analysis**

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Mass after second heating: 25.253 g
Mass after third heating: 25.252 g

**How can you use this data, and the law of conservation of mass, to determine the formula of the product?**

1) Determine the number of moles of tin. 5.200/118.7 = 0.0438 moles. Sn

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1.402 g/16.00 g/mol of atoms = 0.0876 moles. 0.0876/0.0438 = 2.00

The formula must be SnO₂.

LO 3.6: The student is able to use data from synthesis or decomposition of a compound to confirm the conservation of matter and the law of definite proportions.
Bronsted-Lowery Acids & Bases

According to Bronsted-Lowery (B.L.) an acid is a "proton donor" and a base is a "proton acceptor." The proton here is shown as a hydrogen.

The acid's conjugate base is the anion.
The base's conjugate acid now has the proton (hydrogen ion).

- Hydrogen fluoride: A Brønsted-Lowry acid
  \[
  \text{HF(aq) + H}_2\text{O(l) }\rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{F}^-(aq)
  \]

- Ammonia: A Brønsted-Lowry base
  \[
  \text{NH}_3\text{(aq) + H}_2\text{O(l) }\rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)
  \]

Amphoteric nature of water
Water acts as both an acid & a base.

LO 3.7: The student is able to identify compounds as Bronsted-Lowry acids, bases and/or conjugate acid-base pairs, using proton-transfer reactions to justify the identification.
According to Bronsted-Lowery (B.L.) an acid is a “proton donor” and a base is a “proton acceptor.” The proton here is shown as a hydrogen.

It is an acid-base reaction. The weak acid HC$_2$H$_5$O$_2$ reacts with the weak base HCO$_3^-$ with HC$_2$H$_5$O$_2$ donating a proton.

OR

It is an acid-base reaction. No solid precipitates, so it is not a precipitation reaction. None of the oxidation numbers change, so it is not a redox reaction.

1 point is earned for identifying the reaction as acid-base.

1 point is earned for the justification.

LO 3.7: The student is able to identify compounds as Bronsted-Lowry acids, bases and/or conjugate acid-base pairs, using proton-transfer reactions to justify the identification.
According to Bronsted-Lowery (B.L.) an acid is a “proton donor” and a base is a “proton acceptor.” The proton here is shown as a hydrogen.

Lo 3.7: The student is able to identify compounds as Bronsted-Lowry acids, bases and/or conjugate acid-base pairs, using proton-transfer reactions to justify the identification.
Redox Reactions

- When an electron is transferred, it is called a **redox reaction**. When something is reduced, the RED part of redox, it gains electrons. You may have a difficult time with this definition because when something is reduced, it usually means that it is losing something. In this case, it is a reduction in charge. Remember, electrons are negatively charged so if something is being reduced, it's getting more negatively charged by receiving more electrons. The other reaction that is coupled with this is called **oxidation**—the "OX" part of redox. Whenever something is reduced, the electron it gains has to come from somewhere. The oxidation is the loss of an electron, so if an atom is oxidized it loses its electron to another atom. And these are always coupled reactions. If one molecule is oxidized, another molecule must be reduced and vice versa: the electron must go somewhere.

**Half Equations**

- Redox reactions involve the transfer of electrons.
- Equations written to show what happens to the electrons during oxidation and reduction are called half-equations.

\[ \text{magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide} \]

\[ 2\text{Mg(s)} + \text{O}_2(g) \rightarrow 2\text{MgO(s)} \]

**oxidation:** \( \text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^- \)

**reduction:** \( \text{O}_2 + 4\text{e}^- \rightarrow 2\text{O}^2- \)

LO 3.8: The student is able to identify redox reactions and justify the identification in terms of electron transfer.
Redox Reactions

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LO 3.8: The student is able to identify redox reactions and justify the identification in terms of electron transfer

Question:
Zinc ions will react with aluminum metal according to the following chemical reaction:

$$2 \text{Al(s)} + 3 \text{Zn}^{2+}(aq) \rightarrow 2\text{Al}^{3+}(aq) + 3 \text{Zn(s)}$$

Based on this chemical reaction how many moles of electrons would be transferred when 1.0 mol of Zn^{2+} ions are consumed?

a. 6.0 moles  
b. 3.0 moles  
c. 2.0 moles  
d. 1.0 moles  
e. 0.33 moles

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LO 3.8: The student is able to identify redox reactions and justify the identification in terms of electron transfer
Redox Titrations

A redox titration (also called an oxidation-reduction titration) can accurately determine the concentration of an unknown analyte by measuring it against a standardized titrant. A common example is the redox titration of a standardized solution of potassium permanganate (KMnO₄) against an analyte containing an unknown concentration of iron (II) ions (Fe²⁺). The balanced reaction in acidic solution is as follows:

\[
\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow 5\text{Fe}^{3+} + \text{Mn}^{2+} + 4\text{H}_2\text{O}
\]

In this case, the use of KMnO₄ as a titrant is particularly useful, because it can act as its own indicator; this is due to the fact that the KMnO₄ solution is bright purple, while the Fe²⁺ solution is colorless. It is therefore possible to see when the titration has reached its endpoint, because the solution will remain slightly purple from the unreacted KMnO₄.

LO 3.9: The student is able to design and/or interpret the results of an experiment involving a redox titration.
Evidence of Chemical Change

LO 3.10: Evaluate the classification of a process as a physical, chemical, or ambiguous change based on both macroscopic observations and the distinction between rearrangement of covalent interactions and noncovalent interactions.

Chemical Changes:
Production of a gas:
\[ 2 \text{KClO}_3 (s) + \text{heat} \rightarrow 2\text{KCl} (s) + 3\text{O}_2 (g) \]

Formation of a precipitate:
\[ \text{AgNO}_3 (aq) + \text{KCl} (aq) \rightarrow \text{AgCl} (s) + 2\text{KNO}_3 (aq) \]

Change in color:
Two white solids react to produce a mixture of a yellow and a white solid when shaken forcefully!
\[ \text{Pb(NO}_3)_2 (s) + 2\text{KI} (s) \rightarrow \text{PbI}_2 (s) + 2\text{KNO}_3 (s) \]

Production of heat*:
\[ 2 \text{Mg} (s) + \text{O}_2 (s) \rightarrow 2\text{MgO} (s) + \text{heat} \]
*can also include the absorption of heat

Physical Changes:
may produce similar visible evidence (i.e. boiling water creates “bubbles,” but bonds are not broken and reformed. No new substances are made.

An exothermic chemical reaction rearranges electrons and nuclei to a lower potential energy configuration; the excess energy is liberated as heat.
Evidence of Chemical Change

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**Chemical Changes:**

- Production of a gas:
  \[2\text{KClO}_3 (s) + \text{heat} \rightarrow 2\text{KCl}(s) + 3\text{O}_2 (g)\]

- Formation of a precipitate:
  \[\text{AgNO}_3 (aq) + \text{KCl} (aq) \rightarrow \text{AgCl} (s) + 2\text{KNO}_3 (aq)\]

- Change in color:
  Two white solids react to produce a mixture of a yellow and a white solid when shaken forcefully!
  \[\text{Pb(NO}_3)_2 (s) + 2\text{KI} (s) \rightarrow \text{PbI}_2 (s) + 2\text{KNO}_3 (s)\]

- Production of heat*:
  \[2\text{Mg} (s) + \text{O}_2 (s) \rightarrow 2\text{MgO} (s) + \text{heat}\]

*can also include the absorption of heat

**Physical Changes:**

- May produce similar visible evidence (i.e. boiling water creates “bubbles,” but bonds are not broken and reformed. No new substances are made.

---

Note: it is a common misconception that boiling water makes \(\text{O}_2\) and \(\text{H}_2\) gas.

Notice that the water molecule stays intact as the water boils. Covalent bonds are not broken with during this phase change- only intermolecular attractions (hydrogen bonds) between water molecules.
Energy Changes

- Chemical reactions involve the formation of new products
- Bonds between atoms or ions in the reactants must be BROKEN (the enthalpy of the system is increasing ... ENDOThERMIC process)
- Bonds are then FORMED between atoms or ions to make the products of the reaction. (the enthalpy of the system is decreasing...EXOTHERMIC process)

LO 3.11: The student is able to interpret observations regarding macroscopic energy changes associated with a reaction or process to generate a relevant symbolic and/or graphical representation of the energy changes.
Question:
The following question is based on combining the three different half-cell reactions listed below:

Half Cell 1: \( \text{Sn}^{2+} + 2e^- \rightarrow \text{Sn} \)
Half Cell 2: \( \text{Ag}^+ + e^- \rightarrow \text{Ag} \)
Half Cell 3: \( \text{Cr}^{3+} + 3e^- \rightarrow \text{Cr} \)

<table>
<thead>
<tr>
<th>Galvanic Cell</th>
<th>Half Cells</th>
<th>Reaction</th>
<th>( E_{\text{cell}}^{\circ} ) (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>1 &amp; 2</td>
<td>( \text{Sn} + 2\text{Ag}^+ \rightarrow 2\text{Ag} + \text{Sn}^{2+} )</td>
<td>0.94</td>
</tr>
<tr>
<td>Y</td>
<td>2 &amp; 3</td>
<td>( \text{Cr} + 3\text{Ag}^+ \rightarrow 3\text{Ag} + \text{Cr}^{3+} )</td>
<td>1.54</td>
</tr>
<tr>
<td>Z</td>
<td>1 &amp; 3</td>
<td>( 2\text{Cr} + 3\text{Sn}^{2+} \rightarrow 3\text{Sn} + 2\text{Cr}^{3+} )</td>
<td>?</td>
</tr>
</tbody>
</table>

What is the cell potential of galvanic cell Z?

a. 0.25 V  
b. 0.60 V  
c. 2.48 V  
d. 5.90 V

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Click reveals answer and explanation.

LO 3.12: Make qualitative or quantitative predictions about galvanic or electrolytic reactions based on half-cell reactions and potentials and/or Faraday’s laws.
**Question:**
The following question is based on combining the three different half-cell reactions listed below:

- **Half Cell 1:** \( \text{Sn}^{2+} + 2e^- \rightarrow \text{Sn} \)
- **Half Cell 2:** \( \text{Ag}^+ + e^- \rightarrow \text{Ag} \)
- **Half Cell 3:** \( \text{Cr}^{3+} + 3e^- \rightarrow \text{Cr} \)

<table>
<thead>
<tr>
<th>Galvanic Cell</th>
<th>Half Cells</th>
<th>Reaction</th>
<th>( E_{\text{cell}} ) (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>1 &amp; 2</td>
<td>( \text{Sn} + 2\text{Ag}^+ \rightarrow 2\text{Ag} + \text{Sn}^{2+} )</td>
<td>0.94</td>
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<td>Y</td>
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<td>1.54</td>
</tr>
<tr>
<td>Z</td>
<td>1 &amp; 3</td>
<td>( 2\text{Cr} + 3\text{Sn}^{2+} \rightarrow 3\text{Sn} + 2\text{Cr}^{3+} )</td>
<td>?</td>
</tr>
</tbody>
</table>

What is the cell potential of galvanic cell Z?

- a. 0.26 V
- b. 0.60 V
- c. 2.48 V
- d. 5.90 V

**Answer:**
The correct answer is “b”, 0.60 V. The potential of cell Z can be calculated by combining cells X and Y. Cell X needs to be reversed changing its potential from 0.94 to -0.94, then the reactions of the reversed X and Y will combine to give cell Z. So the value -0.94 from the reversed cell X can be added to the potential of cell Y giving a value of 0.60 for cell Z. It is important to note that even though cell X would need to be multiplied by 3 and cell Y would need to be multiplied by 2 in order to produce cell Z those changes do not effect the voltage of either cell.

---

**LO 3.12:** Make qualitative or quantitative predictions about galvanic or electrolytic reactions based on half-cell reactions and potentials and/or Faraday’s laws.
Redox Reactions and Half Cells

**Question:**
The following question is based on combining the three different half cells listed below:

- Half Cell 1: $\text{Sn}^{2+} + 2e^- \rightarrow \text{Sn}$
- Half Cell 2: $\text{Ag}^+ + e^- \rightarrow \text{Ag}$
- Half Cell 2: $\text{Cr}^{3+} + 3e^- \rightarrow \text{Cr}$

<table>
<thead>
<tr>
<th>Galvanic Cell</th>
<th>Half Cells</th>
<th>Reaction</th>
<th>$E^{\circ}_{\text{cell}}$ (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>1 &amp; 2</td>
<td>$\text{Sn} + 2\text{Ag}^+ \rightarrow 2\text{Ag} + \text{Sn}^{2+}$</td>
<td>0.94</td>
</tr>
<tr>
<td>Y</td>
<td>2 &amp; 3</td>
<td>$\text{Cr} + 3\text{Ag}^+ \rightarrow 3\text{Ag} + \text{Cr}^{3+}$</td>
<td>1.54</td>
</tr>
<tr>
<td>Z</td>
<td>1 &amp; 3</td>
<td>$2\text{Cr} + 3\text{Sn}^{2+} \rightarrow 3\text{Sn} + 2\text{Cr}^{3+}$</td>
<td>?</td>
</tr>
</tbody>
</table>

In galvanic cells X and Z, which of the following takes place in half cell 1?

- a. Oxidation occurs in both cell X and cell Z.
- b. Reduction occurs in both cell X and cell Z.
- c. Oxidation occurs in cell X and reduction in cell Z.
- d. Reduction occurs in cell X and oxidation in cell Z.

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**LO 3.13:** The student can analyze data regarding galvanic or electrolytic cells to identify properties of the underlying redox reactions.
Redox Reactions and Half Cells

LO 3.13: The student can analyze data regarding galvanic or electrolytic cells to identify properties of the underlying redox reactions.
While cleaning up after the experiment, the student wishes to dispose of the unused solid I₂ in a responsible manner. The student decides to convert the solid I₂ to I⁻(aq) anion. The student has access to three solutions, H₂O₂(aq), Na₂S₂O₃(aq), and Na₂S₄O₆(aq), and the standard reduction table shown below.

<table>
<thead>
<tr>
<th>Half reaction</th>
<th>E° (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>S₄O₆²⁻(aq) + 2 e⁻ → 2 S₂O₃²⁻(aq)</td>
<td>0.08</td>
</tr>
<tr>
<td>I₂(s) + 2 e⁻ → 2 I⁻(aq)</td>
<td>0.54</td>
</tr>
<tr>
<td>O₂(g) + 2 H⁺(aq) + 2 e⁻ → H₂O₂(aq)</td>
<td>0.68</td>
</tr>
</tbody>
</table>

(e) Which solution should the student add to I₂(s) to reduce it to I⁻(aq)? Circle your answer below. Justify your answer, including a calculation of E° for the overall reaction.

H₂O₂(aq)       Na₂S₂O₃(aq)       Na₂S₄O₆(aq)

(f) Write the balanced net-ionic equation for the reaction between I₂ and the solution you selected in part (e).
(e) Which solution should the student add to I₂(s) to reduce it to I⁻(aq)? Circle your answer below. Justify your answer and include a calculation of $E^\circ$ for the overall reaction.

<table>
<thead>
<tr>
<th>H₂O₂(aq)</th>
<th>Na₂S₂O₃(aq)</th>
<th>Na₂S₄O₆(aq)</th>
</tr>
</thead>
</table>

[Na₂S₂O₃(aq) should be circled.]

The reaction between S₂O₃²⁻(aq) and I₂(s) will be thermodynamically favorable because $E^\circ$ for the reaction is positive ($E^\circ = 0.54 - 0.08 = +0.46$ V), from which it follows that $\Delta G^\circ$ is negative because $\Delta G^\circ = -nFE^\circ$.

1 point is earned for the correct choice.

1 point is earned for a correct justification.

(f) Write the balanced net-ionic equation for the reaction between I₂ and the solution you selected in part (e).

$$I_2 + 2 \text{S}_2\text{O}_3^{2-} \rightarrow 2I^- + \text{S}_4\text{O}_6^{2-}$$

1 point is earned for the correct equation.
3. To determine the molar mass of an unknown metal, M, a student reacts iodine with an excess of the metal to form the water-soluble compound $\text{MI}_2$, as represented by the equation above. The reaction proceeds until all of the $\text{I}_2$ is consumed. The $\text{MI}_2(aq)$ solution is quantitatively collected and heated to remove the water, and the product is dried and weighed to constant mass. The experimental steps are represented below, followed by a data table.

![Diagram](image)

### Data for Unknown Metal Lab

<p>| | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of beaker</td>
<td>125.457 g</td>
</tr>
<tr>
<td>Mass of beaker + metal M</td>
<td>126.549 g</td>
</tr>
<tr>
<td>Mass of beaker + metal M + $\text{I}_2$</td>
<td>127.570 g</td>
</tr>
<tr>
<td>Mass of $\text{MI}_2$, first weighing</td>
<td>1.284 g</td>
</tr>
<tr>
<td>Mass of $\text{MI}_2$, second weighing</td>
<td>1.284 g</td>
</tr>
</tbody>
</table>

(a) Given that the metal M is in excess, calculate the number of moles of $\text{I}_2$ that reacted.

(b) Calculate the molar mass of the unknown metal M.
3. To determine the molar mass of an unknown metal, M, a student reacts iodine with an excess of the metal to form the water-soluble compound MI₂, as represented by the equation above. The reaction proceeds until all of the I₂ is consumed. The MI₂(aq) solution is quantitatively collected and heated to remove the water, and the product is dried and weighed to constant mass. The experimental steps are represented below, followed by a

(a) Given that the metal M is in excess, calculate the number of moles of I₂ that reacted.

\[
\begin{align*}
127.570 - 126.549 &= 1.021 \text{ g I}_2 \\
1.021 \text{ g I}_2 \times \frac{1 \text{ mol I}_2}{253.80 \text{ g I}_2} &= 0.004023 \text{ mol I}_2
\end{align*}
\]

1 point is earned for the number of moles.

(b) Calculate the molar mass of the unknown metal M.

\[
\begin{align*}
\text{Number of moles of I}_2 &= \text{number of moles of M} \\
1.284 \text{ g MI}_2 - 1.021 \text{ g I}_2 &= 0.263 \text{ g M} \\
\text{Molar mass of M} &= \frac{0.263 \text{ g M}}{0.004023 \text{ mol M}} = 65.4 \text{ g/mol}
\end{align*}
\]

1 point is earned for the number of grams of M.  
1 point is earned for the molar mass.