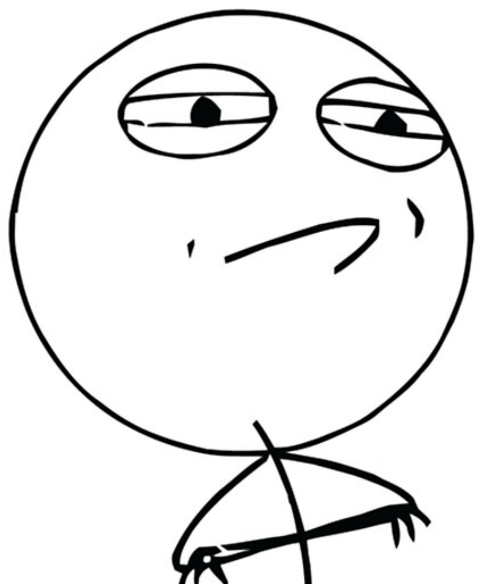
Name: ____ Period: _____

AP Chemistry

Challenge Accepted



Unit 1:

Chemistry "Fun"damentals

Unit 2:

Stoichiometry & Reactions

Unit 3:

Electrochemistry

<u>Table of Contents</u>: Units 1 & 2

Unit 3: Table of Contents

Class Info

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Remind Info:

- A Day: Text **@APLestikA** to **81010**
- B Day: Text **@APLestikB** to **81010**

1st Marking Period: August-October 2023

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
13 August	**14**	**15 A** First day of school	**16 B**	**17 A**	**18 B**	19
20	**21 A**	**22 B**	**23 A**	**24 B**	**25 A**	26
27	**28 B**	**29 A**	**30 B**	**31 A**	**1 Sept. B**	2
3	**4** Labor Day- No school	**5 A**	**6 B**	**7 A**	**8 B**	9
10	**11 A**	**12 B**	**13 A**	**14 B**	**15 A**	16
17	**18 B**	**19 A**	**20 B**	**21 A**	**22 B**	23

1

Sunday	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday
24 Sept.	**25** Student Holiday/ Staff Development	**26** **A**	**27** **B**	**28** **A**	**29** **B**	30
1 Oct.	**2** **A**	**3** **B**	**4** **A**	**5** **B**	**6** **A**	7
8	**9** Student Holiday/ Staff Development	**10** **B**	**11** **A**	**12** **B**	**13** **A** End of 1st MP	14
15	**16** **B**	**17** **A**	**18** **B**	**19** **A**	**20** **B**	21
22	**23** **A**	**24** **B**	**25** **A**	**26** **B**	**27** **A**	28
29	**30** **B**	**31** **A**	**1 Nov.** **B**	**2** **A**	**3** **B**	4

AP Chemistry Unit 1 Objectives

BIG IDEA 1 - The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

- Enduring Understanding 1.A: All matter is made of atoms. There are a limited number of types of atoms; theses are elements.
- Enduring Understanding 1.C: Elements display periodicity in their properties when the elements are organized according to increasing atomic number. This periodicity can be explained by the regular variations that occur in the electronic structures of atoms. Periodicity is a useful principle for understanding properties and predicting trends in periods. Its modern-day uses range from examining the composition of materials to generating ideas for designing new materials.
- Enduring Understanding 1.D: Atoms are so small that they are difficult to study directly; atomic models are constructed to explain experimental data on collections of atoms.
- Enduring Understanding 1.E: Atoms are conserved in physical and chemical processes.

BIG IDEA 2 - Chemical and physical properties of materials can be explained by the structure and arrangement of atoms, ions, molecules and the forces between them.

- Enduring Understanding 2.A: Matter can be described by its physical properties. The physical properties of a substance generally depend on the spacing between the particles (atoms, molecules, ions) that make up the substance and the forces of attraction among them.
- Enduring Understanding 2.C: The strong electrostatic forces of attraction holding atoms together in a unit are called chemical bonds.

Unit 1: Fundamental Vocab

Types of Particles

Atom – a neutral, monatomic element

Compound – neutral particle; two or more ___________________ atoms bonded together

Molecule – particle of a covalent compound; two or more atoms ___________________ bonded together

Diatomic elements – **Br I N Cl H O F**; these elements naturally exist in covalently bonded pairs

Formula Unit – particle of an ________________ compound; cation and anions bonded together in a neutral compound

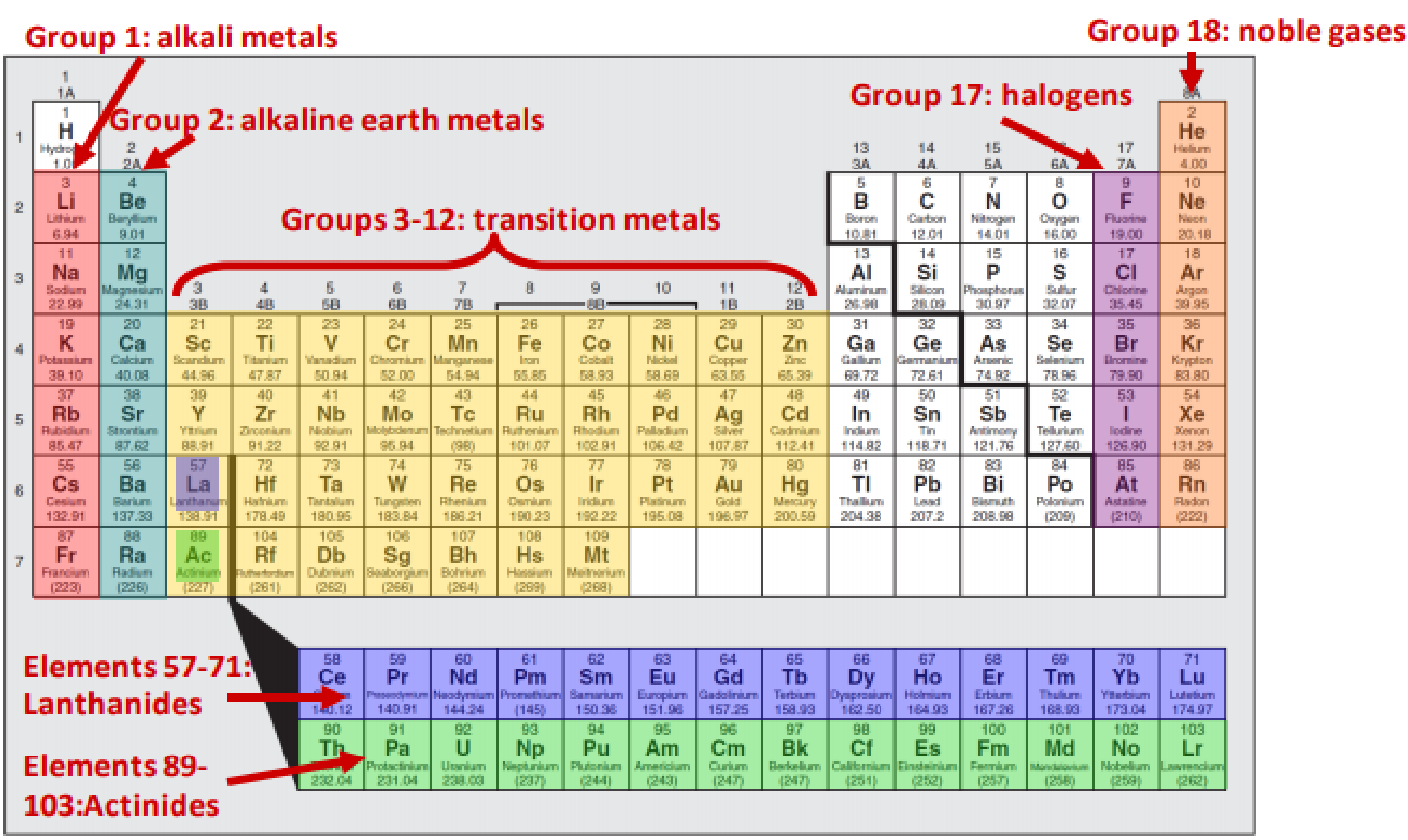
Ion – an atom, or group of atoms, that has a _______________ because they lost or gained electrons

Cation – lost electron(s), ______ charge

Anions – gained electron(s), ______ charge

Periodic Table Families

Family Name	# Valence Electrons	Common Charge Formed
alkali metals		
alkaline earth metals		
halogens		
noble gases		

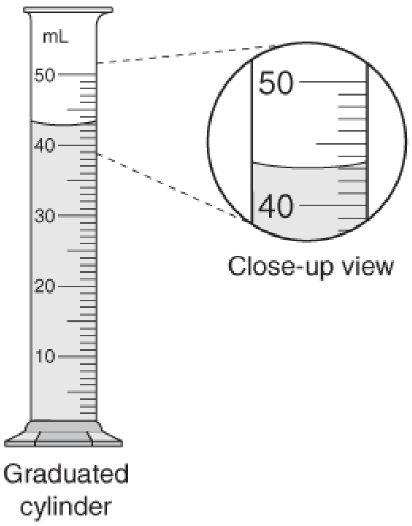


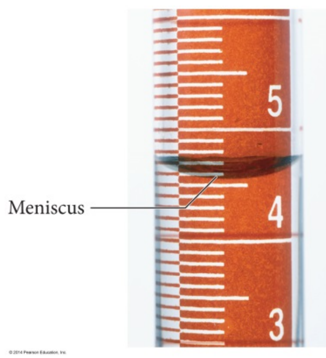
<u>Jigsaw Station 1: Sig Fig Review</u>!

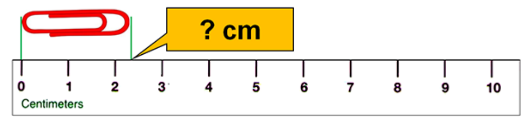
Significant Figures (otherwise and forever known as _____________) in a measurement consist of all the digits known with certainty plus one final digit, which is estimated.

Rule of Thumb: Always measure to the place value measured by the markings on your instrument (e.g. ruler, graduated cylinder) **PLUS ONE MORE.**

Let's Practice! Using correct significant figures, what is the measurement that is represented in each picture?

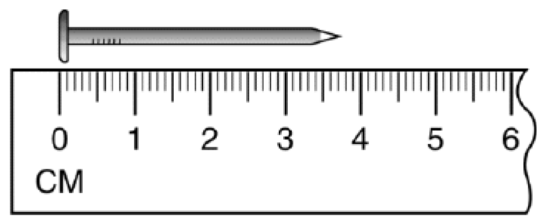






3. Markings? _________________

Measurement? _______________



1. Markings? _______________

Measurement? _____________

2. Markings? _______________

Measurement? _____________

4. Markings? _______________

Measurement? _____________

Rules for Significant Figures

1. <u>Non-zero digits</u> and <u>zeros between non-zero digits</u> are always significant.

   Examples: 102 → ____ sig fig(s), 1.73005 → ____ sig fig(s)

2. <u>Leading zeros</u> are not significant.

   Examples: 0.37 → ____ sig fig(s), 0.0001 → ____ sig fig(s)

3. <u>Zeros to the right</u> of all non-zero digits are only significant if a decimal point is shown.

   Examples: 100 → ____ sig fig(s), 100. → ____ sig fig(s), 0.0100 → ____ sig fig(s)

4. For values written in <u>scientific notation</u>, the digits in the coefficients are significant.

   Examples: 5×10^4 → ____ sig fig(s), 1.30×10^{-13} → ____ sig fig(s)

5. <u>Counting numbers</u> (one kitten, two kittens) and <u>conversion factors</u> (6.022×10^{23} atoms/1 mol) are considered exact value and have an infinite number of sig figs!

   Examples: 3 cats → ____ sig fig(s), 1 student → ____ sig fig(s)

Now You Try! Fun with Tasty Sig Figs.

Number	How many Sig Figs?	Number	How many Sig Figs?
3.0800 mL		55 puppies	
0.00418 g		1,800. m	
0.00000040 L		1,800 m	
3 people		2.998 x 10^8 m/sec	

Rules for Sig Fig Calculations (Hint: *Alpha order!*)

1. **<u>Adding/subtracting</u>**: round to least precise <u>place value</u> (A → P) *Awesome People*
2. **<u>Multiplying/dividing</u>**: round to least precise <u>total number</u> (M → T) *Memorization Technique* ☺
3. <u>Note</u>: **do not round any of the numbers you are given until the very end** after you have plugged them into your equations in their full, precise glory!

Let's Try! *Example #1* *Example #2* *Example #3*

Example #1

$$\begin{array}{r} 2.348 \\ 0.07 \\ +\ 2.9975 \\ \hline \end{array}$$

Example #2

$$\begin{array}{r} 5.9 \\ -\ 0.261 \\ \hline \end{array}$$

Example #3

$$\begin{array}{r} 1{,}010 \\ 2.9 \\ -\ 0.76 \\ \hline \end{array}$$

Example #4: 1.052 x 12.504 x 0.53 =

Example #5: 2.0035 ÷ 3.20 =

Example #6: 6.78 x 5.903 x (5.489 – 4.99) =

More Tasty Calculations Practice!

	Calculator Answer	**Rounded Answer** (with Correct # of Sig Figs)
1. 170 + 3.5 – 28		
2. 47.0 ÷ 2.2		
3. 691,300 ÷ (5.022 – 4.31)		
4. (0.054 + 1.33) × 5.4		

Jigsaw Station 2: Atomic Structure and Types of Matter

<u>Coulomb's Law</u>: fundamental relationship between electrostatic _______________ and ________________.

- It applies to charged particles, magnets, gravitation, etc.
- In chemistry, we are most interested in the _____________ of attraction or repulsion between __________________ particles

$$E \propto \frac{Q_1 Q_2}{r}$$

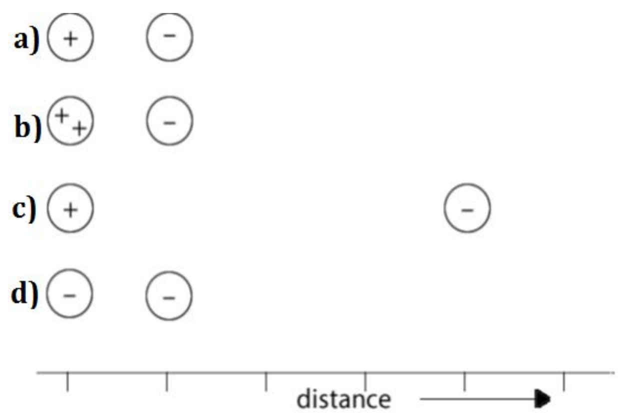
E= energy of attraction or repulsion between particles
Q_1 = charge of first particle
Q_2 = charge of second particle
r = distance between charged particles

In short:

- Energy of attraction/repulsion ______________ as the magnitudes (sizes) of the charges ______________
- Energy of attraction/repulsion ______________ as the distance between the charges ______________

Thought question: Will an electron be more attracted to the nucleus of a hydrogen atom or a helium atom, and why?

<u>Examples:</u>



1. Consider the particles in the diagram to the right.
   a. Compare the particles shown in (a) and (b). Which pair is more attracted to each other and why?

   b. Compare the particles shown in (a) and (c). Which pair is more attracted to each other and why?

   c. Compare the particles shown in (a) and (d). Which pair is more attracted to each other and why?

2. An electron would be most attracted to a nucleus containing which of the following?
   a. 7 protons, 5 neutrons
   b. 8 protons, 5 neutrons
   c. 5 protons, 5 neutrons
   d. 7 protons, 8 neutrons

Atomic Theory

- ______________________ is anything that has mass and takes up space.
- All matter is made up of _________________.

Structure of the Atom

- The atom can be divided into two regions: the ___________________ and the _________________________________.
  1. The __________________ is a very small region near the center of an atom that is positively charged.
  2. The ________________________ is a very large region that surrounds the nucleus and is negatively charged. It consists mostly of empty space.
- The atom is composed of __________________ subatomic particles

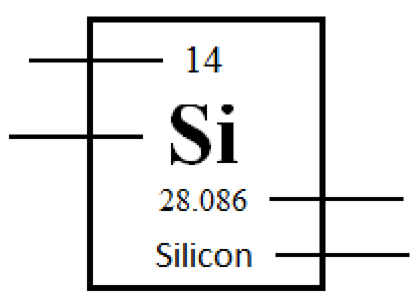
The Three Subatomic Particles

Particle	Symbol	Location	Charge	Mass (amu)

- The unit of mass for atomic particles is the ____ (amu)
  - 1 amu = one-twelfth the mass of a carbon atom containing six protons and six neutrons.

Understanding the Periodic Table

- **Atomic Number**: the **number** that tells you the ______________ of the element; number of _____________
- **Average Atomic Mass**: average mass of all of the element's _________________
  - To find the mass of a SPECIFIC (isotope) atom, you must add up the protons + neutrons
- **Symbol**: shortened element name; starts with a ______________ letter
- **Name**: the ______________ of an atom (not a proper noun = not capitalized in sentences! ☺)



Isotopes

- What are **isotopes**? Atoms of the ___________ element, but different ____________________
- This means the number of ____________________ is the same, and the number of ____________________ is different.

→Mass of an isotope = # protons + # neutrons ←

Two ways to write isotopes:

Type	**hyphen-notation**		**isotope notation/ nuclide symbol**	
Definition	name-mass		$^{\text{mass \#}}_{\text{atomic \#}}\text{Symbol}$	
Example	carbon-12	**vs**	$^{12}_{6}C$	
More examples				

Examples: Consider the following sets of isotopes and then explain the similarities and differences between *each* set.

Set I

$^{12}_{6}C$ $^{13}_{6}C$

Similarities:

Differences:

Set II

$^{24}_{12}Mg$ $^{24}_{11}Na$

Similarities:

Differences:

Practice: Complete the following table using your knowledge of atomic structure.

Element	Hyphen notation	Atomic Number	Number of Protons	Number of Neutrons	Number of Electrons	Mass Number
$^{1}_{1}H$						
	chlorine-35					
		27		33		

Classifying Matter

- **Matter**: anything that occupies space and has mass.
- We classify matter according to its **composition** (the basic components that make it up).

Matter

Can it be physically separated?

NO! → **Pure Substances**

YES! → **Mixtures**

Pure Substances

Can it be chemically separated into simpler substances?

NO! → **Element**

YES! → **Compound**

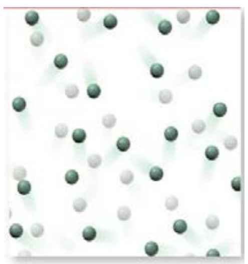
Mixtures

Uniform throughout?

NO! → **Heterogenous Mixture**

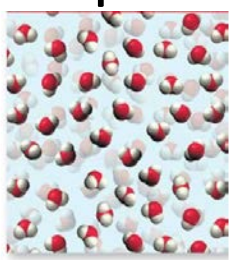
YES! → **Homogenous Mixture (Solution)**

Element



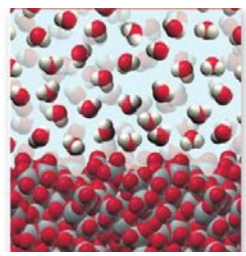
Examples
1.
2.
3.

Compound



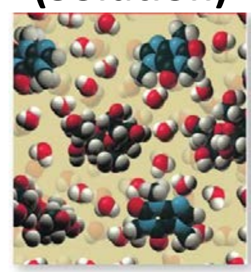
Examples
1.
2.
3.

Heterogenous Mixture



Examples
1.
2.
3.

Homogenous Mixture (Solution)



Examples
1.
2.
3.

Pure Substances - Must be separated chemically (bonds broken)

- Same/fixed ______________________________
- **Elements**
  - Cannot be broken down and still maintain identity
  - One type of atom (basic building blocks)
  - Found on the Periodic Table!
- **Compounds**
  - Chemical combination of ____________ or more elements in fixed, definite proportions
  - Cannot be separated by physical means
  - Properties of compound are different than individual elements

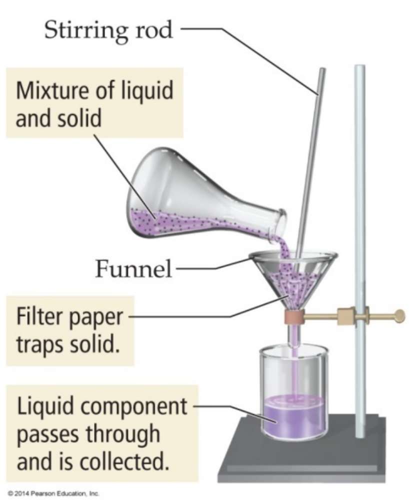
Mixtures - Can be separated by physical means

- Formed when two or more substances (s, l, g, aq) are physically combined
- All substances in mixture retain their own ___________________ properties
- **Heterogenous**
  - Parts of the mixture are NOT evenly distributed (poorly mixed)
  - Does _____________ look the same throughout
- **Homogenous**
  - Parts of the mixtures are ___________________ distributed (evenly mixed)
  - Also called a ____________________

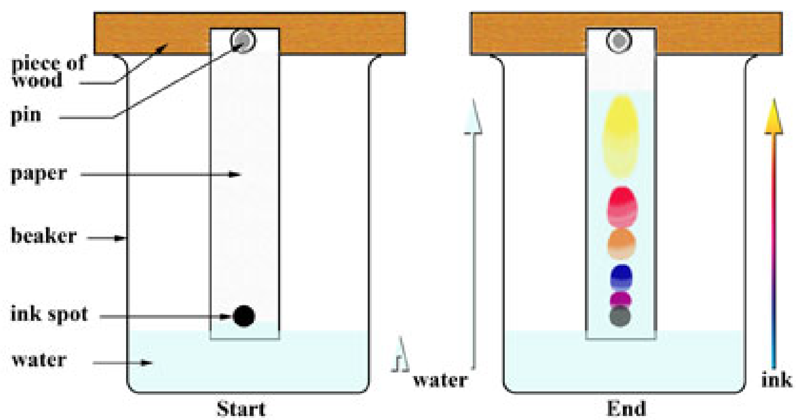
Methods for Separating a Mixture: Both heterogeneous and homogeneous mixtures can be separated by ________________ means into the component parts that make up the mixture.

1. A solid and liquid mixture can be separated by pouring the mixture through a ____________ paper designed to allow only the liquid to pass.
2. A homogeneous mixture of liquids can be separated using ___________________, a process in which the mixture is heated and the more volatile (more easily vaporized) liquid is boiled off first. A condenser is then used to recollect the vaporized component.
3. Paper _______________________ takes advantage of the fact that different components of a homogeneous mixture have different attractions to a solvent and paper.

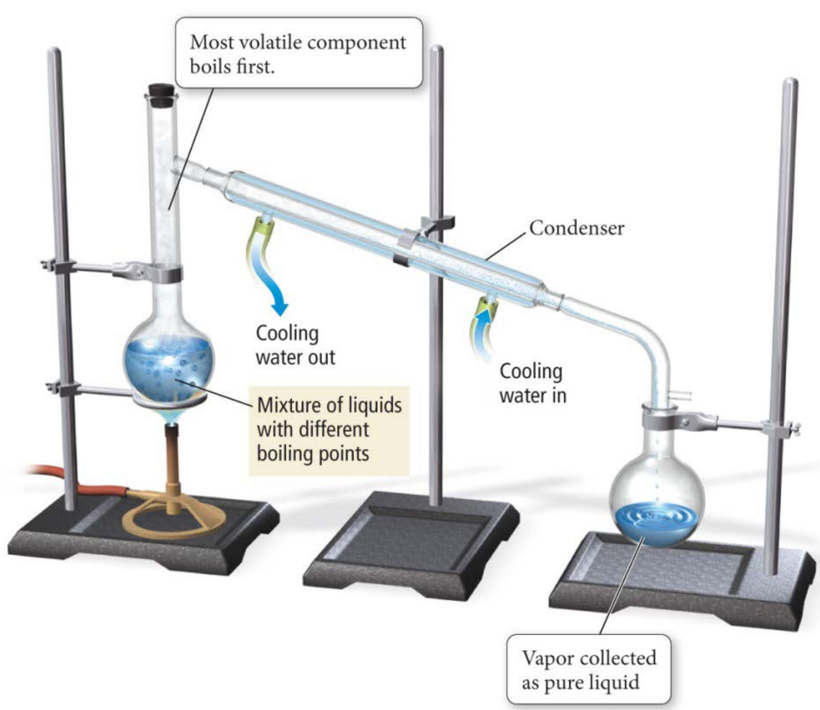
Filtration



Simple chromatography



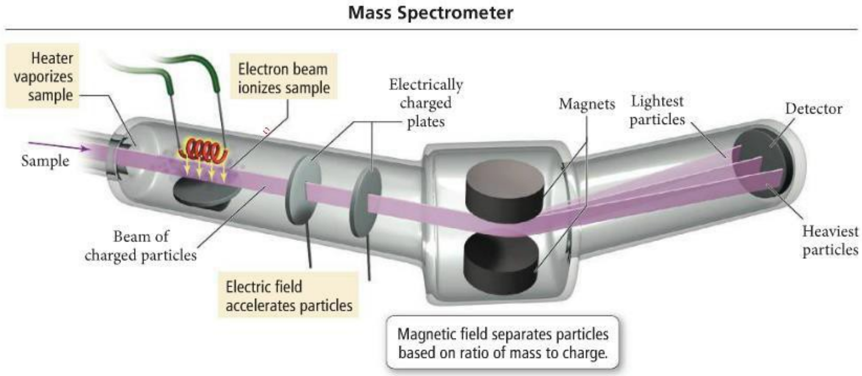
Distillation

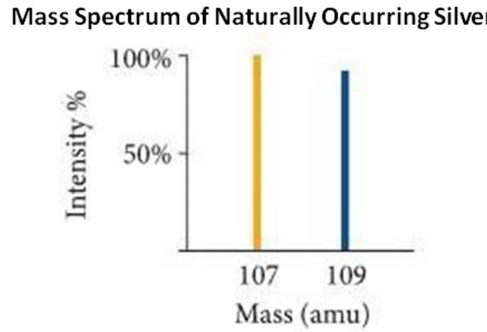


Jigsaw Station 3: Average Atomic Mass and Mass Spec

Mass Spectroscopy

The masses of elements and their percent abundances of isotopes of elements are measured using ___________ **spectrometry**: a technique that separates particles according to their mass, producing a mass ________________.





a. The position (location) of each __________on the x-axis indicates the ________ of the isotope.

b. The intensity (indicated by the___________________of the peak) indicates the relative abundance (how common that isotope is in __________________).

Average Atomic Mass: the weighted average mass of an element's isotopes and is the mass found on the periodic table.

$$Average\ atomic\ mass = \ mass_1\left(\frac{\%\ Abundance}{100}\right) + mass_2\left(\frac{\%\ Abundance}{100}\right) + \cdots$$

- The average atomic mass will be _______________ the mass of the largest and the mass of the smallest isotope.
- The average atomic mass will generally be ______________ to the most abundant isotope.
- Note: It is important to understand that the masses of a proton and neutron are ***approximately*** 1 amu, but the actual mass of each isotope is ___________ a whole number (mmm, nuclear binding energy). When specific, non-whole number masses are provided for each isotope, use the specific masses!

Guided Practice: Delicious average atomic mass practice.

1. Silicon has three, stable, naturally occurring isotopes. These are silicon−28, silicon−29, and silicon−30. The relative abundance of each is 92.21%, 4.70%, and 3.09% respectively.
   a. ESTIMATE the value of the answer before you begin. Will the weighted average be closer to 28, 29, or 30? Why?

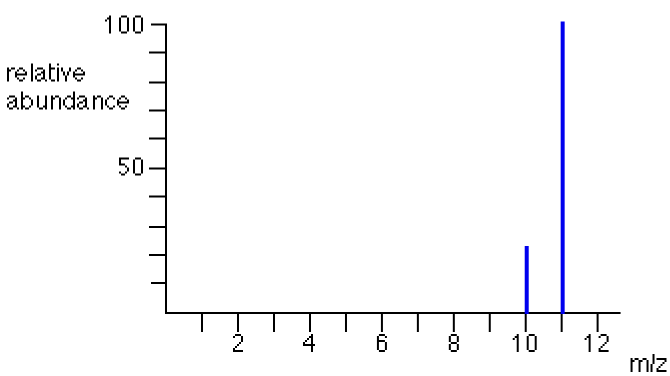
   b. What is the average atomic mass of silicon?

2. Calculate the average atomic mass of magnesium using the following data for three, stable magnesium isotopes.

Isotope	*mass (amu)*	*relative abundance*
Mg-24	23.985	78.70%
Mg-25	24.986	10.13%
Mg-26	25.983	11.17%

3. Copper has only two stable isotopes, and an average atomic mass of 63.546 amu. Cu-63 has 69.17% abundance and 62.940 amu. What is the mass of the second isotope of copper?

4. Given the mass spectrum and data for boron below, estimate the average atomic mass of boron.



Isotope	Peak Intensity
boron−10	23%
boron−11	100.%

5. Europium has two stable isotopes: ^{151}Eu with a mass of 150.9196 amu and ^{153}Eu with a mass of 152.9209 amu. If elemental Europium is found to have a mass of 151.96 amu, calculate the percent of each of the two isotopes. (Hint: Use a system of equations. ☺)

Jigsaw: All Together Practice!

1. A halide is a binary compound, of which one part is a halogen atom and the other part is an element or radical that is less electronegative than the halogen, to make a fluoride, chloride, bromide, iodide or astatide compound. What is the general formula for alkali metal halides?

   a. MX b. M_2X c. MX_2 d. MX_3

2. Which electrons account for many of the chemical and physical properties of elements?

   a. Innermost b. Intermediate c. Outermost d. Transition

3. Which of the following elements is not correctly paired with its group (family) name?

   a. Bismuth (Bi), halogens
   b. Strontium (Sr), alkaline earth metals
   c. Lithium (Li), alkali metals
   d. Radon (Rn), noble gases

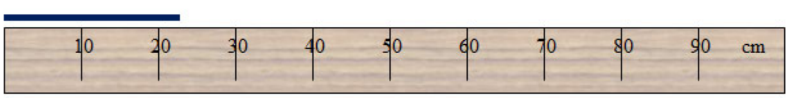
4. The inertness of noble gases is due to:

   a. The number and arrangement of their electrons
   b. The unique structure of their nuclei
   c. The special number of protons and neutrons
   d. The bonds they form with other elements

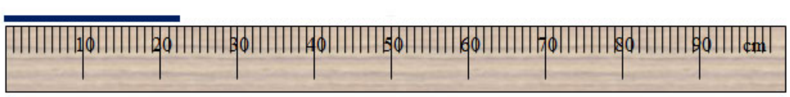
Sig Fig Practice

Directions: Measure the following to the correct number of significant figures.

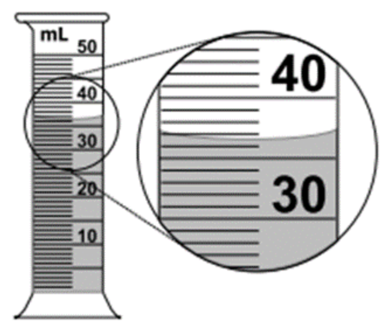
1. ___________



2. ___________



3. ___________



4. ___________

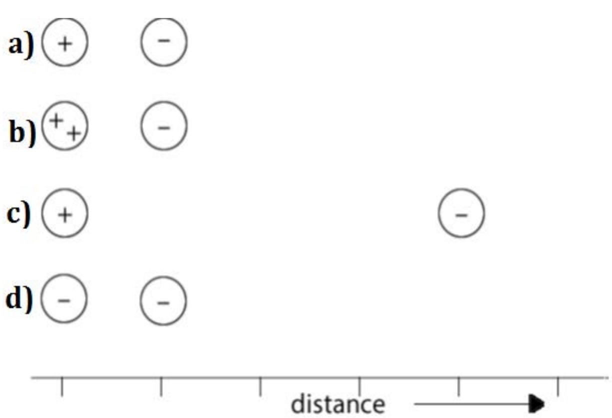
Directions: Determine the number of significant figures.

1. 0.02 ________	**2.** 142 ________	**3.** 0.02020 ________	**4.** 0.073 ________
5. 501 ________	**6.** 1.071 ________	**7.** 501.0 ________	**8.** 10810 ________
9. 5000 ________	**10.** 5.00 ________	**11.** 5000. ________	**12.** 1.20×10^3________

Directions: Round your answer to the correct number of sig figs.

	Calculator Answer	**Rounded Answer** (with Correct # of Sig Figs)
1. 140 × 35		
2. 0.003 + 0.0048 + 0.100		
3. 67.35 ÷ (1.401 – 0.399)		
4. (6.23 + 3.111 – 0.05) × 14.99		
5. 3.14159 × (4.17 + 2.150)		

Coulomb's Law



1. Which set of particles shown the left will experience:
    a. the greatest attraction to each other? _______
    b. the greatest repulsion from each other? _____
2. The particles in (a) will experience ______________ attraction to each other than the particles in (c) because:
    a. greater, the distance between them is less.
    b. smaller, the distance between them is less.
    c. the same, distance is irrelevant to force of attraction.

3. The particles shown above in (a) will experience ______________ attraction to each other than the particles in (b) because:
    a. greater, the nucleus has one proton instead of two.
    b. smaller, the nucleus has one proton instead of two.
    c. the same, charge is irrelevant to force of attraction.

4. An electron in the lowest energy level would be most attracted to the nucleus of which element?
   a. lithium
   b. sodium
   c. potassium
   d. rubidium

5. A proton would be least repulsed by the nucleus of which element?
   a. helium
   b. neon
   c. argon
   d. krypton

Isotopes! Complete the chart below.

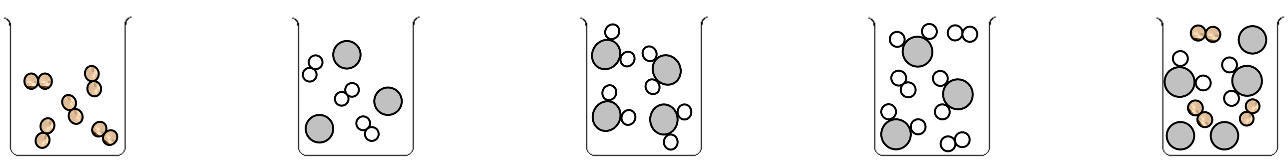
Hyphen notation	Isotope Notation	Mass #	Atomic #	Protons	Neutrons	Electrons
	$^{10}_{5}\mathrm{B}$					
phosphorus–31						
		66		30		
			27		32	

Types of Matter

Directions: Identify the following as: element (E), compound (C), heterogeneous mix (He) or homogeneous mix (Ho).

a. Table salt __________
b. Nitric acid (HNO_3) __________
c. Sugar (Glucose) __________
d. Carbon dioxide (CO_2) __________
e. Milk __________
f. Air __________
g. Nitrogen gas (N_2) __________
h. Zinc (Zn) __________
i. Pulpy orange juice __________

Directions: Use the particle representations below to answer #1-5. **Answer choices may be used more than once!**



A B C D E

1. Compound ______
2. Nitrogen, N_2 ______
3. Mixture of two elements ______
4. Element ______
5. Mixture of water, H_2O, and hydrogen H_2 ______

Directions: Choose the best possible answer for each question.

1. A forensic technician is examining a fine white powder found at a crime scene. It appears to be of uniform texture and consistency. He finds that upon heating, the substance decomposes, releasing a colorless gas and leaving a black residue. Based on these observations, the substance is:
   A a homogeneous mixture
   B a heterogeneous mixture
   C either a homogeneous mixture or a compound
   D a compound

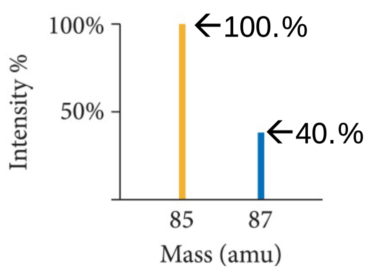
2. An example of a heterogeneous mixture is
   A bronze. **B** concrete. **C** brass. **D** water.

3. How is a mixture different from a compound?
   A Particles of a mixture are combined chemically.
   B Components of a mixture can only be separated chemically.
   C Components of a mixture can be separated by physical means.
   D Composition of a mixture may be constant.

4. A transparent liquid has uniform color. Under a microscope, no difference in uniformity is apparent. The liquid all boils away at the same temperature. Based on this information, it can be concluded that the liquid is:
   A an element
   B a homogeneous mixture
   C either a pure substance or a homogeneous mixture
   D a pure substance

Average Atomic Mass and Mass Spec

1. Consider the mass spectrum shown below.
   a. Determine the average atomic mass of the element using the mass spec shown below.



   b. Which element is most likely shown here? ___________________
   c. How did you identify the element?

2. There are two stable isotopes of calcium: Ca-40 (39.96) and Ca-46 (45.95). Using the average atomic mass of calcium from the periodic table, calculate the % abundance of each isotope of calcium. (Yum, algebra.)

3. Silver consists of two stable isotopes, one with a mass of 106.905 and an abundance of 51.84%.

a. What is the abundance and mass of the other isotope?

b. How many silver-107 atoms are present in a 2.00 gram sample of pure silver?

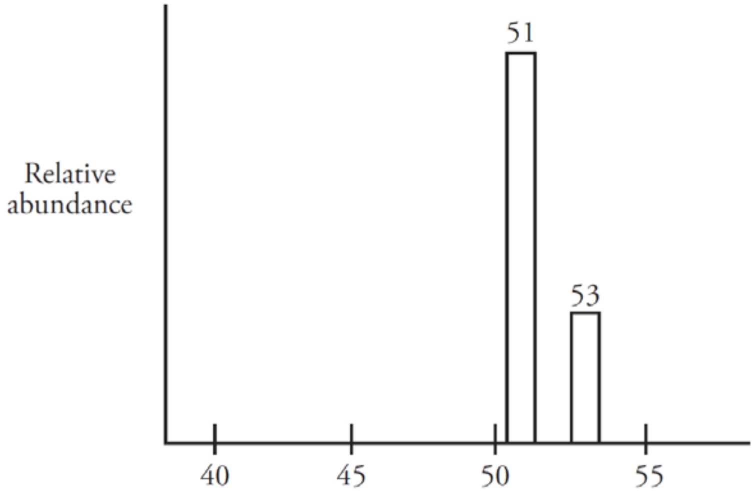
4. There are three naturally occurring isotopes of carbon: carbon−12, carbon−13, and carbon−14. In a natural sample of carbon, which is most likely to be the most abundant isotope? (No math needed! ☺)

5. Lithium and bromine each have two naturally occurring isotopes. Each is described in the table below.

Isotope	Mass (amu)	Abundance
lithium−6	6.015	7.59%
lithium−7	7.016	92.41%
bromine−79	78.918	50.69%
bromine−82	81.916	49.31%

Lithium and bromine combine to form lithium bromide (LiBr). How many peaks will be present in the mass spectrum of LiBr?

a. 1 peak b. 2 peaks c. 3 peaks d. 4 peaks



6. The above mass spectrum is for the hypochlorite ion, ClO^-. Oxygen has only one stable isotope, which has a mass of 16 amu. Using the spectrum, calculate the average mass of a hypochlorite ion.

a. 52.5 amu b. 52.0 amu c. 51.5 amu d. 51.1 amu

7. The table below shows the atomic mass and natural abundance of the two naturally occurring isotopes of lithium.

Naturally Occurring Isotopes of Lithium

Isotope	Atomic Mass (u)	Natural Abundance (%)
Li-6	6.015	7.6
Li-7	7.016	92.4

Which numerical setup can be used to determine the atomic mass of naturally occurring lithium?

a. (7.6)(6.015 amu) + (92.4)(7.016 amu)

b. (0.076)(6.015 amu) + (0.924)(7.016 amu)

c. $\frac{(7.6)(6.015\ amu) + (92.4)(7.016\ amu)}{2}$

d. $\frac{(0.076)(6.015\ amu) + (0.924)(7.016\ amu)}{2}$

8. The average atomic mass of naturally occurring neon is 20.18 amu. There are two common isotopes of naturally occurring neon as indicated in the table below. Using the information provided, calculate the percent abundance of each isotope.

Isotope	Mass (amu)
Ne-20	19.99
Ne-22	21.99

a. 90.5% ^{20}Ne and 9.5% ^{22}Ne

b. 98.2% ^{20}Ne and 1.2% ^{22}Ne

c. 10.5% ^{20}Ne and 89.5% ^{22}Ne

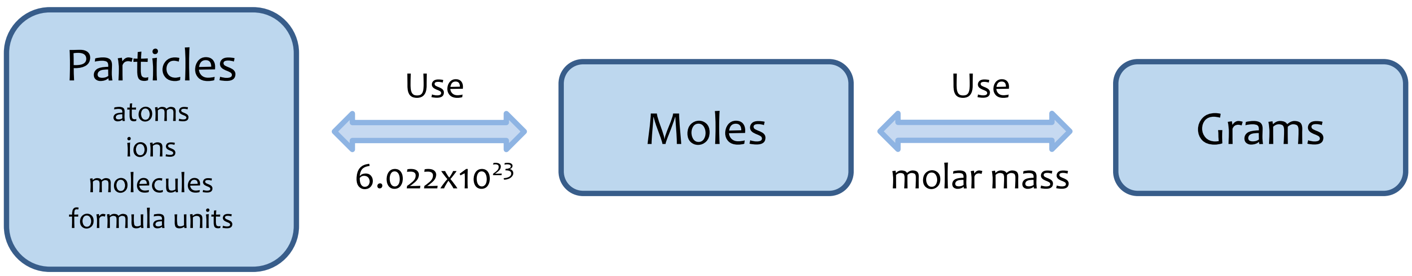
d. 56.4% ^{20}Ne and 43.6% ^{22}Ne

Here and Back Again: A Mole-ish Review

Part I: The Mole – Massively Important

Type of Particle	Definition	Example
	single element	Si, C, Mg, Ca, Cu, Zn, S Br, I, N, Cl, H, O, F
	charged element/polyatomic	N^{3-}, SO_4^{2-}, PO_4^{3-}
	covalent compound (diatomic elements are considered molecules)	C_2H_6, H_2O Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2
	ionic compound	$CuCl_2$, Na_2SO_4, KBr

→ Don't forget about your 7 diatomic molecules! Yay, $Br_2I_2N_2Cl_2H_2O_2F_2$ ←



Examples and Practice:

1. What is the mass in grams of 1.60 x 10^{24} molecules of B_3Br_6?

2. How many moles of nitrogen gas are found in a 57.2 g sample?

3. How many formula units of $MgCO_3$ are contained in a 2.34 mol sample?

Mole Conversions Using a Chemical Formula

- A chemical ________________ shows the kinds of elements and numbers of atoms or moles of each element in the smallest representative unit of a substance.
- You can use a chemical formula as a ________ ratio to convert between moles of a compound and moles of an atom ___________ a compound

****If a problem give a ____________________ but asks for _________ or _________ (or the opposite), then an extra conversion fraction is required!**

Examples and Practice:

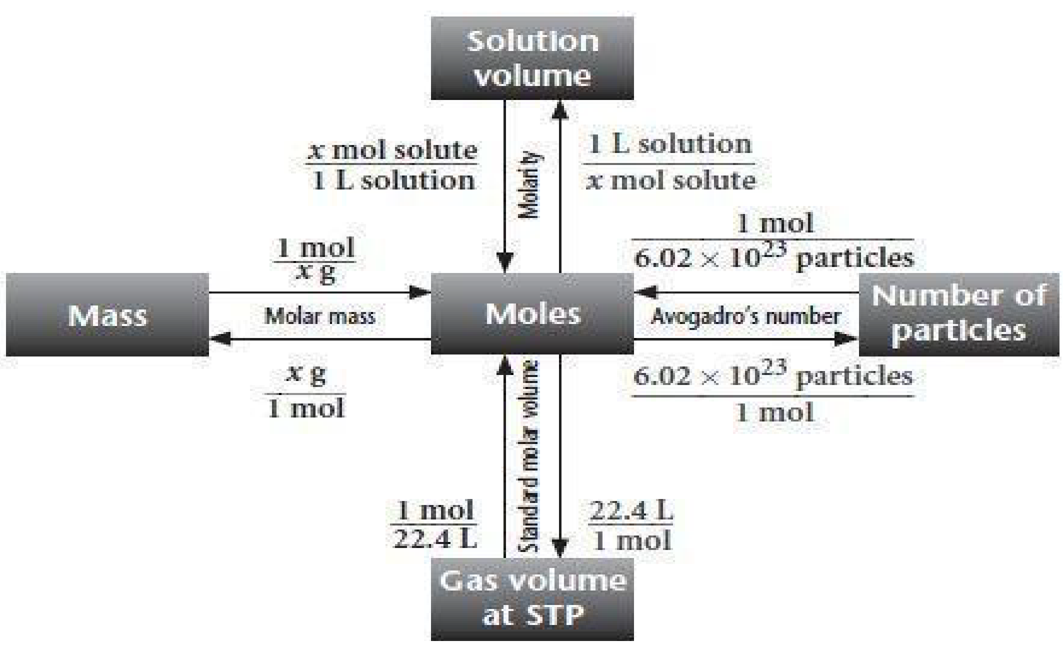
1. A student measures out 2.0 moles of Li_3PO_4. How many moles of lithium ions, Li^+, are in the sample?

2. Freon, which has the formula CCl_2F_2, is used as a refrigerant in air conditioners and as a propellant in aerosol cans. Given a 5.56 mg sample of Freon, calculate the number of molecules of freon in that sample.

3. If two atoms of carbon combine with four atoms of hydrogen in the compound ethene (C_2H_4), how many grams of hydrogen would be needed to combine completely with 6.0 grams of carbon?

4. How many total protons are present in a 4.10 mole sample of CH_4?

5. Given a sample of 3.56 x 10^{24} molecules of C_4H_8, how many moles of carbon are in the sample?



Part II: The Mole – A Solution to Every Problem

- A ____________________ is a homogeneous mixture: the properties are the same no matter what part of the sample one examines.
- Solution concentration is usually expressed in terms of _____________ (M), i.e., the number of **moles** of solute per **liter** of solution.

$$M = \frac{mol}{L}$$

- Molarity has the units of _____, ________________, or:

1. You prepare a solution by dissolving 48.05 g of $MgBr_2$ in enough water to make 800. mL of solution.

a. Calculate the molarity of Mg^{2+} and Br^-.

b. How many bromide ions are found in this solution?

c. What fraction of the total number of ions are bromide ions?

2. How many OH^- ions are contained in 2.5 L of a 0.52 M $Al(OH)_3$ solution?

Part III: The Mole – It's Such a Gas!

- At ________, **one mole** of any ideal gas occupies a volume of ______________.
  - STP: Standard temperature is 0°C (273K) and standard pressure is 1 atm = 760 mm Hg = 760 torr.
- Gases must have their quantitative properties calculated using temperatures in _________________.
- The ideal gas equation, _____________________, can be used to relate the properties [pressure (P), volume (V), number of moles (n), and *Kelvin* temperature (T)] of ideal gases to one another.
  - R = universal gas constant = 8.314 J mol^{-1} K^{-1} = 0.08206 L atm mol^{-1} K^{-1} = 62.36 L torr mol^{-1} K^{-1}
  - The numerical value of R is *dependent upon the units used*.
- A rhyme to help you remember! **When you're _______ at STP, use PV = nRT.**

3. Given 7.81 x 10^{22} molecules of chlorine gas (Cl_2):

a. What is the volume of your sample at STP?

b. What is the volume of your gas sample in Texas in late August? Let's say 1.00 atm and 38°C (about 100°F, sigh).

4. Three identical containers are filled with a single gas each: the first container with O_2, the second with CO_2, and the last with N_2. All of the containers are at the same temperature and pressure.

a. Which container has the greatest number of molecules?

b. Which container has the smallest mass?

Mole Review: Multiple Choice Practice Problems

1. A single sodium atom has an average mass of 22.99 amu (taken from the periodic table). How does this number relate to a mole of sodium atoms?

a. One mole of sodium atoms will have a mass of 22.99 amu.

b. One mole of sodium atoms will have a mass of 22.99 grams.

c. One mole of sodium atoms will have a mass of 22.99 x 6.022 x 10^{23} grams.

d. You cannot relate the mass of a single sodium atom to the mass of a mole of atoms.

2. How many protons are in 3.50 moles of lithium atoms?

a. 3.50 b. 10.5 c. 3.50 x (6.022 x 10^{23}) d. 10.5 x (6.022 x 10^{23})

3. Which of the following samples contains the largest number of atoms?

a. 3.0 mol of H_2O b. 3.0 g of H_2O c. 3.0 x 10^{22} molecules of N_2

4. What volume will 37.8 g of oxygen gas occupy at STP?

a. 847 L b. 52.9 L c. 26.5 L d. 1.69 L

5. At 0°C and 1 atm, a 0.25 L container would hold how many grams of carbon monoxide, CO? (The molar mass of carbon monoxide is 28.01 g/mol).

a. 7.0 g b. 3.2 g c. 0.31 g d. 0.011 g

6. If you need 13.9 grams of lithium atoms, how many grams of lithium sulfide, Li_2S, would be required?

a. 27.8 g   b. 39.1 g   c. 46.0 g   d. 92.0 g

7. Calculate the concentration of a solution prepared by dissolving 11.85 g of solid $KMnO_4$ in enough water to make 750. mL of solution. (The molar mass of $KMnO_4$ is 158.04 g/mol).

a. 0.100 M   b. 0.0562 M   c. 1.00 M   d. 0.562 M

8. How many moles of a gas at 247^{o}C would occupy a volume of 6.4 L at a pressure of 260 mmHg?

a. 0.051 mol   b. 0.11 mol   c. 0.28 mol   d. 3.9 mol

9. How many total ions are contained in 250 mL of a 0.100 M sodium chloride, NaCl, solution?

a. 1.5×10^{22} ions   b. 3.0×10^{22} ions 1.5×10^{25} ions   b. 3.0×10^{25} ions

10. What is the molar concentration of $Al(OH)_3$ if a 500. mL solution of $Al(OH)_3$ contains 1.8×10^{24} ions of OH^-?

a. 0.020 M   b. 0.060 M   c. 2.0 M   d. 6.0 M

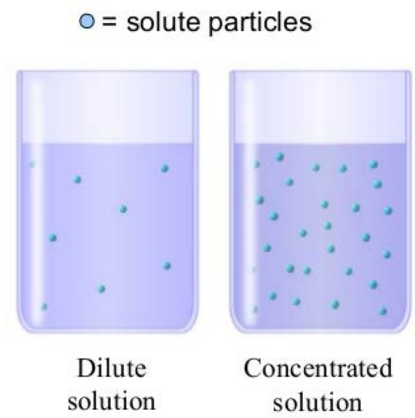
11. What is the total number of atoms in 123 grams of sulfur trioxide, SO_3?

a. 9.25×10^{23} atoms   b. 3.70×10^{24} atoms   c. 5.93×10^{27} atoms   d. 2.37×10^{28} atoms

Unit 10 Part 3: Molarity

Concentration of Solutions: a measure of the amount of solute in a given amount of solvent.

- **Dilute solution:** has a relatively ______________ amount of solute in a solvent.
- **Concentrated solution:** has a relatively ____________ amount of solute in a solvent.



Molarity: the number of _____________ of solute in ____________ of solvent.

- Molarity = "molar concentration"
- Usually the solvent is water.
- Molarity of a solution is represented using square brackets:
  - For example, "the molar concentration of NaCl" is written as [NaCl]

$$Molarity = \frac{moles\ of\ solute}{liters\ of\ solution\ (volume)} \rightarrow M = \frac{mol}{L}$$

→ To solve for molarity you need to measure: 1) mass of the solute and 2) total volume of the solution! ←

Examples

1. *You have 4.00 L of solution that contains 116.9 g of sodium chloride, NaCl. Calculate [NaCl].*

2. *You have 800. mL of a 0.50 molar HCl solution. How many moles of HCl does this solution contain?*

3. *To produce the desired amount of silver chromate in a reaction, you will need at least 97.1 g of potassium chromate in solution as a reactant. All you have on hand in the stock room is 5.00 L of a 5.0 M K_2CrO_4 solution. What volume (in liters) of the solution is needed to give you the 97.1 g of K_2CrO_4 needed for the reaction?*

Mini-Lab: Let's Make a Solution!!

Purpose: (fill in assigned solution): To prepare 100. mL of a _________M sodium chloride solution.

Calculations:

Procedure: (Write a detailed procedure. Someone else should be able to make the solution only using your steps below. Be sure to include proper lab equipment and the amounts used.)

Practice Makes Perfect!

1. What is the molarity of a solution composed of 33.2 g of potassium iodide, KI, dissolved in enough water to make 125 mL of solution?

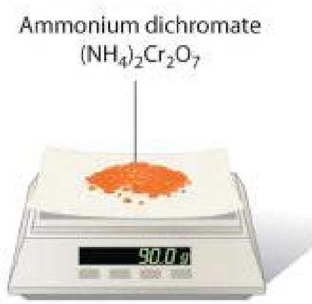
2. How many moles of H_2SO_4 are present in 0.500 L of a sulfuric acid solution if $[H_2SO_4]$ = 0.150 M?

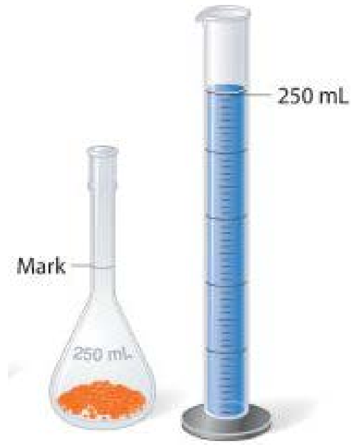
3. How many grams of NaCl are required to make 2.0 L of 4.00 molar NaCl?

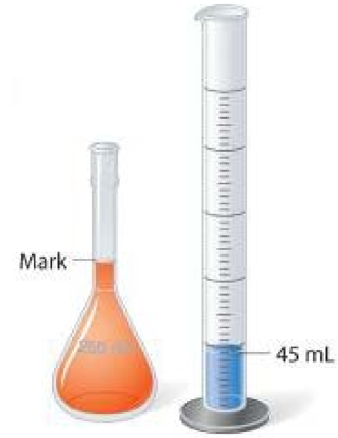
Molarity Mini-Lab

Part 1: How to Use a Volumetric Flask to Make a Solution

Purpose: (fill in assigned solution): To prepare ____________ mL of a ____________ M solution of ____________







How to Make a Solution with a Volumetric Flask

1. Calculate the mass of solute needed to make the desired solution.
2. Put a weigh boat (or weigh paper) on balance. Press the tare (or zero) button. The balance should now read zero with the weigh boat on top.
3. Measure out the calculated grams of solute.
4. Carefully pour the measured solid into the volumetric flask. Use a wash bottle filled with DI water to rinse the weigh boat into the volumetric flask (to make sure all remaining solid is added).
5. Add distilled water to the volumetric flask until the water level reaches the bottom of the skinniest part of the flask.

   Note: Wash bottles are easily contaminated: NEVER touch the tip of the bottle to anything (including your hands or the container you're adding water to).

6. Use a wash bottle or pipette to add water drop-by-drop until water reaches EXACTLY to the calibration line on the neck of the flask (the bottom of the meniscus should touch the line).
7. Cap the volumetric flask and mix thoroughly.

Note: Volumetric flasks are both accurate and precise.
However, a given volumetric flask can only measure **ONE** specific volume!

Solution Time!

1. Using your knowledge of chemistry, calculate the mass of solute needed to prepare ____________ mL of a ____________ M solution (the values assigned to your group at the beginning of lab).

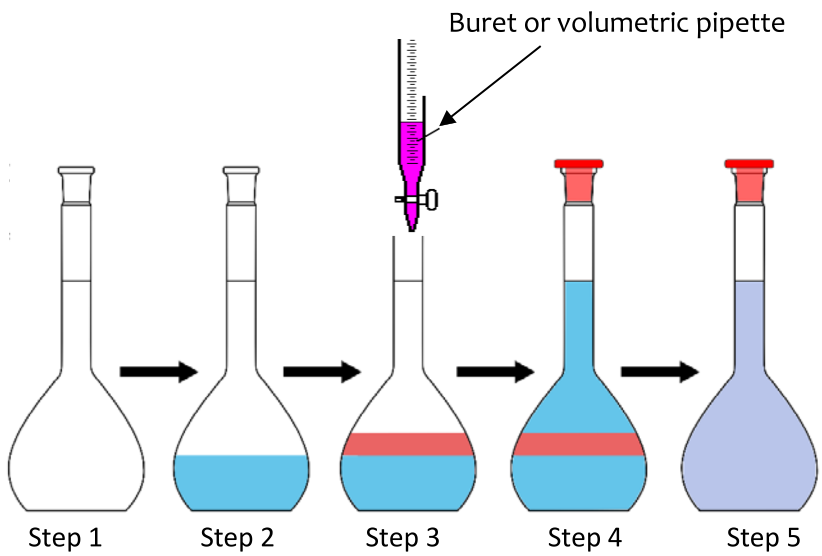
   Show your calculations below.

2. Make your solution. When you're done, label your completed solution with concentration and chemical formula.

Part 2: How to Use a Volumetric Flask for Dilution

Dilution Equation: (not given on the formula chart, but SOOOOO worth memorizing!)

$$M_1V_1 = M_2V_2$$



Step 1: Calculate the volume of stock solution needed to make the desired solution.

Step 2: Add distilled water to the volumetric flask until it is about 1/3 full.

Step 3: Add the pre-measured quantity of stock solution to the volumetric flask. For most precise results, measure the stock solution with a buret or volumetric pipette!

Step 4: Use a wash bottle or pipette to fill with water EXACTLY to the calibration line on the neck of the flask (the bottom of the meniscus should touch the line).

Step 5: Cap the volumetric flask and mix thoroughly.

Dilution Time!

1. Transfer your solution from part 1 into a beaker.
2. Add two drops of food coloring.
3. Use the dilution equation to calculate the amount of ____________ M stock solution of ____________ to create the diluted solution assigned to your group. (The solution you created in part 1 is your stock solution.)

   Show your calculations below.

4. Using the results of your calculations, answer the following questions:
   a. What volume of stock solution will you add to the volumetric flask? _________________
   b. How much deionized water will you add to complete the dilution? _________________
5. Make your dilute solution. When you're done, label your completed solution with concentration and chemical formula.

Analysis:

1. Why did we add food coloring to our stock solution before diluting?

2. How is a volumetric flask different from most other glassware we use in chemistry, like a beaker or an Erlenmeyer flask? When is a volumetric flask useful?

3. Write a step-by-step procedure explaining what you did: someone should be able to make the solution only using your steps below. Be sure to include proper lab equipment and the chemicals/amounts used. Choose your lab equipment from the list below (you won't need to use everything).

   50 mL beaker, 250 mL beaker, wash bottle, funnel, 100 mL graduated cylinder, pipette, volumetric flask with lid, balance, weigh boat, deionized water

Percent Composition by Mass

Percent Composition: the percent by ____________ of each element in a compound.

- According to the law of ____, a given chemical compound always contains the exact same elements in the exact same ratio by _____________.

$$\% \textbf{ compostion of an element} = \frac{\boldsymbol{total\ mass\ of}\ \textbf{element}\ \boldsymbol{in\ compound}}{\boldsymbol{total\ mass\ of}\ \textbf{compound}} \times \mathbf{100}$$

Guided Practice

1. Find the percentage composition of each element in the compound copper (I) sulfide, Cu_2S.

2. Find the mass percentage of water in sodium carbonate decahydrate, $Na_2CO_3 \bullet 10H_2O$, which has a molar mass of 286.15 g/mol.

3. When ammonia, NH_3, is formed, 1.0 gram of hydrogen reacts with about 5.0 grams of nitrogen. How much nitrogen would be needed to react with 2.5 grams of hydrogen in the production of ammonia?

4. Lake Superior is the largest lake in North America and contains about 1.2×10^{16} kg of water. What mass of hydrogen is contained in Lake Superior?
   a. 6.0×10^{15} kg
   b. 1.1×10^{16} kg
   c. 1.3×10^{15} kg
   d. 1.2×10^{16} kg

5. For a 150 g sample of glucose, $C_6H_{12}O_6$, there is 60 g of carbon. How many grams of carbon are there for a 300 g sample of glucose?
   a. 30 g
   b. 60 g
   c. 90 g
   d. 120 g

Empirical and Molecular Formulas

<u>Empirical Formula</u>: the symbols for the __________________ combined in a compound, with __________________ showing the smallest whole-number mole ratio of the different atoms in the compound.

<u>Molecular Formula</u>: the ______________formula of a compound which shows the total number of each atom in the molecule.

*** It is possible for the empirical formula and the molecular formula to be the ______________! ***

Molecular	Simplify by dividing	Empirical
$C_6H_{12}O_6$	6	CH_2O
P_4O_{10}		
	3	C_3HON_4
NaCl		

Steps for the Calculation of Empirical Formula

Step 1: Convert % composition to a mass composition by assuming 100 g of sample.

Step 2: Convert the mass into moles. **<u>DO NOT ROUND!! Keep at least 3 #</u>.**

Step 3: Divide by each number of moles by the smallest number of moles calculated in step #2.

Step 4: If not a whole number, multiply to get rid of the decimal (see chart on right) and use those whole numbers as subscripts in compound.

Fractional Subscript	Multiply by This
0.20	5
0.25	4
0.33	3
0.40	5
0.50	2
0.66	3
0.75	4
0.80	5

The Empirical Song! *(sung with the melody from Twinkle, Twinkle Little Star)*

Percent to mass and mass to mole,

Divide by small then multiply til whole.

That's how you find the empirical

Smallest whole-number ratio.

Let's Practice!

1. Analysis of a sample of a compound known to contain only phosphorus and oxygen indicates that it contains 43.67% phosphorus by mass. What is the empirical formula of this compound?

How to Determine the Molecular Formula

Step 1: Find the mass (formula weight) of the empirical formula.

Step 2: Take the molecular mass and divide it by empirical mass (this will always give you a whole number).

Step 3: Multiply the whole # by the empirical formula's subscripts to determine the molecular formula.

Let's Practice!

2. What is the molecular formula for a compound with the empirical formula H_2O and a molecular mass of 54 g/mol?

3. A sample of NutraSweet is 57.14% C, 6.16% H, 9.52% N, and 27.18% O. Calculate the empirical formula of NutraSweet and find the molecular formula. (The molar mass of NutraSweet is 294.30 g/mol)

4. The analysis of a rocket fuel sample showed that it contained 87.4% nitrogen and 12.6% hydrogen by weight. Mass spectral analysis showed the fuel to have a molar mass of 32.06 g/mol. What is the molecular formula of the fuel?

Multiple Choice Practice: Tasty and Delicious

1. An element X combines with oxygen to form a compound of formula XO_2. If 24.0 g of element X combine with exactly 16.0 g of O to form this compound, what is the atomic weight of element X?

   a. 48.0 amu   b. 24.0 amu   c. 16.0 amu   d. 12.0 amu

2. A new ore contains 52.3% silver by mass. How many grams of the ore are needed to obtain 10.0 moles of silver?

   a. 2,060 g   b. 1,080 g   c. 564 g   d. 10.0 g

3. A compound is made up of entirely silicon and oxygen atoms. If there are 14.0 g Si and 32.0 g O present, what is the empirical formula of the compound?

   a. SiO_2 b. SiO_4 c. Si_2O d. Si_2O_3

4. A chemist suspects that a given sample of $CaSO_4$ is impure. Upon testing, the chemist finds that the sample contains 31.6% calcium, but pure $CaSO_4$ is 29.4% calcium by mass. Which of the following might account for the measured percent mass of the sample?

   a. The sample is composed entirely of $CaSO_4$.

   b. The sample is a mixture of $CaSO_4$ and $CaCO_3$.

   c. The sample is a mixture of $CaSO_4$ and $MgSO_4$.

   d. The sample is a mixture of $CaSO_4$ and $Ca(BrO_3)_2$.

5. Styrene has the empirical formula CH with a molar mass of 104.13 g/mol. Approximately how many hydrogen atoms are present in a 52 g sample of styrene?

   a. 5.0×10^{22} H atoms
   b. 2.4×10^{24} H atoms
   c. 8.0×10^{23} H atoms
   d. 6.1×10^{23} H atoms

6. What is the empirical formula for a compound that contains 7.48 g N and 1.08 g H?

   a. N_2H
   b. NH_2
   c. NH
   d. NH_5

7. If the compound in #6 has a molecular mass of 32.05 g/mol, what is its molecular formula?

   a. N_2H_2
   b. N_4H_2
   c. N_2H_{10}
   d. N_2H_4

8. An Olympic medal contains 71.5% of gold by mass. How much gold could be extracted from a medal that weighs 115 g?

   a. 0.417 mol
   b. 0.817 mol
   c. 2.43 mol
   d. 4.86 mol

9. Two different samples are analyzed. Sample A contains 2.8 g of N and 1.6 g of O. Sample B contains 14.0 g of N and 8.0 g of O. Which of the following statements is most likely to be true?

   a. Sample A and B are the same compound because they contain the same types of atoms.
   b. Sample A and B are the same compound because their mass ratios indicate they both contain the same ratio of atoms within the molecule.
   c. Sample A and B are different compounds because they contain different numbers of atoms indicating a different ratio of atoms within the molecule.
   d. Sample A and B are different compounds because their molar masses are different.

Hydrates: Salty salts with a hidden surprise!

A **hydrate** is a ________ substance (often ionic) that contains a fixed composition of water molecules (known as "waters of hydration") embedded in its crystal structure.

- ➔ Heating a hydrate "drives off" the water molecules, and the solid that remains behind is called anhydrous, meaning "without water."
- ➔ By measuring the mass of water removed when dehydrating a hydrate, we can determine the ratio of water molecules to anhydrous salt for a given hydrate, which allows us to find the formula of the hydrate!

Notes about Language: Talking about hydrates can be tricksy! Here's a quick guide to the terminology used.

Word/ Phrase	Meaning/ Context
Hydrate	Pure substance, typically crystalline, containing a fixed ratio of water molecules within its structure → this term is only used *before* water is removed by heating!
Water molecules "driven off"	The process of forcing out the embedded water molecules in a hydrate through heating (do NOT call this evaporation – different context!)
Anhydrous salt	Compound remaining after all water molecules have been 'driven off' (removed)

Hydrate Math

The percent of water in a hydrate can be determined in a manner similar to determining the percent composition of a compound.

$$\%\ water = \frac{mass\ of\ water\ lost}{mass\ of\ hydrate} \times 100\%$$

Steps to gravimetrically (by mass) **determine the formula of a hydrate:**

1. Determine the ________ of the water that has left the compound.
2. Convert the mass of water to __________.
3. Convert the mass of anhydrate that is left over to moles.
4. Find the water-to-anhydrate mole _______ (just like finding an _________________ formula, but be careful: you can't multiply til whole! The mole ratio of the anhydrous salt must always be ____; only the number of waters can be a whole number greater than 1)
5. Use the mole ratio to write the formula.

Example: The following data were obtained when a sample of $BaCl_2$ hydrate was analyzed:

Mass of empty test tube	18.42 g
Mass of test tube and hydrate (before heating)	20.75 g
Mass of test tube and anhydrous salt (after heating)	20.41 g

a. Calculate the mass of water lost from the hydrate.

b. Calculate the percentage of water in the original sample.

c. Calculate the moles of water lost from the sample.

d. Calculate the moles of anhydrous salt remaining after the sample was heated.

e. Determine the formula for the hydrate.

Common Lab Errors when Determine the Formula of a Hydrate

Error	Effect on Calculated % H_2O
Excess ________________ caused the dehydrated sample to decompose. • Often times, a gas will be released during the decomposition	• Gas from the decomposition will be lost *as well as* the expected water loss from heating the hydrate. • The calculated % H_2O will be ______________than the actual % H_2O in the hydrate.
The dehydrated sample absorbed moisture from the air after heating (but before the mass is measured).	• Not all of the waters of hydration will be removed. • The calculated % H_2O will be ______________than the actual % H_2O in the hydrate.
The hydrate is not heated to "__________________ mass" • The hydrate should be heated multiple times and the mass measured each time, to ensure all of the water molecules have been driven off.	• Not all of the water molecules will have been driven off, so the remaining salt is not completely anhydrous. • The calculated % H_2O will be ______________than the actual % H_2O in the hydrate.

Practice with Hydrates: Thirst-Quenching!

Error Analysis Practice:

1. A student wants to experimentally determine the number of moles of water in one mole of $BeC_2O_4 \cdot 3\ H_2O$. After heating the sample and measuring the new mass, the student calculated the amount of water driven off. Using that value and her experimental data, she derives the formula of the hydrate as $BeC_2O_4 \cdot 2\ H_2O$. Provide a reasonable explanation for an error that might have caused this outcome and explain how this error affected the student's results.

2. Another student repeats the attempt to experimentally derive the empirical formula of of $BeC_2O_4 \cdot 3\ H_2O$. He decides to heat the sample multiple times over a higher heat to be certain to drive off all water from the hydrate sample. Using his data, this student calculated the formula of the hydrate to be $BeC_2O_4 \cdot 4\ H_2O$. Explain what error might have resulted in this outcome.

Multiple Choice Practice:

1. A sample of a hydrate of $CuSO_4$ with a mass of 250 grams was heated until all the water was removed. The sample was then weighed and found to have a mass of 160 grams. What is the formula for the hydrate?

   a. $CuSO_4 \cdot 10\ H_2O$ b. $CuSO_4 \cdot 7\ H_2O$ c. $CuSO_4 \cdot 5\ H_2O$ d. $CuSO_4 \cdot 2\ H_2O$

2. The anhydrous salt X_2CO_3 has a molar mass of 106 g/mol. A hydrated form of this salt is heated until all of the water is removed and it loses 54% of its mass. The formula of the hydrate is:

   b. $X_2CO_3 \cdot 7\ H_2O$ b. $X_2CO_3 \cdot 5\ H_2O$ c. $X_2CO_3 \cdot 3\ H_2O$ d. $X_2CO_3 \cdot H_2O$

Formula of an Unknown Hydrate Lab

Introduction

In this lab, you will be heating a hydrate of copper (II) sulfate ($CuSO_4 \cdot xH_2O$) to drive off the water. Masses are taken before heating to determine the mass of the original sample (the hydrate) and after heating to determine the mass of the anhydrous copper (II) sulfate ($CuSO_4$) remaining. The difference between these two masses is equal to the mass of the water lost.

$$CuSO_4 \cdot xH_2O$$

Potential Sources of Lab Error:

Heating time and temperature are critically important for this experiment.

- If not enough heat is applied, some water will remain attached to the copper (II) sulfate, producing a calculated mass percent of water that is lower than the actual value.
- If too much heat is applied, the anhydrous copper (II) sulfate, which has a greyish white color, can be decomposed to copper (II) sulfide, a black colored compound. This will produce a calculated mass percent of water that is higher than the actual value.

Important Note about Language:

A hydrate is a pure substance that contains water molecules embedded in its crystal structure. Thus, when water is removed from the hydrate through heating, water does not “evaporate” or “boil” away from the hydrate (those words signify a mere physical change). Instead, the correct phrasing to use states that the water is “driven off” from the hydrate and the solid that remains behind is called anhydrous, meaning "without water."

Chemicals

copper (II) sulfate unknown hydrate ($CuSO_4 \cdot xH_2O$)

Equipment

Hot plate

Crucible and cover

Crucible tongs

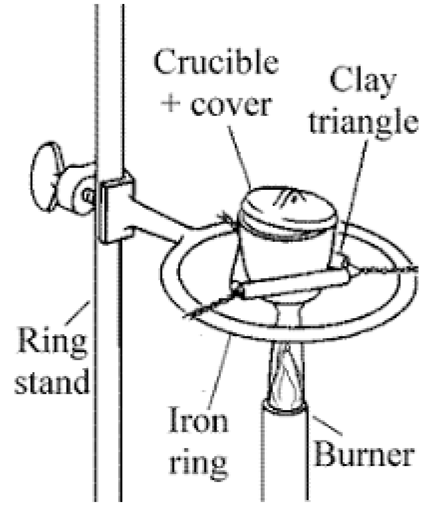
Stir rod

Balance

Procedure

A. Preparation of the Crucible

1. Obtain a clean, dry ceramic crucible and cover. Measure the mass of the crucible and cover (together) and **record in your data table**.



B. Heating the Hydrate

2. Add approximately 2.5−3.5 g of the copper (II) sulfate into the crucible (the cover should be on the balance as well). **Record the total mass of the sample, crucible and cover** in your data table.
3. Remove the crucible lid, then place the crucible and hydrate sample on the clay triangle over the Bunsen burner. Light the Bunsen burner, then heat the crucible and hydrate sample for about 12 minutes. (The crucible should NOT glow red; use the top of the flame and not the inner flame).
4. Turn off the Bunsen burner and CAREFULLY cover the crucible with the lid (don't touch the burning hot crucible!) Allow the closed crucible and sample to cool for 10 minutes, then measure the total mass of the sample, crucible and cover. **Record in your data table.** (WARNING: Do not measure the mass of a hot object on the balance!)

C. Heating to Constant Mass

5. To determine if all the waters of hydration have been driven off, reheat the sample for 5 minutes (WITHOUT cover), cool (with cover closed) and mass again. (WARNING: Do not measure the mass of a hot object on the balance!)
6. This should be repeated until the successive masses are constant within 0.03 g. When the masses are constant, record the lowest mass as the final mass.

D. Rehydration

7. ONCE YOU HAVE FINISHED HEATING TO CONSTANT MASS, use a pipette to transfer a few drops of water to your anhydrous sample to "rehydrate" it. Notice what happens to the color: the anhydrous and hydrated samples often have different colors.
8. Place the cooled $CuSO_4$ in the waste container provided. Clean all equipment and return to the specified location.

Data Table: Percentage of Water in a Hydrate

	Mass (g)
Mass of crucible and cover	
Mass of hydrate sample, crucible and cover	
Mass of dehydrated sample, crucible and cover after 1st heating	
Mass of dehydrated sample, crucible and cover after 2nd heating	
Mass of dehydrated sample, crucible and cover after 3rd heating (if necessary)	

Calculations (Show all work!)

1. Calculate the mass of your hydrate sample.

2. Calculate the mass of water driven off from the hydrate.

3. Calculate the mass of copper (II) sulfate *without water* in your original sample.

4. Calculate the percentage of water in the original sample.

5. Calculate the moles of copper (II) sulfate, $CuSO_4$, and the moles of water in your original sample.

6. Determine the formula of the unknown hydrate; that is, determine the value of "x" in the formula ($CuSO_4 \cdot xH_2O$).

Ask your teacher for the correct formula of the hydrated compound: ______________________________

7. Using the correct hydrate formula, calculate the actual (accepted) percentage of water in the hydrate.

8. Calculate your percent error (on the percentage of water in the original sample, #4).

Analysis

1. In this experiment, overheating causes a **high** calculated percent value for water. Why is the high reading obtained? (*Hint: overheating causes copper (II) sulfate, $CuSO_4$, to decompose into copper (II) sulfide, CuS. What is lost from the $CuSO_4$ in this process? Where did it go?*)

2. Suppose that you did not completely convert the hydrate to the anhydrous compound. Explain how this would affect the calculated percent by mass of water in the compound and the hydrate formula you determined.

AP Free Response

or How I Learned to Stop Worrying and Love AP Chemistry

The AP Chemistry test absolutely love love loves to give you very long questions, where many of the answers rely on answers determined in earlier parts of the question. This has stymied many a determined AP chemistry student in the past and can significantly lower your overall score. You need to know how to best show AP chemistry graders (and me!) all of the chemistry you know as quickly as possible, which means not getting stuck on one tiny part of a problem.

Question: What do you do when you know exactly how to answer parts d. – g., but you have no idea how to get the answer to part c. (which you need for the rest of the question)?

Answer: Make up an answer! For reals. The AP Chemistry scorers will give you FULL POINTS if you correctly show your work and chemistry knowledge on a question part, *even* if your final answer is wrong because you started with an incorrect value from a previous wrong answer. (The wrong calculation will still be counted wrong, as it should be.)

AP Free Response Sample Question (10 points)

1. In the laboratory, a sample of pure nickel was placed in a clean, dry, weighted crucible. The crucible was heated so that the nickel would react with the oxygen in the air. After the reaction appeared complete, the crucible was allowed to cool and the mass was determined. The crucible was reheated and allowed to cool. Its mass was then determined again to be certain that the reaction was complete. The following data was collected during the experiment. Use this data to answer the questions below.

Data Table 1: Calculation of Empirical Formula of Nickel Oxide

Mass of crucible	30.02 g
Mass of nickel and crucible	31.07 g
Mass of nickel oxide and crucible	31.36 g

a. What is the mass of nickel used? [1 point]

b. What is the mass of nickel oxide produced? [1 point]

c. What mass of oxygen reacted? [1 point]

d. Based on your calculations, what is the empirical formula for the nickel oxide? [2 points]

e. Based on your knowledge of reduced layered perovskite synthesis, what is the molecular weight of the compound? [2 points]

f. Determine the molecular formula of the compound. [1 point]

g. While waiting for the sample to cool, the student begins to clean up the lab station using a nearby sink. If, unknown to the student, some water splashed into the crucible, how would this affect the calculated empirical formula? [1 point]

Combustion Analysis

<u>Combustion Analysis</u>: Technique used to obtain the empirical formula of a ____________________

➔ Remember a standard (unbalanced) combustion reaction? (This formula is unbalanced!)

$$C_xH_yO_z + O_2 \rightarrow \_\_\_\_\_ + \_\_\_\_\_\_$$

How to Solve a Combustion Analysis Problem

1. Convert mass of CO_2 and mass of H_2O to ___________ of each compound.
2. Convert moles of CO_2 to moles of _______________, and moles of H_2O to moles of _______________.
3. **<u>If</u>** the compound contains something which is _________ C or H, find its mass by subtraction, and convert the mass to moles.
4. Now you have mole numbers! Complete the _____________ formula calculation (divide by small, multiply til whole).

Practice:

1. Upon combustion, a 0.8233 g sample of a compound containing only carbon, hydrogen, and oxygen produces 2.445 g CO_2 and 0.6003 g H_2O. What is the empirical formula of the compound?

Practice Makes Perfect!

2. Combustion analysis determined that a compound containing only carbon and hydrogen produces 1.83 g CO_2 and 0.901 g H_2O. Find the empirical formula of the compound.

3. When the <u>unbalanced</u> reaction below occurs at STP, 1.5 L of CO_2 and 1.0 L of H_2O are created. What is the empirical formula of the hydrocarbon?

$$C_xH_y(g) + O_2 \rightarrow CO_2(g) + H_2O(g)$$

a. CH_2 b. C_2H_3 c. C_2H_5 d. C_3H_4

4. Combustion analysis of 0.800 g of an unknown hydrocarbon yields 26 g CO_2 and 7.8 g H_2O. What is the formula of the hydrocarbon?

a. CH_2 b. C_2H_3 c. C_2H_5 d. C_3H_4

Unit 1 MC Practice

1. 2.500 grams of $MgSO_4 \cdot x\, H_2O$, a hydrated salt with an unknown water content, is dried in an oven to constant mass until all water has been removed. After drying, the anhydrous salt has a mass of 1.221 grams. How many moles of water are present per mole of hydrated magnesium sulfate? (FW of $MgSO_4$ = 120 g/mol)

   a. 1 b. 3 c. 5 d. 7

2. Naturally occurring rubidium consists of just two isotopes. One of the isotopes consists of atoms having a mass of 84.912 amu, the other of 86.901 amu. What is the percent natural abundance of the heavier isotope?

   a. 15% b. 28% c. 37% d. 72%

3. When hafnium metal is heated in an atmosphere of chlorine gas, the product of the reaction is found to contain 62.2 percent Hf by mass and 37.4 percent Cl by mass. What is the empirical formula for this compound?

   a. Hf_2Cl_3 b. $HfCl_2$ c. $HfCl_3$ d. $HfCl_4$

4. Which of the following represents the correct method for converting 11.0 g of copper metal to the equivalent number of copper atoms?

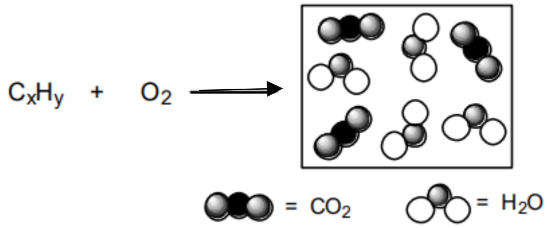
   a. $11\left(\frac{1}{63.55}\right)\left(\frac{6.02\times10^{23}}{1}\right)$

   b. $11\left(\frac{1}{63.55}\right)$

   c. $11\left(\frac{1}{63.55}\right)\left(\frac{63.55}{6.02\times10^{23}}\right)$

   d. $11\left(\frac{63.55}{1}\right)\left(\frac{6.02\times10^{23}}{1}\right)$

5. A hydrocarbon of unknown formula C_xH_y was submitted to combustion analysis: the results are shown in the diagram below. What is the empirical formula of the compound?



a. CH   b. CH_2 c. C_3H_4 d. C_3H_8

6. How many grams of calcium nitrate, $Ca(NO_3)_2$, contains 48 grams of oxygen atoms?

a. 82 g   b. 120 g   c. 190 g   d. 320 g

7. If 200.0 mL of 0.60 *M* $MgCl_2$(aq) is added to 400. mL of distilled water, what is the concentration of Mg^{2+}(aq) in the resulting solution? (Assume volumes are additive.)

a. 0.20 *M*   b. 0.30 *M*   c. 0.40 *M*   d. 1.2 *M*

8. Which set of molecules is in order from lowest to highest percent mass of oxygen?

a. $CH_3OH < CH_3CH_2OH < CH_3CH_2CH_2OH < HOCH_2CH_2OH$

b. $HOCH_2CH_2OH < CH_3OH < CH_3CH_2OH < CH_3CH_2CH_2OH$

c. $CH_3CH_2CH_2OH < CH_3CH_2OH < CH_3OH < HOCH_2CH_2OH$

d. $CH_3CH_2CH_2OH < CH_3CH_2OH < HOCH_2CH_2OH < CH_3OH$

9. Given that there are two naturally occurring isotopes of gallium, ^{69}Ga and ^{71}Ga, the natural abundance of the ^{71}Ga isotope must be approximately:

a. 25%   b. 40%   c. 50%   d. 71%

10. What pressure (in atm) would be exerted by 76 g of fluorine gas, $F_2(g)$, in a 1.50 liter vessel at −37°C?

a. 4.1 atm b. 8.2 atm c. 26 atm d. 84 atm

11. Which of the following statements is true?

I. The molar mass of $CaCO_3$ is 100.1 g mol^{-1}.
II. 50 g of $CaCO_3$ contains about 9×10^{23} oxygen atoms.
III. A 200 g sample of $CaCO_3$ contains about 2 moles of $CaCO_3$.

a. I only b. II only c. I and III only d. I, II, and III

12. In which of the following compounds is the mass ratio of element X to oxygen closest to 2.5 to 1? (The molar mass of X is 40.0 g/mol.)

a. X_3O_2 b. X_2O c. XO_2 d. XO

13. If 87 grams of potassium sulfate, K_2SO_4, (molar mass 174 g/mol) is dissolved in enough water to make 250 mL of solution, what are the concentrations of the potassium and the sulfate ions?

	$[K^+]$	$[SO_4^{2-}]$
a.	0.020 *M*	0.020 *M*
b.	1.0 *M*	2.0 *M*
c.	2.0 *M*	1.0 *M*
d.	4.0 *M*	2.0 *M*

14. A compound contains 30.% sulfur and 70.% fluorine by mass. The empirical formula of the compound is:

a. SF b. SF_2 c. SF_3 d. SF_4

15. How many carbon atoms are contained in 2.8 g of C_2H_4?

a. 6.0×10^{22} b. 1.2×10^{23} c. 3.0×10^{23} d. 6.0×10^{23}

16. Which of the following statements is(are) false?

    I. The % by mass of each element in a compound depends on the amount of the compound.
    II. The mass of each element in a compound depends on the amount of the compound.
    III. The % by mass of each element in a compound depends on the amount of element present in the compound.

    a. I only b. II and III only c. I and II only d. I, II, and III

17. Guanidin, $HNC(NH_2)_2$, is a fertilizer. What is the percent by mass of nitrogen in the fertilizer?

    a. 45% b. 55% c. 65% d. 71%

18. How many atoms are in one mole of CH_3OH?

    a. 6.0 b. 6.0×10^{23} c. 1.2×10^{24} d. 3.6×10^{24}

19. What is the weight of $MgCO_3$ (formula weight 84.3 g/mol) found in 100. mL of a 5.0 M solution?

    a. 42 g b. 84 g c. 420 g d. 840 g

20. An oxide of lead contains 90.65% Pb by weight. The empirical formula is:

    a. Pb b. PbO c. Pb_3O_4 d. PbO_2

21. The mass in grams of 2.6×10^{22} chlorine atoms is:

    a. 0.76 g b. 1.5 g c. 3.2 g d. 4.4 g

22. Identify the INCORRECT statement.

a. Helium in a balloon: an element
b. Paint: a mixture
c. Tap water: a compound
d. Mercury in a barometer: an element

23. Which one of the following samples contains the most **atoms**?

a. 1 mol of UF_6(g)
b. 1 mol of He(g)
c. 1 mol of CH_3COCH_3(l)
d. all contain the same number of atoms

24. Combustion analysis of 2.400 grams of toluene, an organic solvent, yields 8.023 g CO_2 and 1.877 g H_2O. What is the simplest formula for toluene?

a. C_4H_7
b. C_7H_4
c. C_7H_8
d. C_8H_7

25. The neutral atoms of all of the isotopes of the same element have:

a. different numbers of protons
b. equal numbers of neutrons
c. the same number of electrons
d. the same mass number

26. Which one of the following samples contains the most **molecules**?

a. 1 mol of UF_6(g)
b. 1 mol of He(g)
c. 1 mol of CH_3COCH_3(l)
d. all contain the same number of atoms

27. Analysis of a sample of a covalent compound showed that it contained 14.4% hydrogen and 85.6% carbon by mass. What is the empirical formula of the compound?

a. CH
b. CH_2
c. CH_3
d. C_2H_3

28. What mass of cerussite, $PbCO_3$, would contain 35.0 grams of lead?

a. 27.1 g b. 35.6 g c. 45.1 g d. 51.7 g

29. The simplest formula for an oxide of nitrogen that is 36.8 percent nitrogen by weight is:

a. N_2O b. NO c. N_2O_3 d. NO_2

30. How many aluminum atoms are there in 3.50 g of Al_2O_3?

a. 2.07×10^{22} b. 2.07×10^{23} c. 4.13×10^{22} d. 4.13×10^{23}

31. The density of chlorine gas at STP, in grams per liter, is approximately:

a. 1.3 b. 3.2 c. 4.5 d. 6.2

32. A container with volume 71.9 mL contains water vapor at a pressure of 10.4 atm and a temperature of 465^{o}C. How many grams of the gas are in the container?

a. 0.129 g b. 0.222 g c. 0.363 g d. 0.421 g

33. What is the atomic weight of a hypothetical element consisting of two isotopes, one with mass = 64.23 amu (26.0%), and one with mass = 65.32 amu?

a. 64.8 amu b. 64.8 amu c. 65.0 amu d. 65.3 amu

34. The mass of element X found in 1.00 mole of each of four different compounds is 28.0 g, 42.0 g, 56.0 g, and 70.0 g, respectively. The possible atomic weight of X is:

a. 8.00 b. 14.0 c. 28.0 d. 38.0

35. Consider the species, ^{72}Zn, ^{75}As, and ^{74}Ge. These species have:

a. the same number of electrons
b. the same number of neutrons
c. the same number of protons
d. the same mass number

36. Which of the following includes **all** of the following that are **chemical** changes and not physical changes?

I. freezing of water
II. dropping a piece of iron into hydrochloric acid (H_2 is produced)
III. burning a piece of wood
IV. emission of light by a kerosene lamp

a. I and IV only b. II and III only c. II, III, and IV only d. I, II, and III only

AP Free Response Practice #1 [10 points, 1991 Form B]

1. Answer the following questions about BeC_2O_4 and its hydrate.
   a. Calculate the mass percent of carbon in the hydrated solid with the formula $BeC_2O_4 \cdot 3\ H_2O$. (2 points)

   b. When heated to 220.°C, $BeC_2O_4 \cdot 3\ H_2O$ dehydrates completely as represented below:

$$BeC_2O_4 \cdot 3\ H_2O(s) \rightarrow BeC_2O_4(s) + 3\ H_2O(g)$$

   If 3.21 g of $BeC_2O_4 \cdot 3\ H_2O$ is heated to 220.°C, calculate each of the following:

      i. The mass of BeC_2O_4 formed. (1 point)

      ii. The volume of $H_2O(g)$ released, measured at 220.°C and 735 mmHg. (2 points)

A student repeats the dehydration from part (b) in an attempt to experimentally determine the number of moles of water in one mole of $BeC_2O_4 \cdot 3\ H_2O$. The student collects the data shown in the table below.

Mass of empty crucible	36.48 g
Initial mass of sample and crucible	39.69 g
Mass of sample and crucible after first heating	38.82 g

   c. Use the data above to:

      i. Calculate the total number of moles of water lost when the sample was heated. (1 point)

ii. Determine the formula of the hydrated compound. (2 points)

d. Is the student's experimentally determined waters of hydration greater than, less than, or equal to the waters of hydration in the accepted formula? Provide a reasonable explanation for error and how this error affected the student's results. (2 points)

AP Free Response Practice #2 [10 points]

2. Lysine is an amino acid which has the following elemental composition: C, H, O, N. It is found in the protein of foods such as beans, cheese, yogurt, meat, milk, brewer's yeast, wheat germ, and other animal proteins. The average 70 kg human needs 800 – 3,000 mg of lysine daily. In one experiment, 2.175 g of lysine was combusted to produce 3.94 g of CO_2 and 1.89 g H_2O. The molar mass of lysine is approximately 150 g/mol.

   a) Determine the mass, in grams, of each of the following in the 2.175 g sample of lysine.

      i. carbon [*1 point*]

      ii. hydrogen [*1 point*]

   b) In a separate experiment, 1.873 g of lysine was burned to produce 0.436 g of NH_3.

      i. Determine the mass, in grams, of N in the 1.873 g sample of lysine. [*1 point*]

      ii. Determine mass percent of each element in lysine. [*2 points*]

c) Determine the mass, in grams, of O in the original 2.175 g sample of lysine. [*1 point*]

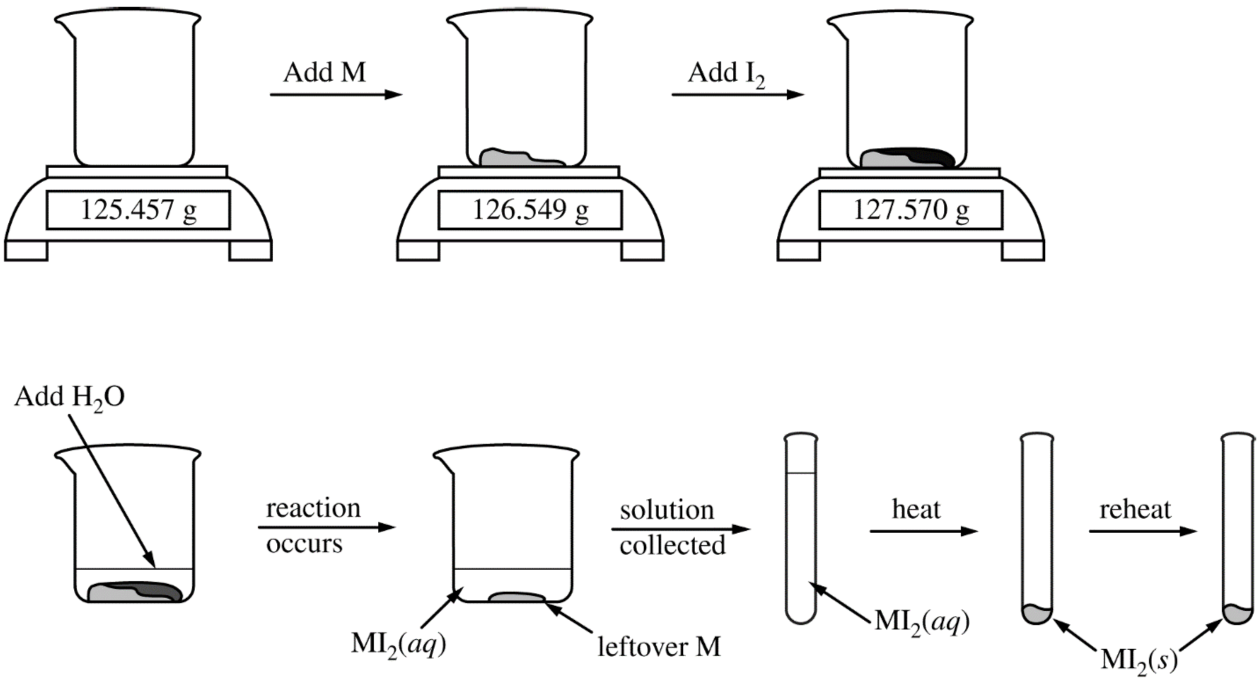
d) Using information derived from the provided data, determine the empirical formula of lysine. [*3 points*]

e) Determine the molecular formula of lysine. [*1 point*]

AP Free Response Practice #3 (2016, modified) [10 points]

$$M + I_2 \rightarrow MI_2$$

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 MI_2, as represented by the equation above. The reaction proceeds until all of the I_2 is consumed. The $MI_2(aq)$ 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.



Data for Unknown Metal Lab	
Mass of beaker	125.457 g
Mass of beaker + metal M	126.549 g
Mass of beaker + metal M + I_2	127.570 g
Mass of MI_2, first weighing	1.284 g
Mass of MI_2, second weighing	1.284 g

a) Given that the metal M is in excess, calculate the number of moles of I_2 that reacted. [*2 points*]

b) Determine the following for the unknown metal M.

i. Calculate the molar mass of the unknown metal M. [*2 points*]

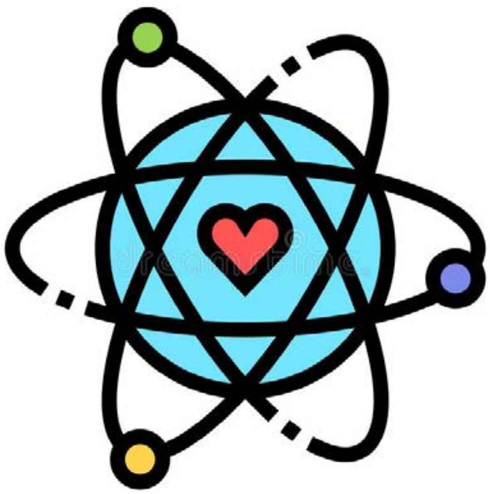
ii. What is the most likely identity of the unknown metal M? [*1 point*]

iii. Calculate the number of moles of unknown metal M that reacted. [*2 points*]

c) Provide a calculation to confirm the empirical formula of the compound MI_2 based on the data shown. [*2 points*]

d) If the student failed to heat to constant mass, would the calculated molar mass of the unknown metal M be greater than, less than, or equal to the actual molar mass? Explain. [*1 point*]

AP Chemistry FTW!



Unit 2:

Stoichiometry & Reactions

AP Chemistry Unit 2 Objectives

BIG IDEA 1 - The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

- Enduring Understanding 1.D: Atoms are so small that they are difficult to study directly; atomic models are constructed to explain experimental data on collections of atoms.
- Enduring Understanding 1.E: Atoms are conserved in physical and chemical processes.

BIG IDEA 2 - Chemical and physical properties of materials can be explained by the structure and arrangement of atoms, ions, molecules and the forces between them.

- Enduring Understanding 2.A: Matter can be described by its physical properties. The physical properties of a substance generally depend on the spacing between the particles (atoms, molecules, ions) that make up the substance and the forces of attraction among them.
- Enduring Understanding 2.B: Forces of attraction between particles (including the noble gases and also different parts of large molecules) are important in determining many macroscopic properties of a substance, including how observable physical state changes with temperature.
- Enduring Understanding 2.C: The strong electrostatic forces of attraction holding atoms together in a unit are called chemical bonds.

BIG IDEA 3 - Changes in matter involve the rearrangement and/or the reorganization of atoms and /or the transfer of electrons.

- Enduring Understanding 3.A: Chemical changes are represented by a balanced chemical equation that identifies the ratios with which reactants react and products form.
- Enduring Understanding 3.B: Chemical reactions can be classified by considering what the reactants are, what the products are, or how they change from one into the other. Classes of chemical reactions include synthesis, decomposition, acid-base, and oxidation-reduction reactions.
- Enduring Understanding 3.C: Chemical and physical transformation may be observed in several ways and typically involve a change in energy.

BIG IDEA 6 - Any bond or intermolecular attraction that can be formed can be broken. These processes are in a dynamic competition, sensitive to initial conditions and external perturbations.

- Enduring Understanding 6.A: Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal.

Physical and Chemical Properties and Changes

Intensive & Extensive Properties

Extensive property: a characteristic that depends on the ______________ of matter present (how much you have).

1. Mass **2.** Volume **3.** Length (width, height)

Intensive property: a characteristic that does _______ depend on the amount of matter present.

1. Color **2.** Density **3.** Malleability **4.** Ductility **5.** Luster

6. Odor **7.** Melting Pt. **8.** Boiling Pt. **9.** Conductivity

Physical Properties/Changes

Physical property: a characteristic that can be observed or measured without changing the ___________________ of the substance.

1. Color **2.** Melting point **3.** Malleability **4.** Ductility **5.** Density

6. Solubility (ability to dissolve) **7.** Mass **8.** Volume **9.** Viscosity

Physical change: a change when chemical bonds within compounds are _________ broken in the reactant and no new chemical bonds are formed in the products. The _______________ of the material does NOT change.

1. Changes of _____________. *Example:* **$Fe(s) \rightarrow Fe(l)$** or **$CO_2(s) \rightarrow CO_2(g)$**

2. Changes in _____________.

Chemical Properties/Changes

Chemical property: a characteristic that can ONLY be observed or measured by subjecting it to a process that might **change** the ___________________ of the substance.

1. Reactivity **2.** Flammability **3.** Toxicity **4.** pH (how acidic or basic it is) **5.** Ability to Ferment
6. Ability to Oxidize, for example: rusting (___) or tarnishing (___)

Chemical change: a change when chemical bonds within the compound ARE broken in the reactants and ______ chemical bonds are formed in the products. The ______________________ of the material changes.

Potential Indications of Chemical Change

1. Energy change (emission or absorption of heat or light)
2. Formation of a gas
3. Formation of a precipitate
4. Formation of an odor
5. Color change

Chemical Reactions can be described in several ways: you must be able to covert from any of these forms to another

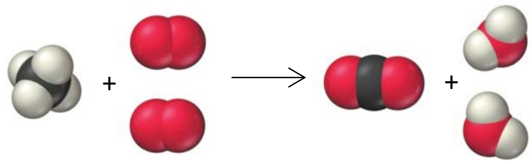
1. In a __________________ (equation)

   $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

2. In a ________________

   Methane reacts with oxygen gas to form carbon dioxide and water vapor.

3. In a picture (_____________________ representation)



Symbols used in chemical equations:

$\longrightarrow$	forms, produces, yields
(s)	solid state
(l)	liquid state
(g)	gaseous state
(aq)	aqueous solution (dissolved in water).
$\xrightarrow{heat}$ or $\xrightarrow{\Delta}$	Reactants are heated.
$\xrightarrow{MnO_2}$	Formula of catalyst used to speed up reaction. The catalyst does NOT undergo permanent chemical change itself.

Steps in a Chemical Reaction: Atoms are _______ created or destroyed! (Law of Conservation of __________)

1. Reactant bonds _____________
2. Atoms rearranged.
3. ______ bonds formed in products!
4. Mass of all Reactants = Mass of all Products

Obeying the Law (of Conservation of Mass): Balancing Chemical Equations!

1. Word equations must be translated into a formula equation.

   Do not forget nomenclature:
   - Covalent compounds: prefixes
   - Ionic compounds: balance charges
   - Diatomic elements: **Br_2, I_2, N_2, Cl_2, H_2, O_2, F_2** ← But only when **alone** (i.e. not bonded to another element)
2. Balance 1 atom at a time by writing ____________________ in front of compound.
   - Do NOT ever change subscripts!
3. Polyatomic ions treated as a single unit (_____they stay together on both sides)
4. If an equation has C, H, and O's balance them in that respective order.
5. Make sure the coefficients are in the __________________ whole number ratio (reduce).

Practice with Balancing Reactions!

1. $\_\_\_Fe(s) + \_\_\_O_2(g) \rightarrow \_\_\_Fe_2O_3(s)$

2. $\_\_\_C_3H_8(g) + \_\_\_O_2(g) \rightarrow \_\_\_H_2O(l) + \_\_\_CO_2(g)$

3. $\_\_\_C_6H_{14}(g) + \_\_\_O_2(g) \rightarrow \_\_\_H_2O(l) + \_\_\_CO_2(g)$

4. Aluminum metal and molecular chlorine undergo a synthesis reaction. Write and balance the equation for this reaction.

5. An impure sample containing lead (II) nitrate, $Pb(NO_3)_2$, and an inert material is added to sodium hydroxide, NaOH. The two aqueous reactants undergo a chemical change, producing aqueous $Pb(OH)_2$, and one other compound that remains dissolved in the solution. Write and balance the equation for this reaction.

6. An aqueous reaction of hydrochloric acid, HCl, and K_2SO_3 undergoes a gas evolution reaction to form liquid water, SO_2 vapor, and aqueous KCl. Write and balance the equation for this reaction.

Unit 2 Part 1B: Types of Chemical Reactions

1. Synthesis Reactions

$$A + B \rightarrow AB$$

- Also called combination reactions
- Two ________________, or _________________ combine to make one compound.
- We can predict the product if the reactants are two elements.

Example: ____Mg (s) + ____N_2 (g) →

Now you try: ____Al (s) + ____O_2 (g) →

2. Decomposition Reactions

$$AB \rightarrow A + B$$

- Decompose = _____________________
- One compound (reactant) _____________________ into two or more elements or compounds.
- Usually requires energy
- We can predict the products if the reactants break apart into two elements.

Example: ____ H_2O (g) →

Now you try: ____ NaCl (s) →

3. Single Replacement

$$A + BC \rightarrow AC + B$$

- Also referred to as ________________ displacement
- One element _____________________another
- Reactants must be an ____________________ and a _____________________
- Products will be a different __________________ and a different _____________________
- We can predict the products if we know the charge on the cation in the compound being formed:

Example: ____$Cu(NO_3)_2$(*aq*) + ____Ag(*s*) →

Now you try: ____Zn(s) + ____$PbCl_3$(aq) →

4. **Double Replacement**

- Two things ___________________ each other.
- Reactants must be two _______________________.
- Usually in an ________________________ solution
- Two types of DR reactions: precipitate & acid/base
- We can always predict the products of a DR reaction:

AB + CD → AD + CB

Example: ____$Cu(NO_3)_2(aq)$ + ____$Ag_2CO_3(aq)$ →

Now you try: ____$Pb(ClO_3)_2(aq)$ + ____$Al_3(SO_4)_2\ (aq)$ →

5. **Combustion**

CH_4 + O_2 → CO_2 + H_2O

- A reaction in which a compound (often carbon) reacts with __________________ _________ and usually produces _________________________ and ______________________.
- We can always predict the products of a combustion reaction:

Example: ____$C_6H_{12}(g)$ + ____$O_2(g)$ →

Now you try: ____$C_8H_{18}(g)$ + ____$O_2(g)$ →

How to recognize each type of reaction

1. Look at the **reactants**:
   - element + element = ______________________
   - compound = ______________________
   - element + compound = ______________________
   - compound + compound = ______________________
   - C_xH_y or $C_xH_yO_z$ + O_2 = ______________________
2. Look at the **products**: CO_2 + H_2O = ______________________

Practice, Yum!

Directions: Balance the following equations by placing coefficients in the blanks. Classify the reaction.

Type of Reaction:

1. ____$Zn(s)$ + ____$HCl(aq)$ → ____$ZnCl_2(aq)$ + ____$H_2(g)$ ____________

2. ____$C_3H_8(g)$ + ____$O_2(g)$ → ____$CO_2(g)$ + ____$H_2O(g)$ ____________

3. ____$KClO_3(s)$ → ____$KCl(s)$ + ____$O_2(g)$ ____________

4. ____$BaCl_2(aq)$ + ____$KIO_3(aq)$ → ____$Ba(IO_3)_2(s)$ + ____$KCl(aq)$ ____________

5. ____$Fe(s)$ + ____$H_2O(g)$ → ____$Fe_3O_4(s)$ + ____$H_2(g)$ ____________

6. ____$Fe(s)$ + ____$O_2(g)$ → ____$Fe_2O_3(s)$ ____________

7. ____$C(s)$ + ____$H_2(g)$ + ____$O_2(g)$ → ____$C_2H_6O(s)$ ____________

8. ____$O_2(g)$ + ____$C_6H_{14}(g)$ → ____$H_2O(g)$ + ____$CO_2(g)$ ____________

9. ____$HNO_3(aq)$ + ____$Mg(OH)_2(aq)$ → ____$H_2O(l)$ + ____$Mg(NO_3)_2(aq)$ ____________

Practice Predicting Products!

Directions:

1. Determine type: single replacement (SR), double replacement (DR), synthesis (S), decomposition (D), or combustion (C). Write the type on the line to the left of the reaction.
2. Predict the products and write the correct chemical formula for each product. (Don't worry about states of matter!)
3. Balance each reaction using the correct coefficients.

Type?

_______ 1. ____ $AgNO_3$ + ____ $CaCl_2 \rightarrow$

_______ 2. ____ Al + ____ $Cl_2 \rightarrow$

_______ 3. ____ C_2H_4 + ____ $O_2 \rightarrow$

_______ 4. ____ Mg + ____ $BaI_2 \rightarrow$

_______ 5. ____ NaOH + ____ HCl $\rightarrow$

_______ 6. ____ Na + ____ $FeBr_3 \rightarrow$

_______ 7. ____ $PbSO_4$ + ____ $AgNO_3 \rightarrow$

_______ 8. ____ $KMnO_4$ + ____ $ZnCl_2 \rightarrow$

_______ 9. ____ O_2 + ____ $C_5H_{12}O_2 \rightarrow$

_______ 10. ____ $CuCl_3 \rightarrow$

Oxidation-Reduction (Redox) Review

Oxidation–reduction (redox) reactions: where electrons are transferred from one atom to another.

- If a substance accepts an electron, it is____________________.
- If a substance loses an electron, it is ____________________.
- Electrons are always transferred from the species that is oxidized to the species that is reduced.

➔ Reduction (gain in electrons) + Oxidation (loss of electrons) = Redox!

Two great mnemonics!

1. ______ _______: **O**xidation **I**s **L**oss (OIL) and **R**eduction **I**s **G**ain (RIG)
2. ______ **goes** ______: A species **loses electrons** when **oxidized**, and **gains electrons** when **reduced**.

***Almost* all reaction types (except double replacement) are redox.** However, that statement is not always true – the only way to be certain if a reaction is redox is to determine if any species has gained or lost electrons by looking for a ____________ in oxidation state/number.

- If a chemical reaction _________ have a species which changes oxidation number: ________, **it's redox!**

Oxidation Numbers/ Oxidation States

Oxidation states are imaginary charges assigned based on a set of rules simply used to determine _________________ flow.

→ Even though they look like them, oxidation states are **NOT** ___________ charges.

→ Oxidation numbers can be assigned to each atom in an element, ion, or compound…whether the compound is _______________ or _____________________!

How to Assign Oxidation States (aka Oxidation Numbers)

Free Elements or Monatomic Ions

1. Free elements = 0
2. Monoatomic ions = their charge

Examples	Free Elements			Monatomic Ions	
	$Fe(s)$	$Br_2(l)$	$O_3(g)$	Au^{3+}	S^{2-}
Oxidation #	0	0	0	+3	−2

Atoms in a Compound

3. All atoms in a neutral compound add up to 0.
4. All atoms in a polyatomic ion add up to the ion's charge.

	In order of priority	Oxidation State
Metals	Group 1A metals	+1
	Group 2A metals	+2
	Group 3A metals	+3
Non-metals	**fluorine**	−1
	hydrogen	+1
	oxygen	−2
	Group 7A	−1
	Group 6A	−2
	Group 5A	−3

Helpful hints:

- Group 4A (the carbon family) and transition metals are NOT listed – you will ALWAYS have to solve for them.
- Coefficients do NOT affect oxidation numbers.
- You can split up ionic compounds with a polyatomic and use the ion's overall charge to solve for the oxidation states of each element in the ion

Easy practice: Elements, Ions, and Simple Ionic Compounds

1. $Na^+(aq)$
2. $H_2(g)$
3. $Na_2S(s)$

Medium practice: More compounds and polyatomic ions

4. $C_2H_4(g)$
5. $Cr_2O_7^{2-}(aq)$

Conflict-resolution practice: What do you do if the rules don't agree?

6. $SO_3^{2-}(aq)$
7. $H_2O_2(l)$

Fun practice! ☺

8. $Fe(ClO_4)_3(s)$
9. $H_2SeO_3(aq)$

How to Identify What is Oxidized or Reduced in a Reaction

Once you have identified a redox reaction by the change in oxidation state, now you can tell what was oxidized or reduced!

a. A substance that has the element that has been ________________ (LOST electrons) will have an oxidation number that becomes more ________________ (or less negative).

b. A substance that has the element that has been ________________ (GAINED electrons) will have an oxidation number that becomes more ________________ (or less positive).

Balancing Redox Reactions: We split redox reactions into two separate ____________________

- The oxidation half-reaction has electrons as a ________________.
- The reduction half-reaction has electrons as a ________________.

Oxidation Half-Reaction	**Reduction Half-Reaction**
$\mathrm{Zn}(s) \rightarrow \mathrm{Zn}^{2+}(aq) + 2\,e^-$	$\mathrm{Fe}^{2+}(aq) + 2\,e^- \rightarrow \mathrm{Fe}(s)$

Practice!

1. Using oxidation numbers, identify what was oxidized and reduced in each reaction below.

   a) $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

   b) $2\ \mathrm{H_2O}(l) + 4\ \mathrm{MnO_4^-}(aq) + 3\ \mathrm{ClO_2^-}(aq) \rightarrow 4\ \mathrm{MnO_2}(aq) + 3\ \mathrm{ClO_4^-}(aq) + 4\ \mathrm{OH^-}(aq)$

2. In the reaction below, a piece of solid nickel is added to a solution of potassium dichromate.

$$14\ \mathrm{H^+}(aq) + \mathrm{Cr_2O_7^{2-}}(aq) + 3\ \mathrm{Ni}(s) \rightarrow 2\ \mathrm{Cr^{3+}}(aq) + 3\ \mathrm{Ni^{2+}}(aq) + 7\ \mathrm{H_2O}(l)$$

Which species is being oxidized and which is being reduced?

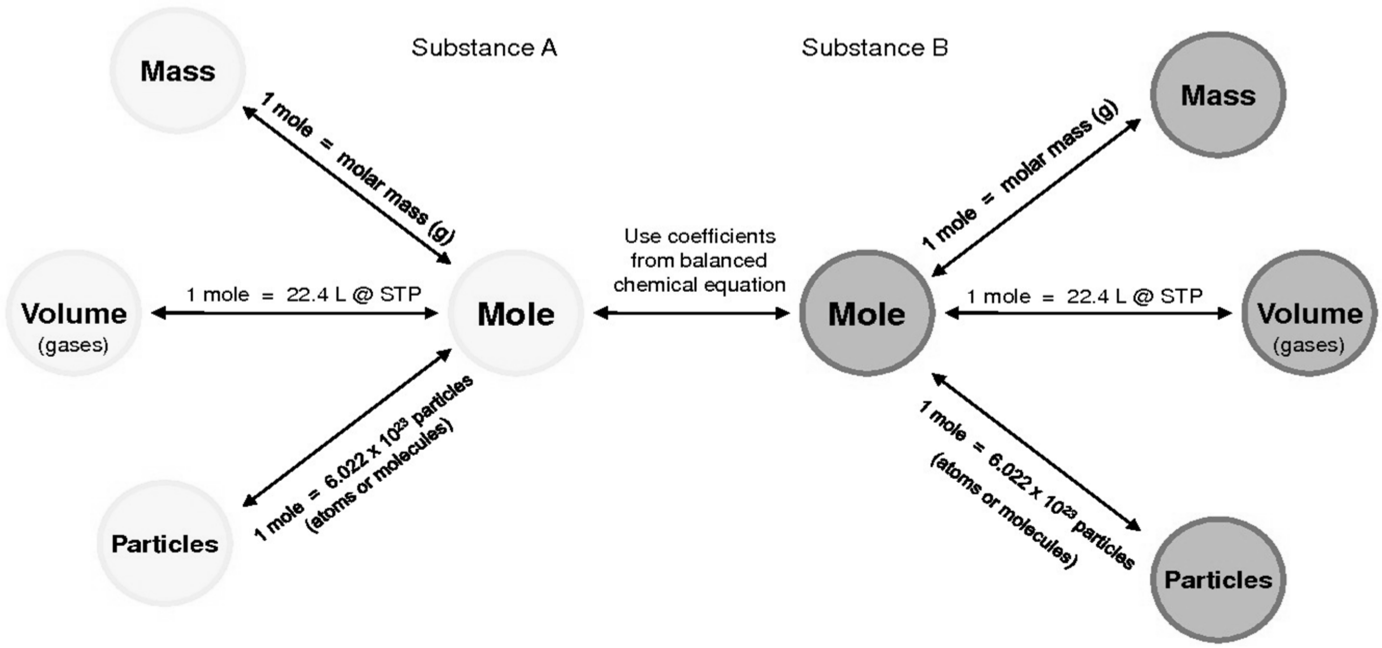
	Oxidized	Reduced
a.	$\mathrm{Cr_2O_7^{2-}}(aq)$	$\mathrm{Ni}(s)$
b.	$\mathrm{Cr^{3+}}(aq)$	$\mathrm{Ni^{2+}}(aq)$
c.	$\mathrm{Ni}(s)$	$\mathrm{Cr_2O_7^{2-}}(aq)$
d.	$\mathrm{Ni^{2+}}(aq)$	$\mathrm{Cr^{3+}}(aq)$

Stoichiometry

Stoichiometry involves the study of the relationships between _____________________ and _____________________ in a chemical reaction.

- The equation that is describing the reaction must be balanced.
- The coefficients in a chemical equation can describe individual _____________________ or amounts in _______________ of the products and reactants.

Mole Ratios and Molar Mass as Conversion Factors: A ______________ ratio is a conversion factor that relates the amounts in moles of any two chemical species involved in a chemical reaction.



<u>Basic Stoichiometry:</u> Gram/Mole/Particle Conversions

1. How many total moles of product are produced if 3.25 g of LiOH reacts with excess H_2SO_4?

   ____$LiOH(aq)$ + ___ $H_2SO_4(aq)$ → ___$H_2O(l)$ + ___$Li_2SO_4(aq)$

2. Consider the following reaction: $Ba(NO_3)_2\,(aq) + 2\,KOH\,(aq) \rightarrow Ba(OH)_2\,(aq) + 2\,KNO_3\,(aq)$

   a. What mass of barium hydroxide could be formed from 1.0 kg of barium nitrate?

   b. How many hydroxide ions would be in this sample?

Gas Stoichiometry: At STP and non-STP

3. 0.500 L of $H_2(g)$ reacts with excess $O_2(g)$ at STP according to the equation: $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$

   What volume of water is produced?

4. How many grams of carbon react if 543 L of gaseous carbon monoxide can be produced at 27°C and 0.247 atm according to the following equation? $2\ C(s) + O_2(g) \rightarrow 2\ CO(g)$

5. Air bags in cars are inflated by the sudden decomposition of sodium azide, NaN_3, by the following reaction: $2\ NaN_3(s) \rightarrow 3\ N_2(g) + 2\ Na(s)$. What volume of nitrogen gas, measured at 1.30 atm and 87°C, would be produced by the reaction of 70.0 g of NaN_3?

Molarity Stoichiometry: Solutions to All Your Problems

***Note:** millimoles (or mmol) can be your BEST friend during solution stoich!

$$\frac{mmol}{mL} = M$$

1. 250 mL of 0.70 M Li_3PO_4 and 250 mL of excess $Ca(OH)_2$ are mixed, producing an aqueous LiOH solution (and a calcium phosphate precipitate). What is the molar concentration of LiOH in this solution?

2. Calculate the mass of Ag_2S produced when 125 mL of 0.200 M $AgNO_3$ is added to excess Na_2S solution.

$$\_\_\ \mathrm{Na_2S}(aq) + \_\_\ \mathrm{AgNO_3}(aq) \rightarrow \_\_\ \mathrm{Ag_2S}(s) + \_\_\ \mathrm{NaNO_3}(aq)$$

3. 50.0 mL of $Mg(OH)_2$ is used to neutralize 35.0 mL of 0.60 M HBr. What is the molarity of the $Mg(OH)_2$?

Molarity Stoich Shortcut: $M_1V_1 = M_2V_2$

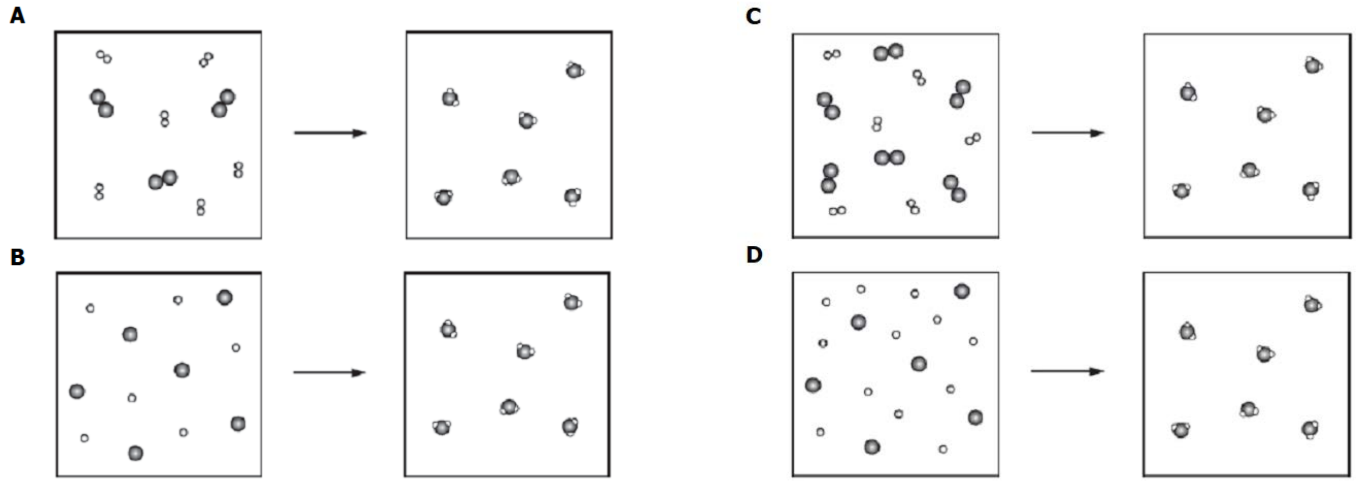
This shortcut is useful iif:

- IF you are starting and ending with molarity, and
- the species compared have a 1:1 mole ratio

4. 125 mL of KOH completely neutralizes 57.0 mL of 1.20 M HBr. Calculate the initial concentration of KOH.

Let's Get Stoiched! Multiple Choice Practice

1. Which of the following particulate diagrams best shows the formation of water vapor from hydrogen gas and oxygen gas in a rigid container at 125°C?



2. A sample of 9.00 grams of aluminum metal is added to an excess of hydrochloric acid, HCl. The volume of hydrogen gas produced at 25°C and 810. mmHg is:

   a. 11.5 L b. 11.2 L c. 7.65 L d. 7.46 L

3. The reaction of 7.8 g benzene, C_6H_6, with excess HNO_3 resulted in 0.90 g of H_2O. What is the percent yield? (Molar masses (g mol^{-1}): C_6H_6 = 78; HNO_3 = 63; $C_6H_5NO_2$ = 123; H_2O = 18.)

$$C_6H_6 + HNO_3 \rightarrow C_6H_5NO_2 + H_2O$$

   a. 90.% b. 50.% c. 12% d. 2.0%

4. A student titrates 20.0 mL of 1.0 M NaOH with 2.0 M formic acid, HCOOH. How much formic acid is needed to completely neutralize the NaOH?

$$\_\_\_\mathrm{HCOOH}(aq) + \_\_\_\mathrm{HNO_3}(aq) \rightarrow \_\_\_\mathrm{NaCOOH}(aq) + \_\_\_\mathrm{H_2O}(l)$$

a. 10.0 mL b. 20.0 mL c. 30.0 mL d. 40.0 mL

Use the following information to answer #5-6: When heated in a closed container in the presence of a catalyst, 1.2 x 10^{23} formula units of potassium chlorate, $KClO_3$, decompose into potassium chloride and oxygen gas via the following reaction:

$$2\ \mathrm{KClO_3}(s) \rightarrow 2\ \mathrm{KCl}(s) + 3\ \mathrm{O_2}(g)$$

5. How many grams of oxygen gas will be generated?

a. 1.60 g b. 4.80 g c. 9.60 g d. 18.37 g

6. Approximately how many liters of oxygen gas will be evolved at STP?

a. 1.24 L b. 3.36 L c. 6.72 L d. 22.4 L

Limiting and Excess Reagents

<u>Limiting reactant</u>: will be completely _________ ____ during the chemical reaction → determines all other amount

<u>Excess reactant</u>: will ______ be completely used up during the chemical reaction → has some left over at the end

<u>Percent Yield</u>: a method to calculate the effectiveness of a chemical reaction.

- **<u>Actual yield</u>**: what you produce from actually doing the reaction in a _______ setting.
- **<u>Theoretical yield:</u>** what "should have been" produced from chemical reaction. This is calculated with _________!

$$\text{Percent Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

How to Find the Limiting Reactant: Two Methods

1. Calculate the number of __________ of each reactant and then divide each one by its _______________________ from the balanced reaction.

   ➔ The reactant which produces the <u>smaller</u> number will be the limiting reactant.

   <u>Advantage</u>: very fast (great for multiple choice!)

   <u>Disadvantage</u>: you still have to actually calculate whatever info is asked for by the problem.

2. Use a BCA table to calculate the ending number of __________ for both reactants and products.

   ➔ The reactant which is <u>entirely used up</u> will be the limiting reactant.

   <u>Advantage</u>: you will calculate the ending amount of your excess reactant AND products, so any calculation you need is already complete.

   <u>Disadvantage</u>: you have to be in MOLES, so you might have extra math to do before and after.

Let's try BOTH methods! Which do you like best?

1. Methanol, CH_3OH, is combusted according to the balanced equation below. If you have 50.0 g of methanol and 50.0 g of molecular oxygen, how many grams of excess reactant remain unreacted?

$$2\,CH_3OH(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

2. Aqueous sodium hydroxide, NaOH, reacts with phosphoric acid to give sodium phosphate and water in the following balanced chemical equation: $3\ NaOH\ (aq) + H_3PO_4\ (aq) \rightarrow Na_3PO_4\ (aq) + 3\ H_2O\ (l)$

   If 20.80 g of NaOH is mixed with 29.40 g of H_3PO_4 (NaOH = 40.00 g/mol, H_3PO_4 = 98.00 g/mol):

   a. How many grams of Na_3PO_4 can be formed?

   b. How many grams of the excess reactant remain unreacted?

   c. If the actual yield of Na_3PO_4 was 15.00 g, what is the percent yield of Na_3PO_4?

3. Hydrofluoric acid solutions cannot be stored in glass containers because HF reacts readily with silica in glass to produce hexafluorosilicic acid (H_2SiF_6).

$$SiO_2(s) + 6HF(aq) \rightarrow H_2SiF_6(aq) + 2H_2O(l)$$

If 40.0 kg silicon dioxide and 40.0 kg of HF react:

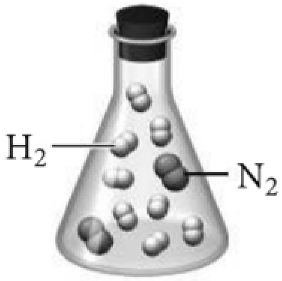
a. Determine the mass of the excess reactant remaining.

b. Determine the theoretical yield of hexafluorosilicic acid, H_2SiF_6.

c. Determine the percent yield if the actual yield is 45.8 kg H_2SiF_6.

Multiple Choice Practice

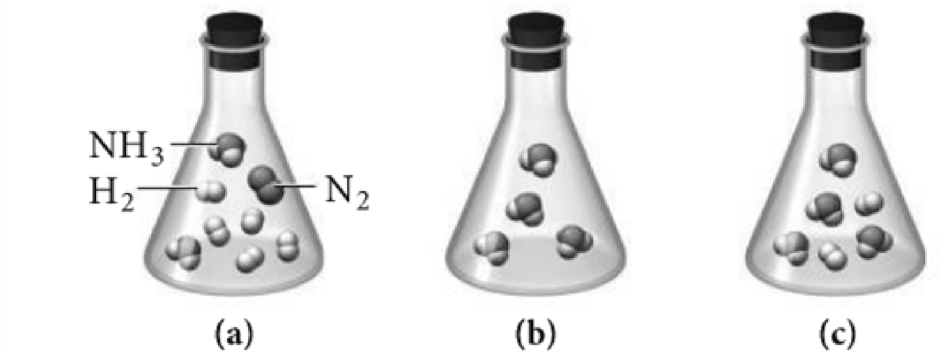
4.



Nitrogen and hydrogen gas react to form ammonia according to the reaction:

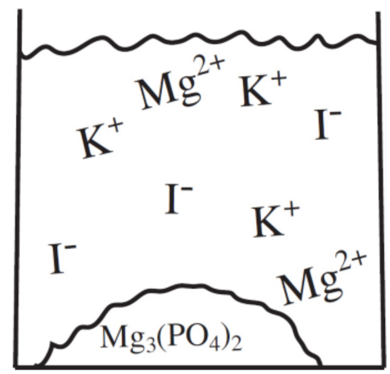
$$N_2(g) + 3H_2(g) \rightarrow 2\,NH_3(g)$$

If a flask contains a mixture of reactants represented by the image at right, which image below best represents the mixture in the flask after the reactants have reacted as completely as possible? What is the limiting reactant? Which reactant is in excess?



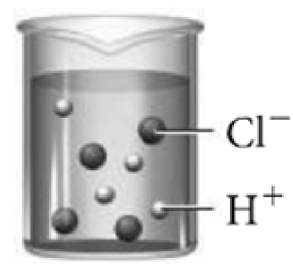
5. Based on the particulate drawing of the products for the reaction below, which reactant is limiting for the following reaction and why?

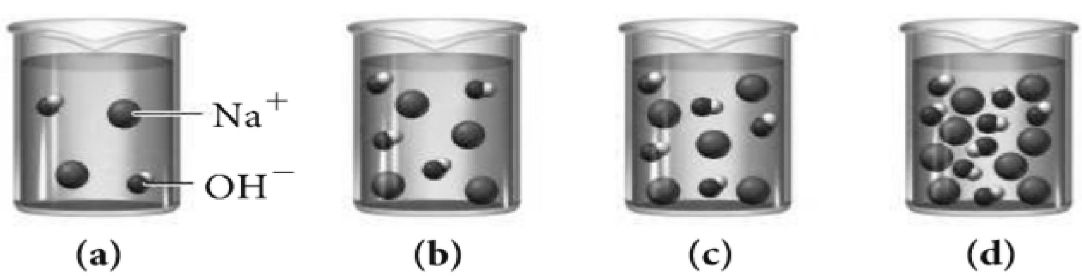
$$2\ K_3PO_4(aq) + 3\ MgI_2(aq) \rightarrow Mg_3(PO_4)_2(s) + 6\ KI(aq)$$



   a. The K_3PO_4, because there are no excess PO_4^{3-} ions after the reaction.
   b. The MgI_2, because there are excess Mg^{2+} cations remaining after the reaction.
   c. The K_3PO_4, because it contains a cation that cannot form a precipitate.
   d. The MgI_2, because it required more of itself to create the products.

6. A hydrochloric acid solution will neutralize a sodium hydroxide solution. Look at the molecular views showing one beaker of HCl and four beakers of NaOH. Which NaOH beaker will just neutralize the HCl beaker? Begin by writing a balanced chemical equation for the neutralization reaction.





7. What is the maximum grams of CO_2 (molar mass 44.0 g/mol) that can be produced from 50.0 mol each of sulfuric acid, H_2SO_4, and sodium hydrogen carbonate, $NaHCO_3$? The balanced equation is:

$$2\ \mathrm{NaHCO_3} + \mathrm{H_2SO_4} \rightarrow 2\ \mathrm{CO_2} + 2\ \mathrm{H_2O} + \mathrm{Na_2SO_4}$$

a. (50.0) (1/2) (44.0)
b. (50.0) (44.0)
c. (50.0) (4) (44.0)
d. (50.0) (2) (44.0)

8. The reaction of silver metal and dilute nitric acid proceeds according to the equation below. If 0.10 mol of powdered silver is added to 10. mL of 6.0 molar nitric acid, HNO_3, the number of moles of NO gas that can be formed is:

$$3\ \mathrm{Ag}(s) + 4\ \mathrm{HNO_3}(aq) \leftrightarrow 4\ \mathrm{AgNO_3}(aq) + \mathrm{NO}(g) + 2\ \mathrm{H_2O}(l)$$

a. 0.015 mol
b. 0.030 mol
c. 0.045 mol
d. 0.020 mol

9. 250. mL of 0.30 M hydrochloric acid is combined with 250. mL of 0.10 M K_2S in the gas evolution shown below.

$$2\ \mathrm{HCl}(aq) + \mathrm{K_2S}(aq) \rightarrow \mathrm{H_2S}(g) + 2\ \mathrm{KCl}(aq)$$

What is the concentration of K^+ ions in the KCl solution formed during this reaction?

a. 0.050 M
b. 0.10 M
c. 0.15 M
d. 0.30 M

10. Which of the following statements about the reaction below is **false**?

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

a. Every methane molecule that reacts produces two water molecules.
b. If 16.0 g of methane react with 64.0 g of molecular oxygen, the combined masses of the products will be 80.0 g.
c. If 11.2 L of methane react with an excess of molecular oxygen, the volume of CO_2 produced at STP is (44/16)(11.2) liters.
d. If 22.4 L of methane at STP react with 64.0 g of molecular oxygen, 22.4 L of CO_2 at STP can be produced.

11. When 12 g of methanol (CH_3OH) was treated with excess oxidizing agent (MnO_4^-), 14 g of formic acid (HCOOH) was obtained. Using the following chemical equation, calculate the percent yield.

$$3\ CH_3OH + 4\ MnO_4^- \rightarrow 3\ HCOOH + 4\ MnO_2$$

a. 100% b. 92% c. 82% d. 70%

12. How many grams of NH_3 can be prepared from 77.3 g of N_2 and 14.2 g of H_2?

a. 47.0 g b. 79.7 g c. 93.9 g d. 120. g

Types of Chemical Reactions

Type of Reaction	What Happens	Common Form
Synthesis Reaction	Two or more substances come together to make _____ new substance.	A + B → C
Decomposition Reaction	One substance breaks down into two or more simpler substances.	C → A + B
Single Replacement Reaction	One element replaces another in a compound.	A + BC → AC + B
Double Replacement Reaction	Ions of two compounds exchange places to form two new compounds.	AB + CD → AD + CB
Combustion reaction	A substance combines with ____________ gas, releasing large amounts of energy in the form of heat and light.	$C_xH_yO_z + O_2 \rightarrow H_2O + CO_2$

There are three subsets of double replacement reactions that you will need to know!

1. Precipitation Reaction: two aqueous solutions mix to form a precipitate (_________) → more about this shortly!

   AB(aq) + CD(aq) → AD(s) + CB(aq)

2. Neutralization (Acid-Base) Reaction: an Arrhenius acid and base react to produce a salt and water.

   $HB + C(OH) \rightarrow H_2O + CB$

3. Gas Evolution Reaction: two aqueous solutions mix to form a gas which _____________ out of solution. There are four general categories of this reaction type!

Compounds that Undergo Gas-Evolution Reactions

Reactant Type	Intermediate Product	Gas evolved (with decomposed products, if applicable)
Sulfides	*None*	$H_2S(g)$
Carbonates and bicarbonates	H_2CO_3	$CO_2(g) + H_2O(l)$
Sulfites and bisulfites	H_2SO_3	$SO_2(g) + H_2O(l)$
Ammonium	NH_4OH	$NH_3(g) + H_2O(l)$

Oxidation/Reduction Reactions: I used to be great at redox, but I got a little rusty. ;D

Oxidation–reduction (redox) reactions: where electrons are transferred from one atom to another.

- If a substance accepts an electron, it is____________________.
- If a substance loses an electron, it is ____________________.
- Electrons are always transferred from the species that is oxidized to the species that is reduced.

Two great mnemonics!

1. ______ _______: **O**xidation **I**s **L**oss (OIL) and **R**eduction **I**s **G**ain (RIG)
2. ______ **goes** ______: A species **loses electrons** when **oxidized**, and **gains electrons** when **reduced**.

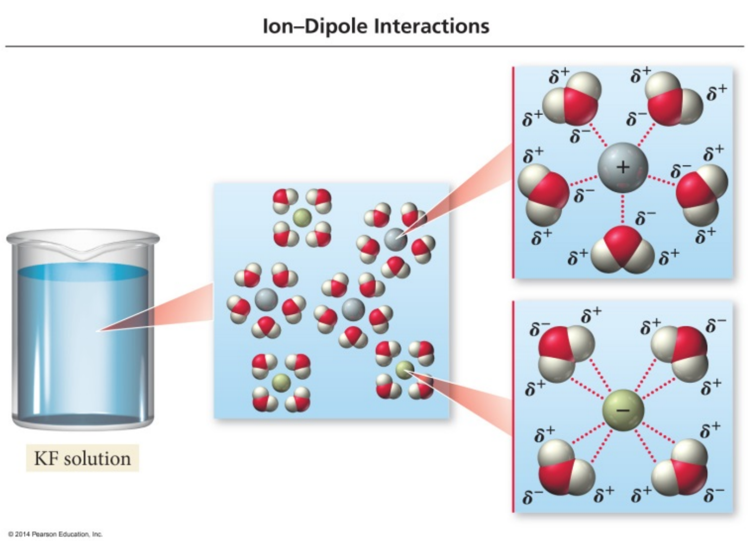
Almost all reaction types (except double replacement) are redox. You will need to understand and interpret oxidation state changes for the test this unit. Next unit, we'll learn so much more!

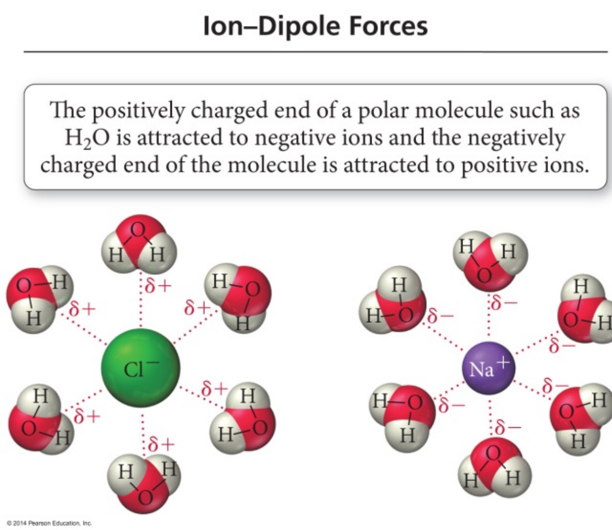
Precipitation Reactions

Precipitation Reactions: a double replacement reaction between aqueous solutions of ionic compounds which produces an ionic compound that is ________________ in water: this insoluble product is called a precipitate.

Quick review: **Ionic Compound Solubility**

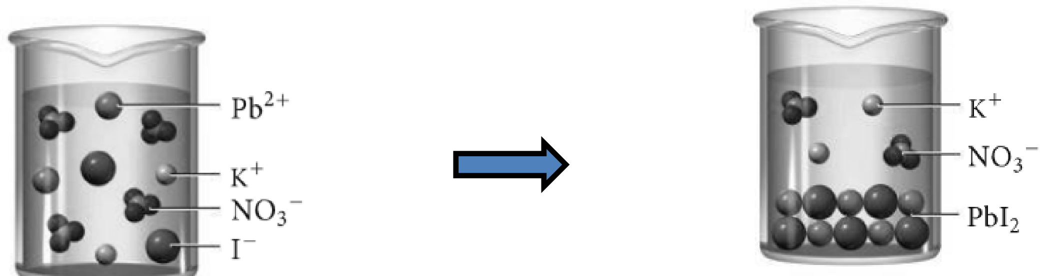
- When we mix a solute with a solvent, there are attractive forces (ion-dipole IMFs) between the solute and solvent particles – if the attraction is strong enough, this is what allows the solute to _________________!
- Ions separate (ionize) from one another when dissolved in water (called _____________________)
- The number of ions produced in solution depends on the ratio in the original __________________.
    - Ex: $Pb(NO_3)_2$ dissociates to form →
        - Thus, **1 formula unit** of $Pb(NO_3)_2$ dissociates to form _____ total ions.





Back to precipitation reactions!

For example, consider what happens if you mix two solutions, KI(aq) and $Pb(NO_3)_2$(aq). Before mixing, both compounds are _______________ in water: they have **dissociated into ions** in their respective solutions. When the two solutions are mixed together:



KI(*aq*) and $Pb(NO_3)_2$(*aq*)

PbI_2(*s*) and KNO_3(*aq*)

In the *instant* that the solutions come into contact, all four ions are present:

In this reaction, KNO_3 is soluble, but PbI_2 is insoluble. Thus, PbI_2 precipitates.

Solubility Rules

There are many rules to determine which ions are soluble in water. Although you no longer need to know the very large list of solubility rules from the AP test of yore, there are some basic solubility rules you MUST know!

<u>Must be memorized</u>:

Alkali metal cations, NH_4^+, and NO_3^- are always soluble

Other solubility rules will be given within a problem or can be figured out by context.

<u>Note</u>: Precipitation reactions do not always occur when two aqueous solutions are mixed!

	Precipitation Rxn?	Why?
RbCl(*aq*) + NaBr(*aq*) → RbBr(aq) + NaCl(aq)	Nope.	No insoluble substance produced.
HCl(*aq*) + $AgNO_3$(*aq*) → AgCl(*s*) + HNO_3(*aq*)	Yes!	Solid produced from aqueous reactants.
3 MnO_2(aq) + 4 Al(s) → 3 Mn(s) + 2 Al_2O_3(aq)	Nope.	Not starting from only aqueous reactants.

<u>Ways to Represent a Precipitation Reaction</u>

<u>Complete ionic equation</u>: Starting from the full equation, split any aqueous substances into their ions, but leave non-aqueous substances intact.

<u>Net ionic equation</u>: Starting from the complete ionic equation, remove all ions that found in exactly the same form on both sides of the equation (called <u>spectator ions</u>, since they do not participate in reaction).

Example

Molecular (Full) Equation: $2AgNO_3(aq) + CaCl_2(aq) \rightarrow 2AgCl(s) + Ca(NO_3)_2(aq)$

Complete Ionic Equation: $2Ag^+(aq) + 2NO_3^-(aq) + Ca^{2+}(aq) + 2Cl^-(aq) \rightarrow 2AgCl(s) + Ca^{2+}(aq) + 2NO_3^-(aq)$

<u>**Net ionic**</u> equation:

$2Ag^+(aq) + 2Cl^-(aq) \rightarrow 2AgCl(s)$ simplifies to:

Net ionic equation

Let's Try!

1. Why is a precipitation reaction best represented by a net ionic equation, rather than the complete or molecular equation?

2. Complete and balance the following reaction, then write the complete AND net ionic equations.

___$H_2SO_4(aq)$ + ___$NaOH(aq) \rightarrow$

3. Write the net ionic equation for the following reactions.

a. $(NH_4)_2SO_4(aq) + 2\ AgNO_3(aq) \rightarrow 2\ NH_4NO_3(aq) + Ag_2SO_4(s)$

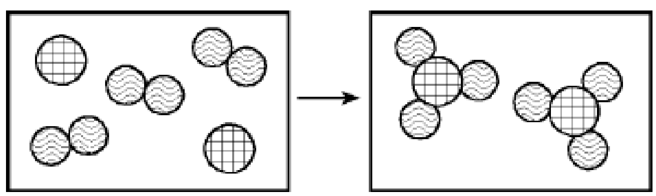
b. $H_2SO_4(aq) + 2\ LiOH(aq) \rightarrow 2\ H_2O(l) + Li_2SO_4(aq)$

More Reactions Practice!

Part I: Complete the following chart. (Balance the reactions if you want more practice!)

Reaction	Type of Reaction S, D, SR, DR, or C?	Redox: Yes or no?
1. ___Ca + ___H_2O → ___$Ca(OH)_2$ + ___H_2		
2. ___ $NH_4F(aq)$ + ___ $Sr(ClO_3)_2$ → ___ NH_4ClO_3 + ___SrF_2		
3. ___HNO_3 + ___$LiOH$ → ___ $LiNO_3$ + ___ H_2O		
4. ___C_6H_{14} + ___O_2 → ___H_2O + ___CO_2		
5. ___LiI + ___$Pb(NO_3)_2$ → ___$LiNO_3$ + ___PbI_2		

Part II: Multiple Choice Practice



1. The diagram above best represents which type of reaction?
   a. Acid/base
   b. Oxidation/reduction
   c. Precipitation
   d. Decomposition

2. If we dissolve 25 grams of salt in 251 grams of water, what is the mass of the resulting solution?
   a. 251 g
   b. 276 g
   c. 226 g

3. Consider the following three equations for chemical reactions:

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

   These are examples of:
   a. three redox reactions
   b. three acid-base reactions
   c. a redox reaction, a precipitation reaction, and an acid-base reaction
   d. a neutralization reaction, then two precipitation reactions

4. Given the net ionic equation

$$3\ \mathrm{Ba^{2+}}(aq) + 2\ \mathrm{PO_4^{3-}}(aq) \rightarrow \mathrm{Ba_3(PO_4)_2}(s)$$

how many grams of iron (II) phosphate must be present to react with 2.0×10^2 grams of barium bromide?

(MW $Fe_3(PO_4)_2$ = 357 g/mol, MW $BaBr_2$ = 297 g/mol)

A

$$200g\ BaBr_2 \times \frac{1\ mol\ BaBr_2}{297g\ BaBr_2} \times \frac{1\ mol\ Fe_3(PO_4)_2}{3\ mol\ BaBr_2} \times \frac{357g\ Fe_3(PO_4)_2}{1\ mol\ Fe_3(PO_4)_2}$$

B

$$200g\ BaBr_2 \times \frac{1\ mol\ BaBr_2}{297g\ BaBr_2} \times \frac{2\ mol\ Fe_3(PO_4)_2}{3\ mol\ BaBr_2} \times \frac{357g\ Fe_3(PO_4)_2}{1\ mol\ Fe_3(PO_4)_2}$$

C

$$200g\ BaBr_2 \times \frac{1\ mol\ BaBr_2}{297g\ BaBr_2} \times \frac{3\ mol\ Fe_3(PO_4)_2}{1\ mol\ BaBr_2} \times \frac{357g\ Fe_3(PO_4)_2}{1\ mol\ Fe_3(PO_4)_2}$$

D

$$200g\ BaBr_2 \times \frac{1\ mol\ BaBr_2}{297g\ BaBr_2} \times \frac{3\ mol\ Fe_3(PO_4)_2}{2\ mol\ BaBr_2} \times \frac{357g\ Fe_3(PO_4)_2}{1\ mol\ Fe_3(PO_4)_2}$$

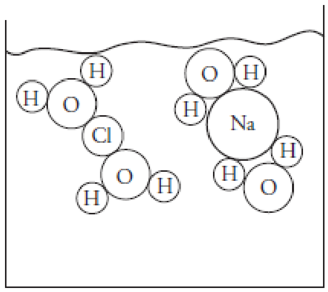
5. If solutions containing equimolar amounts of $AgNO_3$ and KCl are mixed, what is the identity of the spectator ions?

   a. Ag^+, NO_3^-, K^+, and Cl^-
   b. Ag^+ and Cl^-
   c. Ag^+ and K^+
   d. NO_3^- and K^+

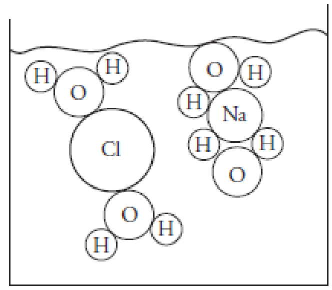
6. Choose the correct net ionic equation representing the precipitation reaction that occurs when solutions of potassium carbonate, K_2CO_3, and copper (I) chloride, CuCl, are mixed.

   a. $\mathrm{K_2CO_3}(aq) + 2\ \mathrm{CuCl}(aq) \rightarrow 2\ \mathrm{KCl}(aq) + \mathrm{Cu_2CO_3}(s)$
   b. $\mathrm{K_2CO_3}(aq) + 2\ \mathrm{CuCl}(aq) \rightarrow 2\ \mathrm{KCl}(s) + \mathrm{Cu_2CO_3}(aq)$
   c. $\mathrm{CO_3^{2-}}(aq) + 2\ \mathrm{Cu^+}(aq) \rightarrow \mathrm{Cu_2CO_3}(s)$
   d. $\mathrm{K^+}(aq) + \mathrm{Cl^-}(aq) \rightarrow \mathrm{KCl}(s)$

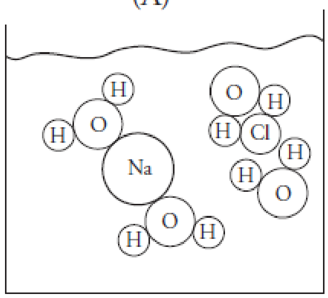
7. Which of the following diagrams best represents what is happening on a molecular level when NaCl dissolves in water?



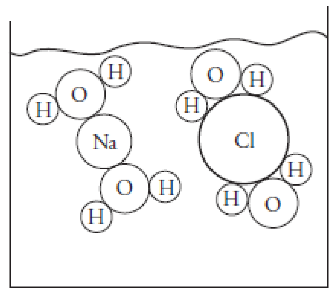
(A)



(C)



(B)



(D)

8. If 0.62 g of sodium sulfite is added to excess nitric acid, how many mL of SO_2 gas will be evolved at 625 mmHg and 27°C? (The molar mass of sodium sulfite is 126 g/mol).

$$\mathrm{Na_2SO_3}(aq) + 2\ \mathrm{HNO_3}(aq) \rightarrow \mathrm{SO_2}(g) + \mathrm{H_2O}(l) + 2\ \mathrm{NaNO_3}(aq)$$

a. 0.018 mL   b. 0.20 mL   c. 13 mL   d. 150 mL

Gravimetric Analysis Lab

Introduction

Chemists can find the identity of unknown compounds using techniques such as qualitative analysis, chromatography, spectroscopy, and gravimetric analysis. Gravimetric analysis, which uses a balance to determine the mass of a substance, is one of the oldest and most accurate quantitative methods for determining the amount of an analyte in a sample.

The Crime

Mr. Groff takes daily supplements to ensure that he gets enough essential nutrients and to maintain and improve his health. One of the supplements is a capsule (pill) that contains potassium carbonate, K_2CO_3, **and** a sugar filler. Last week, just after Mr. Groff took his supplements, he heard an evil laugh and was then knocked unconscious. When he awoke with a splitting headache, he found the following note:

Think you're [11 **Na** 22.99] *funny?*

I've tampered with your supplements and you've just taken an unknown amount of carbonate ion! Who's laughing now?

I can't wait until you [18 **Ar** 39.95]*...until you are just bones so I can* [56 **Ba** 137.33].

Groff realized the culprit was most likely one of his chemistry teachers and quickly contacted the police. The police searched the homes of all of the RRHS chemistry teachers and discovered that the following teachers had potassium carbonate pills in their possession, each with a different percent by mass.

Suspect	**% Carbonate by Mass** (in K_2CO_3 pill)
Ms. Madeline Childs	25%
Mr. Trevor Dado	50%
Ms. Victoria Rodgers	75%
Mr. Patrick Boylan	100%

Your task in this lab is to help Mr. Groff identify the percent by mass of the carbonate in the tampered supplements from his home using gravimetric analysis so that the police can arrest the culprit.

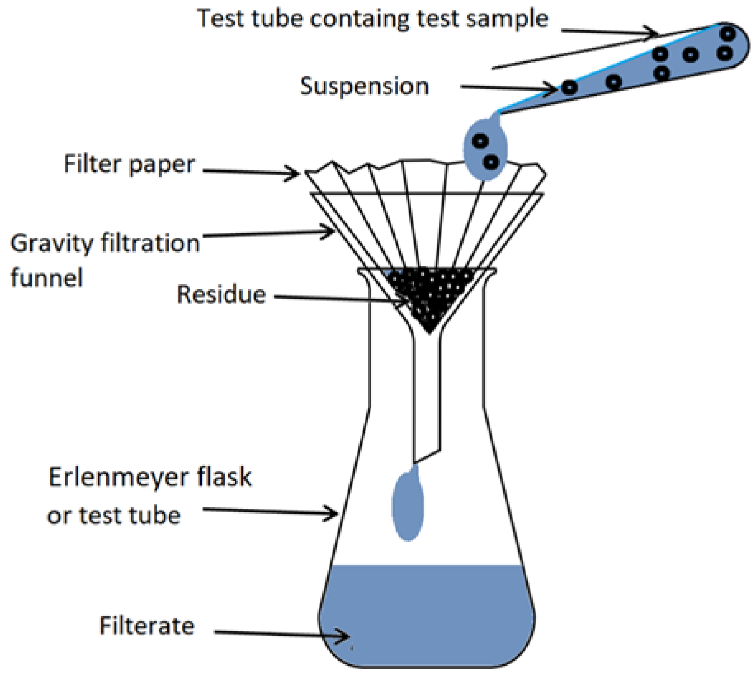
Chemicals

Unknown A (alkali metal carbonate)
0.50 M calcium chloride, $CaCl_2$
Distilled water and wash bottle

Equipment

100 mL beaker
Erlenmeyer flask + Buchner funnel
Stirring rod
Filter paper
Watch glass

Procedure: Day 1



1. Place a clean 100-mL beaker on an analytical balance and record the exact mass.
2. Carefully open the capsule (pill) and add contents into the 100-mL beaker. Record the exact mass in the data table.
3. Use your wash bottle to add about 20 mL of distilled water into beaker and stir with the stirring rod to dissolve capsule contents.
4. Add 30 mL of 0.50 M $CaCl_2$ to the solution containing the capsule components and stir to dissolve.
5. **Write your initials on a piece of filter paper with a pencil.** Record the mass of the dry filter paper.
6. Assemble a gravitation filtration set-up (shown to the right) using the Erlenmeyer flask and filter paper. Place the filter paper into the funnel and wet it with a small amount of distilled water.
7. To filter your sample, slowly pour into the center of the filter paper.
8. Rinse the beaker and stirring rod with deionized (DI) water from your wash bottle to ensure that all of the precipitate is collected on the filter paper. (Reminder: do NOT contaminate the wash bottle.)
9. Continue filtering until all of the filtrate (liquid) has moved through the filter paper into the flask.
10. "Wash" your precipitate by rinsing it with deionized water from your wash bottle.
11. While your precipitate finishes dripping, measure the mass of your watch glass using an analytical balance. Record the exact mass in the data table.
12. When the filter paper is sufficiently dry, use forceps or the tip of a scoopula to carefully remove the wet filter paper with the precipitate from the funnel.
    a. Carefully unfold your filter paper and place it onto a watch glass.
    b. Place your watch glass and filter paper into a drying oven overnight. **This will remove any remaining water to ensure accurate mass measurements of just the precipitate**.
13. Clean-up your lab station.
    a. Rinse the filtrate solution down the drain.
    b. Clean the 100-mL beaker: you will need to scrub the beaker so that all white residue is removed.

Procedure: Next Class Day

14. Now that the filter paper has been thoroughly dried,
    a. Remove the watch glass and filter paper from the drying oven.
    b. Measure its mass using an analytical balance. Record the exact mass in the data table.
15. Clean up:
    a. Dry filter paper with precipitate goes in the trash.
    b. Wash the watch glass: you will need to scrub the watch glass so that all white residue is removed.

Data and Calculations Table

	Mass (g)
Mass of empty beaker	
Mass of beaker and capsule contents	
Calculation: Mass of capsule contents	
Mass of dry filter paper	
Mass of watch glass	
Mass of watch glass, dry filter paper, and precipitate	
Calculation: Mass of precipitate	

Calculations and Analysis (Show all work!)

1. Write the balanced, molecular (complete) equation for the precipitation reaction that occurred, including states.

2. What is the identity of your precipitate? How do you know?

3. Write the balanced net ionic equation for the precipitation reaction that occurred, including states of matter.

4. Calculate the moles of precipitate collected.

5. Calculate the mass percent of carbonate ion, CO_3^{2-}, in the capsule.

6. Based on your results, who was the culprit in this attempt on Mr. Groff's life? How do you know?

7. In gravimetric analysis, we often refer to "drying to constant mass."
    a. What is the point of drying the filtrate to constant mass? (Hint: what chemical are we concerned about?)
    b. Experimentally, what procedure can be followed to be certain the filtrate was dried to constant mass?
    c. If we failed to dry to constant mass, would the experimentally determined mass percent of carbonate, CO_3^{2-}, be greater than, less than, or equal to the actual value? Justify your answer.

8. Identify a potential lab error that would cause your experimentally determined mass percent of carbonate, CO_3^{2-}, to be <u>too low</u>. Explain how this lab error would affect your data and how this change would result in an experimentally determined carbonate mass percent that is smaller than the actual value.

<u>Stoichiometry in the Lab</u>: Titrations and Gravimetric Analysis!

Titration

- <u>Goal</u>: to determine the unknown concentration!
- A substance in a solution of unknown concentration is reacted with another substance in a solution of known concentration
- <u>Vocab</u>
  - **Titrant** – solution of ___________________ concentration (usually in a buret)
  - **Analyte** – solution of _____________________ concentration (usually in flask or beaker)
  - **Equivalence point** – acid and base have completely ____________________________ each other
    - Moles _________ = Moles _________
  - **End point** - ________________________ changes color

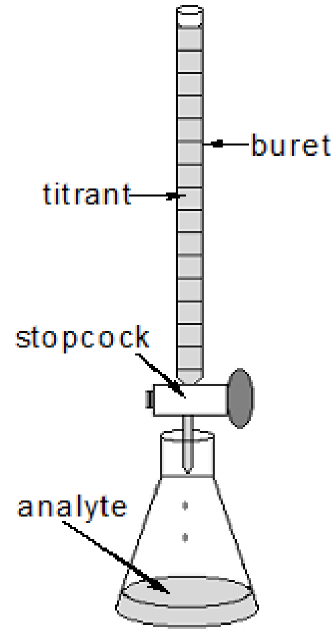


Figure 1: Titration Setup

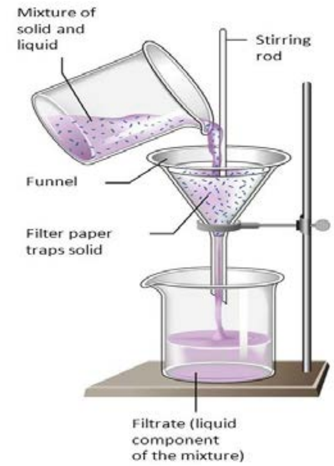
<u>Example:</u> By titration, 0.13 M aqueous sulfuric acid, H_2SO_4, neutralized 27.4 mL of 0.17 M LiOH solution. What was the volume of the acid solution used in the titration?

<u>Gravimetric Analysis</u>: a laboratory procedure in which an ion is precipitated out of a mixture in order to find the percent mass of that ion in an impure substance.

Method:

1. Reactant impure material with a known compound, to form a ________________ containing the ion of interest.
2. Filter and dry precipitate, then measure its mass.
3. Use stoichiometry calculations to determine the _________ of the ion of interest, using the balanced reaction to work backwards from the mass of the precipitate measured in the lab.

Steps in Gravimetric Analysis (Lab Technique)



1. Dry and weigh impure sample.
2. _________________ sample.
3. Add precipitating reagent in _______________.
4. Filter precipitate from solution.
5. __________ precipitate to remove impurities.
6. Dry and weigh to constant ________ (±0.03 g)

Steps in Gravimetric Analysis (Math-y Math)

1. Write the balanced precipitation reaction.
2. Use stoichiometry to convert from the mass of your precipitate to the mass of the ion of interest.
3. Calculate % by mass: $\%\, mass\ of\ ion = \frac{mass\ of\ ion\ (calculated\ in\ step\ 2)}{mass\ of\ original\ sample} \times 100$

Example: The following data were obtained when a sample of unknown alkali metal chloride was analyzed:

Mass of watch glass	8.2030 g
Mass of unknown chloride sample + watch glass	8.3014 g
Mass of AgCl(s) + watch glass (after drying)	8.3922 g

a. Calculate the mass of the precipitate.

b. Calculate the mass of chloride ions, Cl^-, in the original sample.

c. Calculate the percent mass of chloride ions, Cl^-, in the sample.

d. Which of the following compounds is most likely to be the identity of your unknown alkali metal chloride: $LiCl$, $NaCl$, KCl, or $RbCl$? Justify with math.

Common Lab Errors in Gravimetric Analysis

Error	Effect on Calculated % by Mass of Ion
Failure to add precipitating reagent in excess.	• Not all of the ion of interest will precipitate, thus some of the ion will be lost in the filtrate. • The calculated ion % will be _____________than the actual ion % by mass.
Failure to wash the precipitate to remove soluble impurities.	• Some impurities will be left in the precipitate, so measured mass includes both the precipitate and these contaminants. • The calculated ion % will be _____________than the actual ion % by mass.
The precipitate is excessively washed or washed with warm water.	• Although the precipitate is "insoluble", small amounts of precipitate will dissolve into the wash water and be carried into the filtrate. • The calculated ion % will be _____________than the actual ion % by mass.
The precipitate is not dried to "constant mass".	• Not all of the water molecules will have been removed, so measured mass includes both the precipitate and remaining water molecules. • The calculated ion % will be _____________than the actual ion % by mass.

Let's practice!

1. A 3.187 g impure sample containing sulfur and soluble inert material needed to be analyzed for sulfur content. As part of the procedure, the entire impure sample is dissolved in 350. mL of water, and the sulfur is converted to sulfate ion, SO_4^{2-}. Barium nitrate, $Ba(NO_3)_2$, is added, which causes the sulfate to precipitate out as $BaSO_4$. After being rinsed with distilled water and collected, the dried $BaSO_4$ has a mass of 2.005 g.

   a. What mass of sulfur was in the original sample?

   b. What is the percent of sulfur in the original ore?

   c. If a small of amount of the $BaSO_4$ precipitate dissolved while being rinsed, what effect would this have on the calculated percent by mass of sulfur? Explain.

   d. If the original impure sample was instead dissolved into 500. mL of water, what effect would this have on the calculated percent by mass of sulfur? How do you know?

2. During a titration, 920. mL of a $Sr(OH)_2$ solution is needed to neutralize 0.35 kg of $HClO_4$. What is the molarity of the initial strontium hydroxide solution?

3. Suppose that 0.323 g of an unknown salt containing sulfate, SO_4^{2-}, is dissolved in 50 mL of water. The solution is acidified with 6 M HCl, heated, and an excess of aqueous $BaCl_2$ is slowly added to the mixture resulting in the formation of a white precipitate, $BaSO_4$.

    a. Assuming that 0.433 g of precipitate is recovered, calculate the percent by mass of SO_4^{2-} in the unknown salt.

    b. If it is assumed that the salt is an alkali sulfate determine the identity of the alkali cation.

4. A 0.4500 g sample of impure potassium chloride, KCl, was dissolved in water and treated with an excess of silver nitrate, $AgNO_3$. A 0.8402 g precipitate of AgCl was massed after reacting, collecting, washing and drying the precipitate. Calculate the percentage KCl in the original sample.

Multiple Choice Practice:

1. A sample of an unknown chloride compound was dissolved in water, and then titrated with excess $Pb(NO_3)_2$ to create a precipitate. After drying, it is determined there are 0.0050 mol of precipitate present. What mass of chloride is present in the original sample?

   a. 0.177 g b. 0.355 g c. 0.522 g d. 0.710 g

2. 0.60 M HNO_3 was used to neutralize 15 mL of 0.30 M $Sr(OH)_2$. What volume of HNO_3 was needed?

   a. 7.5 mL b. 15.0 mL c. 22.5 mL d. 30.0 mL

3. Which of the chemical reactions described below is/are an example of a redox reaction(s)?

   I. $CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + NaCl(aq)$

   II. $CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$

   III. $2Al(OH)_3(s) \rightarrow Al_2O_3(s) + 3H_2O(g)$

   a. I only b. II only c. II and III only d. I and III only

4. A student mixed equimolar amounts of $Cu(NO_3)_3$(aq) and LiOH(aq), and a blue precipitate formed. Which reaction below shows the correct net ionic equation for this reaction?

   a. $Cu(NO_3)_3(aq) + 3\ LiOH(aq) \rightarrow Cu(OH)_3(s) + 3\ LiNO_3(aq)$

   b. $Cu(NO_3)_3(aq) + 3\ LiOH(aq) \rightarrow Cu(OH)_3(aq) + 3\ LiNO_3(s)$

   c. $Cu^+(aq) + OH^-(aq) \rightarrow CuOH(s)$

   d. $Cu^{3+}(aq) + 3\ OH^-(aq) \rightarrow Cu(OH)_3(s)$

   e. $Li^+(aq) + OH^-(aq) \rightarrow LiNO_3(s)$

   f. $3\ Li^+(aq) + 3\ OH^-(aq) \rightarrow 3\ LiNO_3(s)$

Intro Equilibrium

Equilibrium Fundamental Ideas:

1. Chemical reactions don't always run from reactants to products: many are reversible (can go back and forth between products and reactants)

Language:

- Rate products are being formed (how quickly products are made) = forward rate of reaction
- Rate reactants are being formed (how quickly reactants are made) = reverse rate of reaction
- A reaction that goes to completion turns all reactants into products before it stops

2. Reversible reactions NEVER STOP, but…
   - Eventually the **rate products are formed = rate reactants are formed** (forward rate = reverse rate)
   - This point is called **equilibrium**
   - At equilibrium, it LOOKS (macroscopically) like the reaction has stopped (concentrations, pressures, etc stay constant) but that's because we can't see down to the atomic level.
3. We look at a ratio of products to reactants to see where equilibrium lies, and how close current conditions are to reaching equilibrium.
4. Discussion question: why would chemists in the real world care about the ratio of products/reactants when a reaction stops (macroscopically) changing?

$$aA + bB \leftrightharpoons cC + dD$$

$$\text{Ratio} = \frac{[\textbf{Products}]}{[\textbf{Reactants}]} = K \text{ (at equilibrium)}$$

$$\text{Ratio} = \frac{[\textbf{Products}]}{[\textbf{Reactants}]} = Q \text{ (not sure if equilibrium)}$$

Equilibrium: Let's Get Balanced!

Dynamic Equilibrium

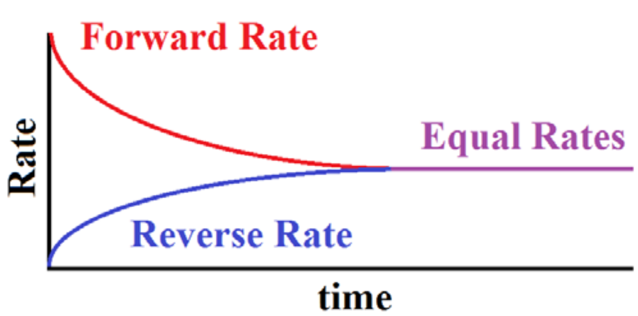
1. Many chemical reactions are at ______________________, which is indicated by double arrows (⇋).
2. At equilibrium, both ____________________ and ____________________ are present.

Forward and backward reactions **rates** are ______________________.

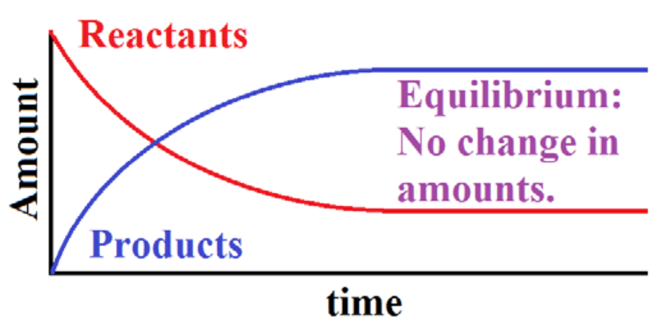
This means that the forward and backward reactions _______________ stop!

- But… concentrations of reactants and products are ______ (usually) equal.
- Instead, the concentrations of the reactants and products are _______________________.

Rate of Reaction vs Time



[Reactants], [Products] vs Time



The Equilibrium Constant, K (the Law of Mass Action): relates the concentrations of ______________________ and ____________________ **at equilibrium** at a given temperature.

For the general reaction: aA + bB ⇋ cC + dD

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b} = \frac{[Products]^{coefficient}}{[Reactants]^{coefficient}}$$

a) Each concentration is raised to the power of its stoichiometric coefficient in the _________________ equation.

b) **Note**: Pure solids and pure liquids (e.g. water) are _________ placed into the expression.

c) There are NO ______________ for equilibrium constant K (they cancel out).

Let's Practice! Write the K_c expressions for the following reactions:

1. $C_6H_{12}(g) + 9\ O_2(g) \rightleftharpoons 6\ CO_2(g) + 6\ H_2O(l)$

2. $Ni(s) + 4\ CO(g) \rightleftharpoons Ni(CO)_4(g)$

3. For $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$, at equilibrium $[N_2] = 1.50$ M, $[H_2] = 2.00$ M, and $[NH_3] = 0.010$ M.

   a. Write the equilibrium constant expression.

   b. Solve for K.

   c. Is the forward or reverse reaction favored, and how do you know?

Important Information about K, the equilibrium constant

→ Regardless of initial conditions, <u>at a given temperature</u> a reaction will reach equilibrium with the ____________ ratio of products to reactants

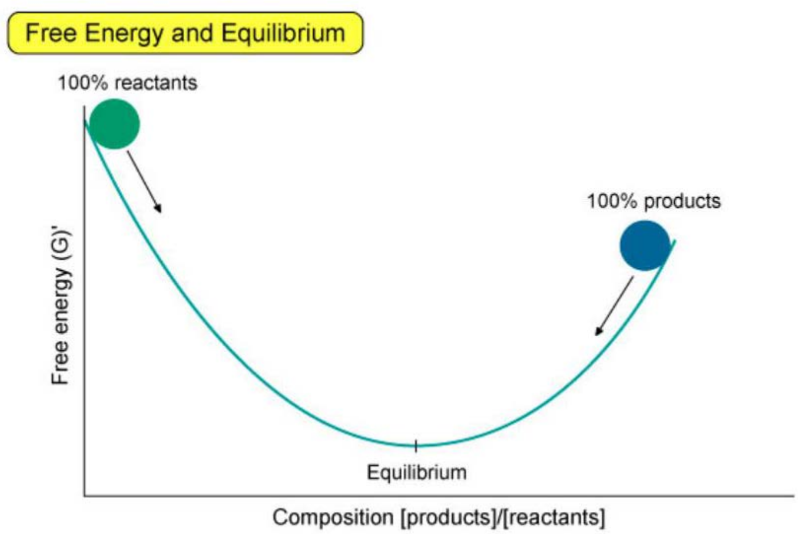
→ **<u>Equilibrium is temperature dependent</u>!!!**

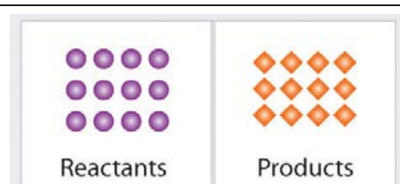
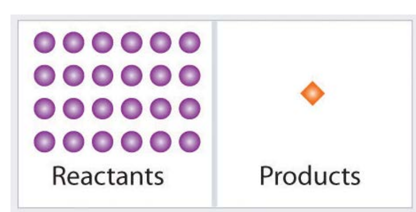
Change the _______________________, change the ________ of products to reactants (K).

What does *K* mean?

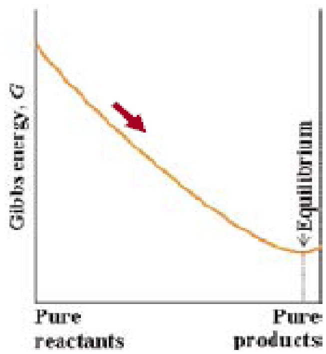
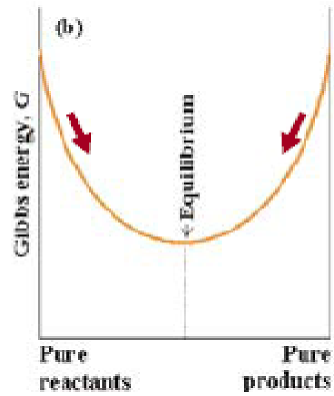
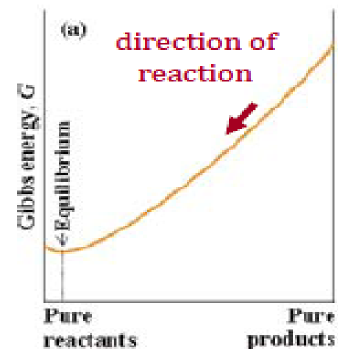
The equilibrium constant, K, tells you:

- the ratio of products to reactants when a given reaction reaches its _____________ free energy state and "stops".

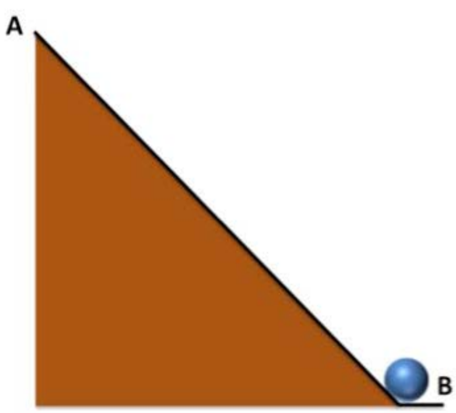


Large *K* (*K* >> 1)	**Intermediate *K*** (*K* ≈ 1)	**Small *K*** (*K* << 1)
Product-favored	Neither	Reactant-Favored
 Mostly products at equilibrium	 Significant amounts of reactants and products at equilibrium	 Mostly reactants at equilibrium

What does *K* mean?

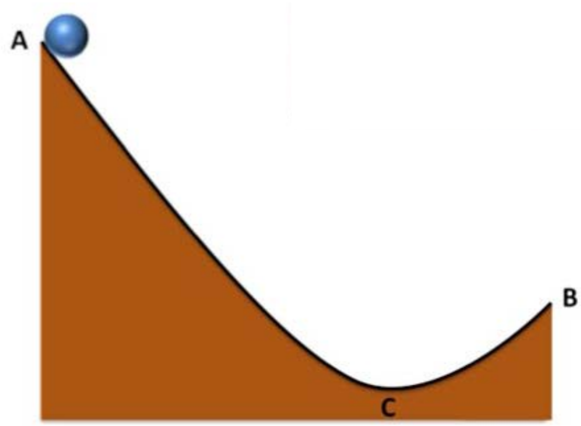
Large *K* (*K* >> 1)	**Intermediate *K*** (*K* ≈ 1)	**Small *K*** (*K* << 1)
Product-favored	Neither	Reactant-Favored
Forward reaction is thermodynamically favorable  	 	Reverse reaction is thermodynamically favorable 

Completion Reactions vs Equilibrium Reactions



Complete Reaction: Complete consumption of reactants.

$$A \rightarrow B$$



Equilibrium: Reaction achieves equilibrium by reaching the lowest state of energy.

$$A \rightleftharpoons B$$

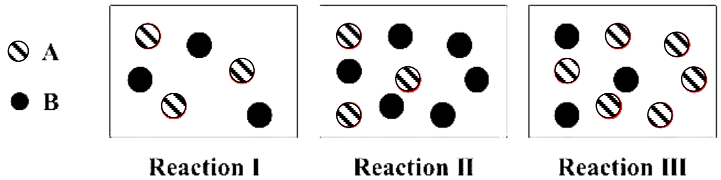
Completion reactions:	Equilibrium reactions:
• Non-reversible • _________ of reactants convert to products • Example: combustion	• Reversible • Reaction will occur until lowest energy state is reached • Example: weak acid dissolution

You can think of K as measuring how close a reaction will go towards 100% completion:

- High K = reaction came pretty close to completely turning reactants into products
- Low K = reaction did NOT come close to turning all reactants into products, mostly reactants just hung around

Let's Practice!

1. The following boxes represent reactions of A ⇋ B at equilibrium.

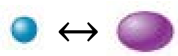


For which reaction shown above is K smallest?

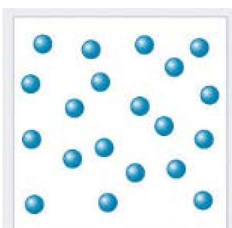
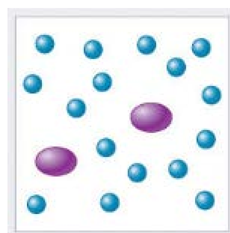
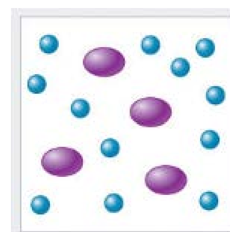
   a. Reaction I
   b. Reaction II
   c. Reaction III
   d. Cannot be determined.

Consider the following reversible reaction to answer #2–3.

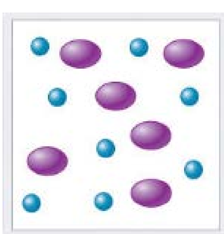
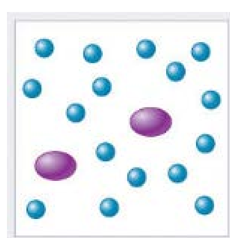
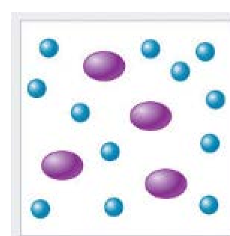
$$A \leftrightarrow B$$



2. Which of the following equilibrium systems has the largest value of K?

a.  b.  c. 

3. Which of the following equilibrium systems has the smallest value of K?

b.  b.  c. 

The Reaction Quotient, Q: The ratio of products to reactants at <u>current conditions</u> (querrent quonditions ;)

Calculate Q if: you need to know the answer to the question, "Is the system at equilibrium?"

A: The answer can be _________ or ________!

For the general reaction: aA + bB ⇋ cC + dD

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

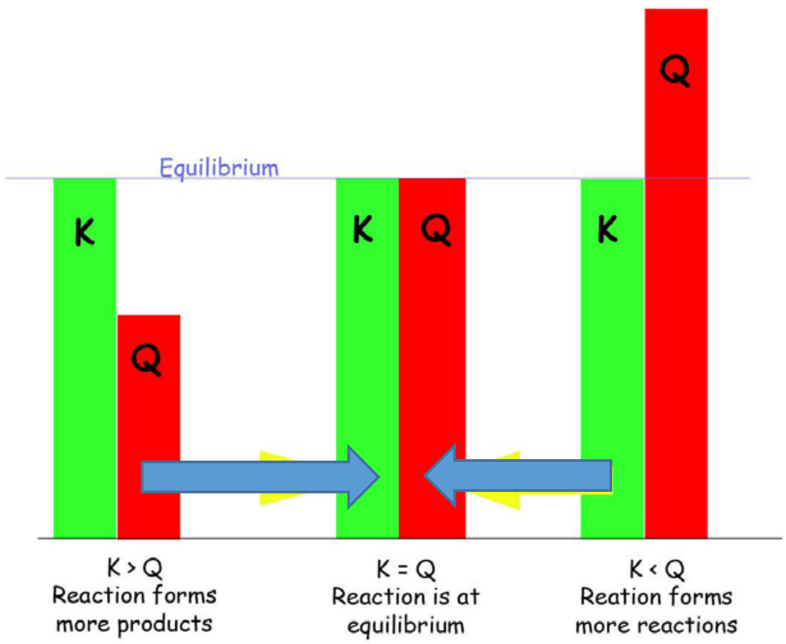
- Q has the appearance of K but the concentrations are not *necessarily* at equilibrium.
- K is _____________ (at constant temperature), but Q can _________________!

Compare K and Q to see if you're at equilibrium AND how to get there!

1. If **K** _____ **Q**, the system is not at equilibrium: forward reaction is favored (shift right, or _____) to make Q = K.
2. If **K** _____ **Q**, the system is at equilibrium.
3. If **K** _____ **Q**, the system is not at equilibrium: reverse reaction is favored (shift left, or _____) to make Q = K.

In Summary

Current conditions	$K > Q$	$K \approx Q$	$K < Q$
change needed for system to reach equilibrium	shift right (make more products)	already at equilibrium	shift left (make more reactants
reaction rates	forward > reverse reaction rate (until equilibrium reached)	forward ≈ reverse reaction rate	forward < reverse reaction rate (until equilibrium reached)



Notes about Language: Talking about equilibrium can be tricky! Here's a quick guide to the terminology used.

Phrases used to describe directions of reaction shift	
If the reaction needs more <u>products</u> to reach equilibrium:	**If the reaction needs more <u>reactants</u> to reach equilibrium:**
1. The reaction will shift right.	1. The reaction will shift left.
2. The forward reaction is occurring more rapidly than the reverse reaction.	2. The reverse reaction is occurring more rapidly than the forward reaction.
3. The reaction will shift to form more products.	3. The reaction will shift to form more reactants.
4. The reaction will proceed to the right.	4. The reaction will proceed to the left.

Let's Practice!

1. For the synthesis of ammonia at 500°C, $N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g)$, the equilibrium constant is 6.0×10^{-2}. In which direction will the system shift to reach equilibrium (at 500°C) if $[NH_3]_{initial} = 1.0 \times 10^{-3}$ M, $[N_2]_{initial} = 1.0 \times 10^{-5}$ M, and $[H_2]_{initial} = 2.0 \times 10^{-3}$ M?

2. For the reaction $2\ NO(g) \rightleftharpoons N_2(g) + O_2(g)$, the equilibrium constant K = 2.4×10^3 at a certain temperature. The initial concentrations are 0.044 *M* NO, 2.0 *M* N_2, and 0.65 *M* O_2. Is the system at equilibrium? If not, which way will the reaction shift and why?

4. The value of the equilibrium constant, K_c, at 25°C is 8.1 for the following reaction:

$$2\ SO_3(g) \rightleftharpoons 2\ SO_2(g) + O_2(g)$$

What must happen for the reaction to reach equilibrium if the initial concentrations of all three species was 2.0 M?

   a. The rate of the forward reaction would increase, and $[SO_3]$ would decrease.
   b. The rate of the reverse reaction would increase, and $[SO_3]$ would decrease.
   c. Both the rate of the forward and reverse reactions would increase, and the value for the equilibrium constant would also increase.
   d. No change would occur in either the rate of reaction or the concentrations of any of the species.

5. Under equilibrium conditions, 0.60 moles of A, 0.60 moles of B and 0.60 moles of C are present in a 3.1 L solution for the reaction shown below. Determine the value of the equilibrium constant, *K*.

$$2\ \mathrm{A} + \mathrm{B} \leftrightarrow 2\ \mathrm{C}$$

a. $K = 17$ b. $K = 5.2$ c. $K = 1.7$ d. $K = 0.52$

Use the following information to answer #6-8: A sample of H_2S gas is placed in an evacuated, sealed container and heated until the following decomposition reaction occurs at 1000 K:

$$2\ \mathrm{H_2S}(g) \leftrightarrow 2\ H_2(g) + \mathrm{S_2}(g) \quad K_c = 86.2$$

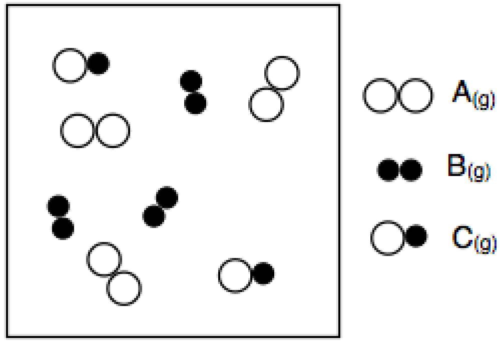
6. Which of the following represents the equilibrium constant for this reaction?

a. $K_c = \frac{[\mathrm{H_2}]^2[\mathrm{S_2}]}{[\mathrm{H_2S}]^2}$ c. $K_c = \frac{2[\mathrm{H_2}][\mathrm{S_2}]}{2[\mathrm{H_2S}]}$

b. $K_c = \frac{[\mathrm{H_2S}]^2}{[\mathrm{H_2}]^2[\mathrm{S_2}]}$ d. $K_c = \frac{2[\mathrm{H_2S}]}{2[\mathrm{H_2}][\mathrm{S_2}]}$

7. Initially, $[H_2S] = 0.1$ M, $[H_2] = 0.01$ M, and $[S_2] = 0.01$ M. Which way will the reaction shift and why?
   a. The reaction will shift to make more products because Q < K.
   b. The reaction will shift to make more reactants because Q < K.
   c. The reaction will shift to make more products because Q > K.
   d. The reaction will shift to make more reactants because Q > K.

8. If, at a given point in the reaction, the value for the reaction quotient *Q* is determined to be 2.5×10^8, which of the following is occurring?
   a. The concentration of the reactant is decreasing while the concentration of the products is increasing.
   b. The concentration of the reactant is increasing while the concentration of the products is decreasing.
   c. The system has passed the equilibrium point, and the concentration of all species involved in the reaction will remain constant.
   d. The concentrations of all species involved are changing at the same rate.



9. The picture above shows the species initially present in a 1.0 L container. The chemical reaction shown below takes place.

$$\mathrm{A + B \leftrightarrow C} \qquad K_c = 2.3 \times 10^{-3}$$

Which of the following statements is true?

a. The reaction shifts towards the products to reach equilibrium.

b. The reaction shifts towards the reactants to reach equilibrium.

c. The reaction mixture is at equilibrium.

d. The direction of shift cannot be determined from the information given.

10. Consider the following reaction:

$$\mathrm{CH_4}(g) + 2\ \mathrm{H_2S}(g) \leftrightarrow \mathrm{CS_2}(g) + 4\ \mathrm{H_2}(g)$$

1.00 M CH_4, 1.00 M CS_2, 2.00 M H_2S and 2.00 M H_2 are mixed in a reaction vessel at 960°C. At this temperature, the reaction will make more methane, CH_4, and more hydrogen sulfide, H_2S. What is a possible K for this reaction?

c. K = 16 c. K = 4
d. K = 8 d. K = 1

Equilibrium: K vs. Q *with Legos™!*

$$H_2(g) + Cl_2(g) \leftrightarrow 2\ HCl(g)$$

yellow + pink ↔ yellow-pink

At Equilibrium: What does this reaction look like at equilibrium? Draw/write below. Calculate K.

Example #1: Record the initial reactant/product quantities below.

1. Is the reaction currently at equilibrium? Why/ why not? If it's not at equilibrium, change the Legos until it is!

2. If you weren't at equilibrium, what did you have to do to make it look like equilibrium? (Use the words 'reactants' and 'products' in your answer.)

3. Which direction did the reaction seem to move? Right towards products or left towards reactants?

Example #2: Record the initial reactant/product quantities below.

1. Is the reaction currently at equilibrium? Why/ why not? If it's not at equilibrium, change the Legos until it is!

2. If you weren't at equilibrium, what did you have to do to make it look like equilibrium? (Use the words 'reactants' and 'products' in your answer.)

3. Which direction did the reaction seem to move? Right towards products or left towards reactants?

Unit 2 MC Practice

1. What total gas volume (in liters) at 520°C and 880 torr would result from the decomposition of 33 g of potassium bicarbonate, $KHCO_3$, according to the equation:

$$2\ KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(g)$$

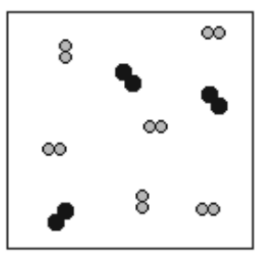
a. 10. L b. 19 L c. 37 L d. 56 L

$$\ldots C_{10}H_{12}O_4S(s) + \ldots O_2(g) \rightarrow \ldots CO_2(g) + \ldots SO_2(g) + \ldots H_2O(s)$$

2. When the equation above is balanced and all coefficients are reduced to their lowest whole-number terms, the coefficient for $O_2(g)$ is

a. 7 b. 12 c. 14 d. 28

3. Consider the contents of the beaker shown below: what state/type of matter is being represented?



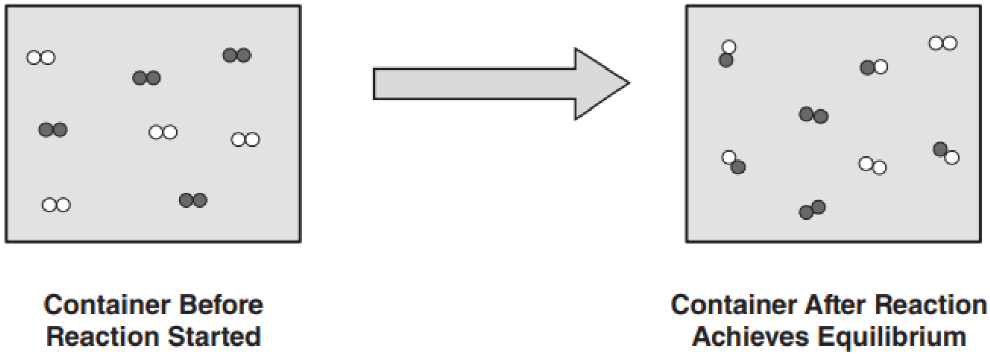
a. element/ gas c. heterogeneous mixture/ liquid

b. compound/ solid d. homogeneous mixture /gas

4. What volume of 0.150 molar HCl is required to neutralize 25.0 milliliters of 0.120 molar $Ba(OH)_2$?

a. 20.0 mL b. 30.0 mL c. 40.0 mL d. 60.0 mL

5. Below are 1.0 L containers for the initial and equilibrium condition for the reaction $A_2(g) + D_2(g) \rightleftharpoons 2\ AD(g)$.



Calculate the magnitude of the equilibrium constant for the reaction.

a. 0.25 b. 1.0 c. 4.0 d. 8.0

6. What is the complete ionic equation for the acid-base reaction that occurs when two aqueous solutions of hydrobromic acid, HCl, and calcium hydroxide, $Ca(OH)_2$, are mixed?

a. $\mathrm{H^+}(aq) + \mathrm{OH^-}(aq) \rightarrow \mathrm{H_2O}(l)$

b. $2\ \mathrm{H^+}(aq) + 2\ \mathrm{OH^-}(aq) \rightarrow 2\ \mathrm{H_2O}(l)$

c. $2\ \mathrm{HBr}(aq) + \mathrm{Ca(OH)_2}(aq) \rightarrow \mathrm{CaBr_2}(aq) + 2\ \mathrm{H_2O}(l)$

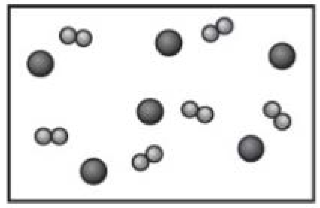
d. $2\ \mathrm{H^+}(aq) + 2\ \mathrm{Br^-}(aq) + \mathrm{Ca^{2+}}(aq) + 2\ \mathrm{OH^-}(aq) \rightarrow \mathrm{Ca^{2+}}(aq) + 2\ \mathrm{Br^-}(aq) + 2\ \mathrm{H_2O}(l)$

7. Calculate the weight of $KClO_3$ that would be required to produce 29.5 L of oxygen at 127°C and 760. torr.

$$2\ \mathrm{KClO_3}(s) \rightarrow 2\ \mathrm{KCl}(s) + 3\ \mathrm{O_2}(g)$$

a. 7.82 g b. 14.6 g c. 24.4 g d. 73.5 g

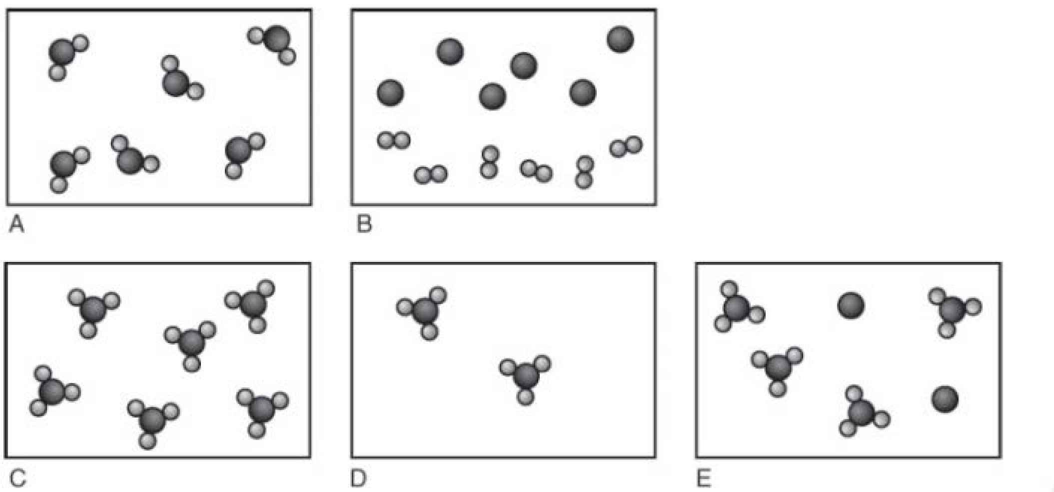
8. Consider a mixture of sulfur atoms and dioxygen molecules in a closed container below:



The equation for the reaction is:

$$2\ \mathrm{S}(g) + 3\ \mathrm{O_2}(s) \rightarrow 2\ \mathrm{SO_3}(g)$$

Which of the following images best represents what will be in the container after the reaction goes to completion?



9. Consider the balanced chemical equation $H_2SO_4(aq) + Fe(s) \rightarrow Fe_2(SO_4)_3(s) + H_2(g)$

Which of the following is a true statement about what happens during this reaction?

a. Oxygen is oxidized, sulfur is reduced.
b. Sulfur is oxidized, oxygen is reduced.
c. Hydrogen is oxidized, iron is reduced.
d. Iron is oxidized, hydrogen is reduced.

10. If 0.40 mol of H_2 and 0.15 mol of O_2 were to react as completely as possible to produce H_2O, what mass of reactant would remain?

a. 0.20 g of H_2 b. 0.40 g of H_2 c. 3.2 g of O_2 d. 4.0 g of O_2

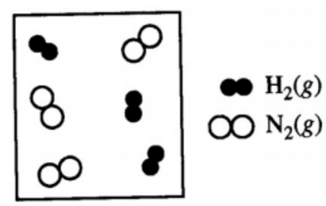
11. All chemical equilibriums must have:

a. $K_{eq} = 1$
c. rate forward = rate reverse
b. [reactants] = [products]
d. mass of reactants = mass of products

12. Consider the following reversible reaction. In a 3.00 liter container, the following amounts are found in equilibrium at 400°C: 0.0420 moles N_2, 0.516 moles H_2 and 0.0357 moles NH_3. Calculate the equilibrium constant, K_c.

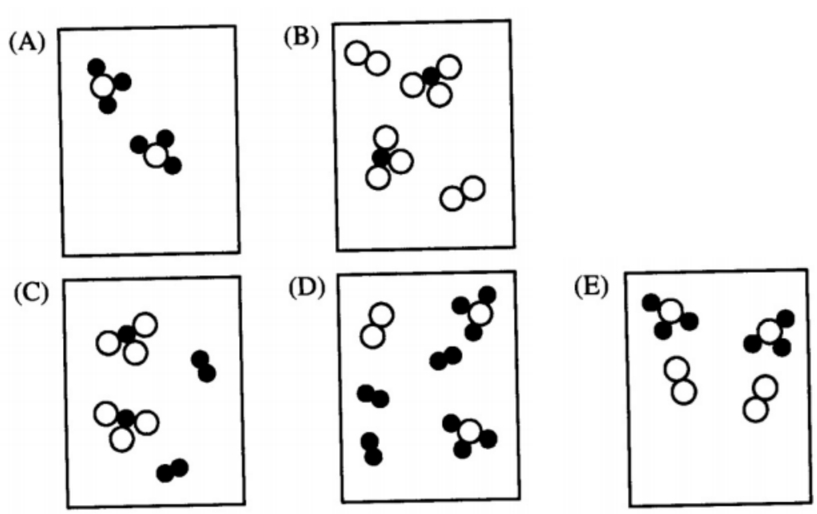
$$N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$$

a. 0.202　b. 1.99　c. 4.94　d. 16.0



13. The diagram above represents H_2(g) and N_2(g) in a closed container. Which of the following diagrams would represent the results if the reaction below was to proceed as far as possible?

$$N_2(g) + 3\,H_2(s) \rightarrow 2\,NH_3(g)$$



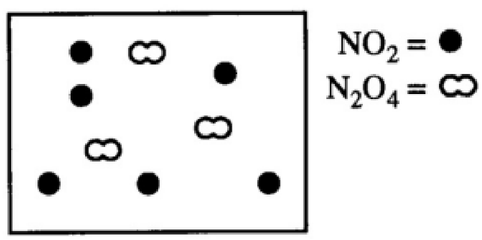
14. A balanced equation for the reaction of copper metal with nitric acid is shown below. Which of the following represents a true statement about the reaction?

$$3\ \mathrm{Cu}(s) + 8\ \mathrm{H}^+(aq) + 2\ \mathrm{NO}_3^-(aq) \rightarrow 3\ \mathrm{Cu}^{2+}(aq) + 4\ \mathrm{H_2O}(l) + 2\ \mathrm{NO}(g)$$

a. The oxidation state of nitrogen changed from +5 to +2.
b. Hydrogen ions are oxidized to form H_2O(l).
c. The oxidation state of oxygen changes from −1 to −2.
d. Copper metal is reduced to a copper (II) ion.

15. When 2.00 g of H_2 reacts with 32.0 g of O_2 in an explosion, the final gas mixture will contain:

a. H_2, H_2O, and O_2 b. H_2 and H_2O only   c. O_2 and H_2O only   d. H_2O only



16. The diagram above represents a mixture of NO_2(g) and N_2O_4(g) in a 1.0 L container at a given temperature. The two gases are in equilibrium according to the equation

$$2\ \mathrm{NO_2}(g) \rightleftharpoons \mathrm{N_2O_4}(g)$$

Which of the following must be true about the value of the equilibrium constant, *K*, for the reaction at this temperature?

a. K = 0   b. 0 < K < 1   c. K = 1   d. K > 1

17. The spectator ions in the following reaction is/are:

$$\mathrm{Na_2CO_3}(aq) + \mathrm{Ba(NO_3)_2}(aq) \rightarrow \mathrm{BaCO_3}(s) + 2\ \mathrm{NaNO_3}(aq)$$

a. Ba^{2+} and CO_3^{2-} b. CO_3^{2-} only   c. Na^+ and NO_3^- d. Na^+ only

18. When the system A + B $\rightleftharpoons$ C + D is at equilibrium,

a. the forward reaction has stopped

b. the reverse reaction has stopped

c. both the forward and reverse reactions have stopped

d. neither the forward nor reverse reactions have stopped

19. Balance the following equation with the smallest whole number coefficients possible. Select the number that is the sum of the coefficients in the balanced equation.

$$\_\_\_\mathrm{KClO_3} \rightarrow \_\_\_\mathrm{KCl} + \_\_\_\mathrm{O_2}$$

a. 3 b. 5 c. 6 d. 7

20. The reaction of 7.8 g benzene, C_6H_6, with excess HNO_3 resulted in 0.90 g of H_2O. What is the percent yield? (Molar masses (g mol^{-1}): C_6H_6 = 78; HNO_3 = 63; $C_6H_5NO_2$ = 123; H_2O = 18.)

$$\mathrm{C_6H_6 + HNO_3 \rightarrow C_6H_5NO_2 + H_2O}$$

a. 90.% b. 50.% c. 12% d. 2.0%

$$2\ \mathrm{SO_3}(s) \rightleftharpoons 2\ \mathrm{SO_2}(g) + \mathrm{O_2}(g)$$

21. The equilibrium constant expression, K_c, for the system described by the above equation is:

a. $\dfrac{[\mathrm{SO_3}]^2}{[\mathrm{SO_2}]^2[\mathrm{O_2}]}$ b. $[\mathrm{SO_2}]^2[\mathrm{O_2}]$ c. $\dfrac{[\mathrm{SO_2}]^2[\mathrm{O_2}]}{[\mathrm{SO_3}]^2}$ d. $\dfrac{1}{[\mathrm{SO_2}]^2[\mathrm{O_2}]}$

22. At 445°C, K_c for the following reaction is 0.020.

$$2\ \mathrm{HI}(g) \rightleftharpoons \mathrm{H_2}(g) + \mathrm{I_2}(g)$$

A mixture of H_2, I_2, and HI in a vessel at 445°C has the following concentrations: [HI] = 2.0 M, [H_2] = 0.50 M and [I_2] = 0.10 M. Which of the following statements concerning the reaction quotient, Q_c, is **TRUE** for the above system?

a. $Q_c < K_c$; more H_2 and I_2 will be produced

b. $Q_c > K_c$; more H_2 and I_2 will be produced

c. $Q_c < K_c$; more HI will be produced

d. $Q_c > K_c$; more HI will be produced

23. How many grams of nitric acid, HNO_3, can be prepared from the reaction of 138 g of NO_2 with 54.0 g H_2O according to the equation below?

$$3\ \mathrm{NO_2} + \mathrm{H_2O} \rightarrow 2\ \mathrm{HNO_3} + \mathrm{NO}$$

a. 108 b. 126 c. 189 d. 279

24. For a specific reaction, which of the following statements can be made about *K*, the equilibrium constant?

a. It always remains the same at different reaction conditions.

b. It increases if the concentration of one of the products is increased.

c. It changes with changes in the temperature.

d. It increases if the concentration of one of the reactants is increased.

25. When 8.0 g of N_2H_4 (32 g mol^{-1}) and 92 g of N_2O_4 (92 g mol^{-1}) are mixed together and react according to the equation below, what is the maximum mass of H_2O that can be produced?

$$2\ \mathrm{N_2H_4}(g) + \mathrm{N_2O_4}(g) \rightarrow 3\ \mathrm{N_2}(g) + 4\ \mathrm{H_2O}(g)$$

a. 9.0 g b. 18 g c. 36 g d. 72 g

$$3\ \mathrm{NaHCO_3}(aq) + \mathrm{C_6H_8O_7}(aq) \rightarrow 3\ \mathrm{CO_2}(g) + 3\ \mathrm{H_2O}(s) + \mathrm{Na_3C_6H_5O_7}(aq)$$

26. Which of the following statements is true about the reaction shown above?

a. 22.4 L of CO_2(g) are produced for every liter of $C_6H_8O_7$(aq) reacted.

b. 1 mole of water is produced for every mole of carbon dioxide produced.

c. 6.02×10^{23} molecules of $Na_3C_6H_5O_7$(aq) are produced for every mole of $NaHCO_3$(aq) used.

d. 54 g of water are produced for every mole of $NaHCO_3$(aq) used.

27. Which of the following applies to a chemical equilibrium?

    I. Forward and reverse reaction rates are equal.
    II. Equilibrium can be achieved from either direction.
    III. Macroscopic properties are constant.

    a. I only b. I and II only c. II and III only d. I, II, and II

28. Calculate the mass of hydrogen formed when 27 g of aluminum reacts with excess hydrochloric acid, HCl, according to the balanced equation below.

$$2\ \mathrm{Al} + 6\ \mathrm{HCl} \rightarrow 2\ \mathrm{AlCl_3} + 3\ \mathrm{H_2}$$

    a. 1.5 g b. 2.0 g c. 3.0 g d. 6.0 g

$$2\ \mathrm{MnO_2} + 4\ \mathrm{KOH} + \mathrm{O_2} + \mathrm{Cl_2} \rightarrow 2\ \mathrm{KMnO_4} + 2\ \mathrm{KCl} + 2\ \mathrm{H_2O}$$

29. For the reaction above, there is 100. g of each reactant available. Which reagent is the limiting reagent? (Molar masses (g/mol): MnO_2 = 86.9; KOH = 56.1; O_2 = 32.0; Cl_2 = 70.9.)

    a. MnO_2 b. O_2 c. KOH d. Cl_2

30. How is the reaction quotient used to determine whether a system is at equilibrium?

    a. At equilibrium, the reaction quotient is undefined.
    b. The reaction is at equilibrium when $Q < K_{eq}$.
    c. The reaction is at equilibrium when $Q > K_{eq}$.
    d. The reaction is at equilibrium when $Q = K_{eq}$.

31. How many grams of H_2O will be formed when 32.0 g H_2 is allowed to react with 16.0 g O_2 according to

$$\_\_\_H_2 + \_\_\_O_2 \rightarrow \_\_\_H_2O$$

a. 9.00 g b. 16.0 g c. 18.0 g d. 32.0 g

32. Methane, CH_4, and hydrogen sulfide form when hydrogen reacts with carbon disulfide, CS_2. Identify the excess reagent and calculate how much remains after 36 L of H_2 reacts with 12 L of CS_2 at constant temperature and pressure.

$$4\ \mathrm{H_2}(g) + \mathrm{CS_2}(g) \rightarrow \mathrm{CH_4}(g) + 2\ \mathrm{H_2S}(g)$$

a. CS_2 is in excess, 3 L remains

b. H_2 is in excess, 9 L remains

c. CS_2 is in excess, 6 L remains

d. H_2 is in excess, 12 L remains

$$\mathrm{Ni^+}(aq) + 2\ \mathrm{NH_3}(aq) \leftrightarrow [\mathrm{Ni(NH_3)_2}]^+(aq) \quad K = 3.2 \times 10^{-7}$$

33. A student mixes 50.0 mL of 0.1 M Ni^+ (in the form of $NiNO_3$) and 50.0 mL of 0.3 M NH_3, and the solution is allowed to reach equilibrium (as shown in the reaction above). List the following species from least to greatest concentrations at equilibrium: Ni^+(aq), NH_3(aq), and $[Ni(NH_3)_2]^+$.

a. $[Ni^+]_{eq}$(aq) < $[NH_3]_{eq}$ (aq) < $[[Ni(NH_3)_2]^+]_{eq}$

b. $[NH_3]_{eq}$ (aq) < $[Ni^+]_{eq}$(aq) < $[[Ni(NH_3)_2]^+]_{eq}$

c. $[[Ni(NH_3)_2]^+]_{eq}$ < $[Ni^+]_{eq}$(aq) < $[NH_3]_{eq}$ (aq)

d. $[Ni^+]_{eq}$(aq) < $[[Ni(NH_3)_2]^+]_{eq}$ < $[NH_3]_{eq}$ (aq)

Unit 2: AP Quiz Free Response Practice #1 [2003 Form B #2, modified, 6 points]

1. Solid iron (III) oxide, Fe_2O_3, can be reduced with gaseous carbon monoxide, producing iron metal and carbon dioxide.
   a. In an experiment, a student combines a 16.2 L sample of CO at 1.50 atm and 200.°C with 15.39 g of Fe_2O_3.
      i. What is the balanced chemical equation for this reaction? [2 points]

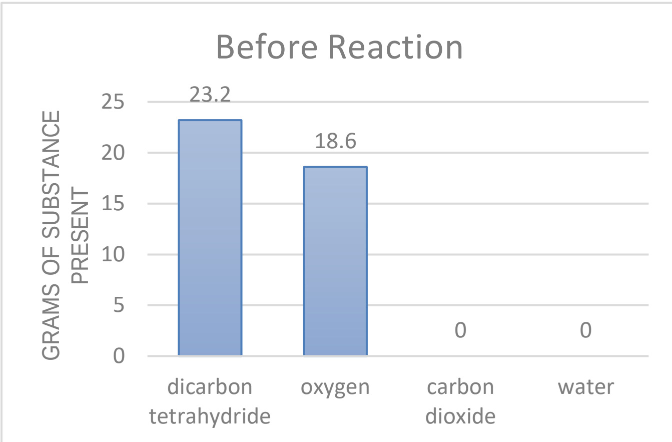
      ii. How many moles of CO(g) are available for the reaction? [1 point]

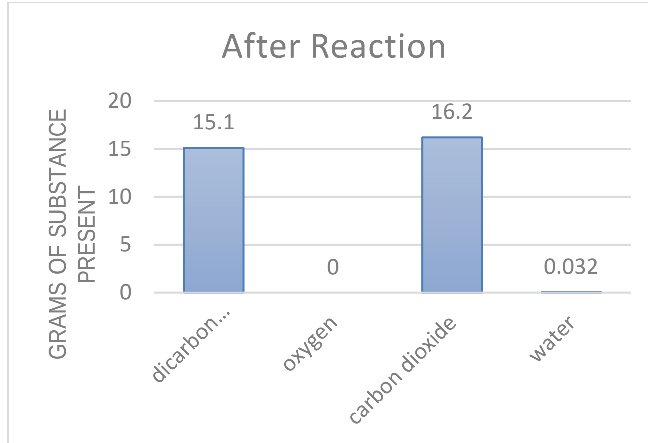
      iii. What is the limiting reactant for the reaction? Justify your answer with calculations. [2 points]

      iv. How many moles of Fe(s) are formed in the reaction? [1 point]

Unit 2: AP Quiz Free Response Practice #2 (10 pts)

In a lab experiment, Skylar completes the combustion of 23.2 g of dicarbon tetrahydride, C_2H_4, in the presence of 18.6 g of oxygen gas. Her data can be seen in the graphs below.





a. Write and balance the equation for this reaction. [2 POINTS]

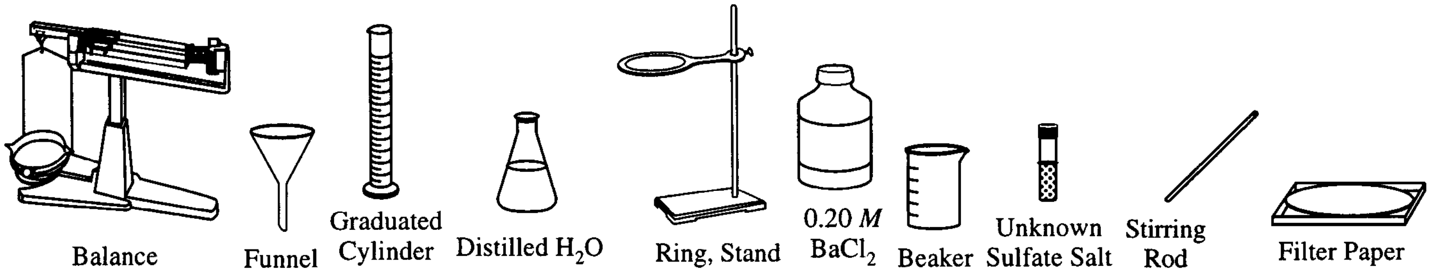
b. Which substance is the limiting reactant? How do you know? [2 POINTS]

c. Which substance is the excess reactant? How do you know? [2 POINTS]

d. What is the theoretical yield, in grams, of carbon dioxide for Skylar's experiment? [1 POINT]

e. Skylar's twin, Ella, tries to replicate Skylar's reaction of 23.2 g of dicarbon tetrahydride with 18.6 g of oxygen. However, when Ella does it, only 0.21 moles of carbon dioxide are formed. What is Ella's percent yield? [2 POINTS]

Unit 2: AP Free Response Practice #1 [8 points]



1. A student is performing an experiment to determine the mass percent of sulfate, SO_4^{2-}, in an unknown soluble sulfate salt. The equipment shown above is available for the experiment. A drying oven is also available. Note: barium sulfate, $BaSO_4$, is insoluble in water.

   a. Briefly list the steps needed to carry out this experiment. [2 points]

   b. What experimental data needs to be collected to calculate the mass percent of sulfate in the unknown? [2 points]

   c. Starting with an initial sulfate salt sample of 8.52 g, the student is able to recover 07.52 g of barium sulfate, $BaSO_4$. What is the mass percent of sulfate, SO_4^{2-}, in the original sample? [3 points]

   d. If the student fails to dry the barium sulfate sample to constant mass, would this cause the calculated mass percent of sulfate to be less than, equal to, or greater than the actual mass percent? Justify your answer. [1 point]

Unit 2: AP Free Response Practice #2 [6 points]

2. A 3.00 g sample of impure solid, containing sodium chloride, NaCl and an inert substance, is dissolved in 200. mL of water. An excess of aqueous silver ions is added, and reaction occurs, producing a white precipitate of silver chloride, AgCl. After filtering and drying, it is found that 0.450 g of solid is produced.

   a. Write the balanced equation for the reaction which occurs, including states. [2 points]

   b. Write the balanced net ionic equation for the precipitation reaction, including states. [1 point]

   c. Calculate the mass of chloride ions, Cl^-, in the original impure sample. [2 points]

   d. What percent by mass of the original sample was due to chloride ions? [1 point]

Unit 2: AP Free Response Practice #3 [2003 Form B #2, modified, 5 points]

3. In a reaction vessel, 0.600 mol of $Ba(NO_3)_2(s)$ and 0.300 mol of $H_3PO_4(aq)$ are combined with deionized water to a final volume of 2.00 L. The reaction represented below occurs.

$$\_\_\_\_\_ \; \mathrm{Ba(NO_3)_2}(aq) + \_\_\_\_\_ \; \mathrm{H_3PO_4}(aq) \rightarrow \_\_\_\_\_ \; \mathrm{Ba_3(PO_4)_2}(s) + \_\_\_\_\_ \; \mathrm{HNO_3}(aq)$$

a. Balance the equation above by writing the correct coefficients in the blanks provided. [1 point]

b. Calculate the mass of $Ba_3(PO_4)_2(s)$ formed. [2 points]

c. Calculate the moles of nitric acid, HNO_3, formed. [1 point]

d. How many moles of excess reactant are left over once the reaction stops? [1 point]

Unit 2: AP Free Response Practice #4 [8 points]

3. 2.54 g of beryllium chloride, $BeCl_2$, are completely dissolved into 50.00 mL of water inside a beaker.

   a) Calculate the concentration of beryllium ions, $Be^{2+}(aq)$, chloride ions, $Cl^{-}(aq)$, and total ions in the beaker. [*2 points*]

A solution of 0.850 M lead (II) nitrate, $Pb(NO_3)_2$, is then titrated into the beaker, causing a precipitate of lead (II) chloride, $PbCl_2$, to form.

b) Identify the net ionic reaction occurring in the beaker. [*1 point*]

c) Determine the volume of lead (II) nitrate, $Pb(NO_3)_2$, in milliliters that must be added to the beaker to cause the maximum precipitate to form. [*2 points*]

d) What is the theoretical yield of precipitate? [*2 points*]

e) A student performing this lab has a percent error of 7.45%. They claim that the source of their error was dissolving the beryllium chloride in only 45.00 mL of water. Do you agree with the student? Justify your response. [*1 point*]

Unit 2: AP Free Response Practice #5 (2008 FR #3, modified) [10 points]

4. A student adds 150. mL of a 0.030 M $Pb(NO_3)_2$ solution to 125 mL of 0.100 M sodium iodide, NaI, solution.

   a. A chemical reaction takes place, and a precipitate is formed.

      i. Write the balanced equation for the reaction, including states. [*2 point]*

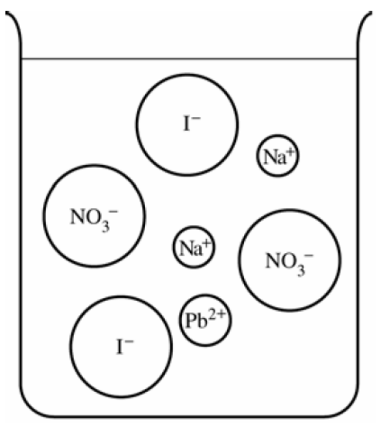
      ii. Write the balanced net ionic equation for the reaction, including states. [*1 point*]

   b. Calculate the number of moles of each reactant. [*2 points*]

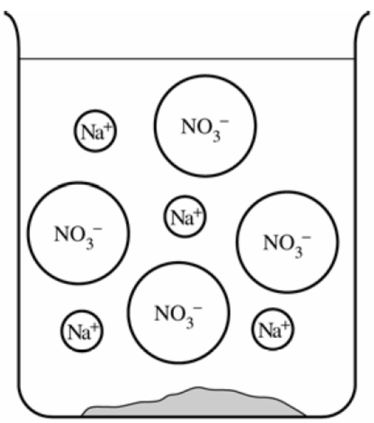
   c. Identify the limiting reactant. Show calculations to support your identification. [*1 point*]

d. Calculate the molar concentration of $NO_3^-(aq)$ in the mixture after the reaction is complete. [*2 points*]

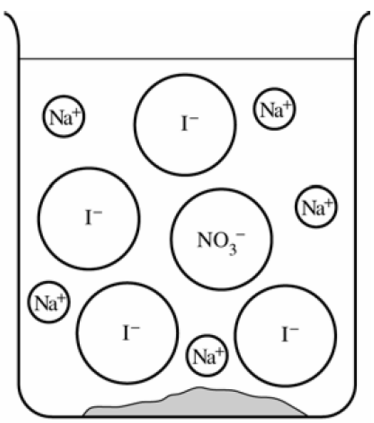
e. Circle the diagram below that best represents the results after the mixture reacts as completely as possible. Explain the reasoning used in making your choice. [*2 points*]



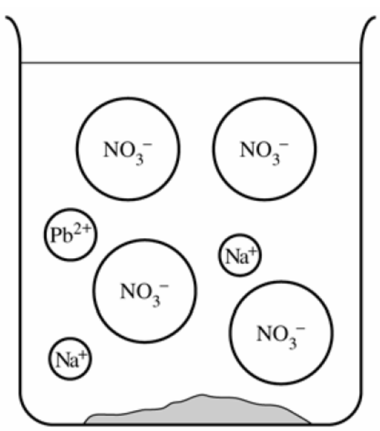
No Precipitate



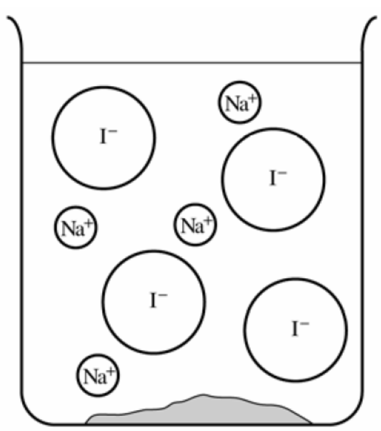
Solid PbI_2



Solid PbI_2

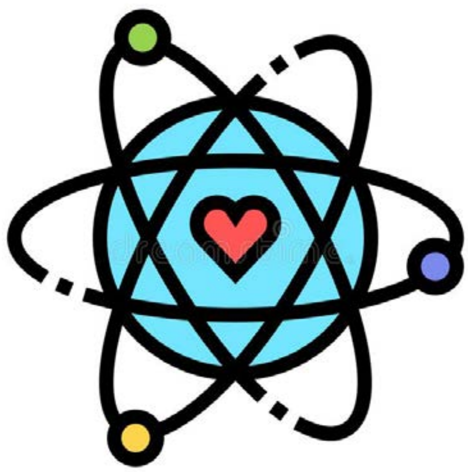


Solid PbI_2



Solid $Pb(NO_3)_2$

AP Chemistry FTW!



Unit 3:

Electrochemistry

AP Chemistry Unit 3 Objectives

BIG IDEA 3 - Changes in matter involve the rearrangement and/or the reorganization of atoms and /or the transfer of electrons.

- Enduring Understanding 3.A: Chemical changes are represented by a balanced chemical equation that identifies the ratios with which reactants react and products form.
- Enduring Understanding 3.B: Chemical reactions can be classified by considering what the reactants are, what the products are, or how they change from one into the other. Classes of chemical reactions include synthesis, decomposition, acid-base, and oxidation-reduction reactions.
- Enduring Understanding 3.C: Chemical and physical transformation may be observed in several ways and typically involve a change in energy.

BIG IDEA 6 - Any bond or intermolecular attraction that can be formed can be broken. These processes are in a dynamic competition, sensitive to initial conditions and external perturbations.

- Enduring Understanding 6.A: Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal.
- Enduring Understanding 6.B: Systems at equilibrium are responsive to external perturbations, with the response leading to a change in the composition of the system.
- Enduring Understanding 6.D: The equilibrium constant is related to temperature and the difference in Gibbs free energy between reactants and products.

Electrochemistry: *Shockingly* Awesome!

Electrochemistry: the study of redox reactions that _____________ or _____________ an electric current.

- Electrochemical ________: where the conversion between chemical potential energy and electrical energy takes place (a fancy way of talking about a chemical reaction that involves the transfer of electrons, aka __________!)
- There are two electrochemical cells you need to know:
  - ______________ (or galvanic) cell: spontaneously undergoes a redox reaction and can be used to produce an electric current if the oxidation and reduction half-reactions are separated.
  - ______________ cell: undergoes a redox reaction only if an electrical current is applied. (i.e. non-spontaneously)

Oxidation-Reduction (Redox) Review

Oxidation–reduction (redox) reactions: where electrons are transferred from one atom to another.

- If a substance accepts an electron, it is____________________.
- If a substance loses an electron, it is ____________________.
- Electrons are always transferred from the species that is oxidized to the species that is reduced.

➔ Reduction (gain in electrons) + Oxidation (loss of electrons) = Redox!

How to Identify What is Oxidized or Reduced in a Reaction

Once you have identified a redox reaction by the change in oxidation state, now you can tell what was oxidized or reduced!

a. A substance that has the element that has been ________________ (LOST electrons) will have an oxidation number that becomes more ________________ (or less negative).

b. A substance that has the element that has been ________________ (GAINED electrons) will have an oxidation number that becomes more ________________ (or less positive).

Balancing Redox Reactions: We split redox reactions into two separate ___________________

- The oxidation half-reaction has electrons as a _______________.
- The reduction half-reaction has electrons as a _______________.

Oxidation Half-Reaction	**Reduction Half-Reaction**
$\mathrm{Zn}(s) \rightarrow \mathrm{Zn}^{2+}(aq) + 2\,e^-$	$\mathrm{Fe}^{2+}(aq) + 2\,e^- \rightarrow \mathrm{Fe}(s)$

Reduction Potential Values: a measure of the tendency of a chemical species to ________ electrons (aka be reduced). Reduction potential is measured in volts (V), or millivolts (mV). Reduction potential values are measured ________________to the potential for H^+ to _____________ electrons to become H_2.

How to read a Reduction Potential Chart

1. The more ______________________ $E°_{red}$ (V), the ______________ likely it is to be reduced (gain electrons).
2. The more ______________________ $E°_{red}$ (V), the ______________ likely it is to be reduced (gain electrons).

Table 1. Standard Reduction Potentials at 25 ° C

Reduction Half Reaction			E°
$Co^{3+}(aq) + e^-$	→	$Co^{2+}(aq)$	1.92 V
$Au^+(aq) + e^-$	→	$Au(s)$	1.69 V
$Mn^{3+}(aq) + e^-$	→	$Mn^{2+}(aq)$	1.54 V
$Au^{3+}(aq) + 3e^-$	→	$Au(s)$	1.498 V
$Ag^+(aq) + e^-$	→	$Ag(s)$	0.80 V
$Cu^+(aq) + e^-$	→	$Cu(s)$	0.52 V
$Cu^{2+}(aq) + 2e^-$	→	$Cu(s)$	0.34 V
$2H^+(aq) + 2e^-$	→	$H_2(g)$	0.00 V
$Pb^{2+}(aq) + 2e^-$	→	$Pb(s)$	−0.13 V
$Ni^{2+}(aq) + 2e^-$	→	$Ni(s)$	−0.26 V
$Co^{2+}(aq) + 2e^-$	→	$Co(s)$	−0.28 V
$Cd^{2+}(aq) + 2e^-$	→	$Cd(s)$	−0.41 V
$Zn^{2+}(aq) + 2e^-$	→	$Zn(s)$	−0.76 V
$Mn^{2+}(aq) + 2e^-$	→	$Mn(s)$	−1.18 V
$Al^{3+}(aq) + 3e^-$	→	$Al(s)$	−1.66 V

When comparing two species:

- Substance with the **greater** (more positive/less negative) **reduction** potential will be ______________
- Substance with the **smaller** (less positive/more negative) **reduction** potential will be ______________

3. Switching the ___________ of the reduction potential will give you the oxidation potential ($E°_{oxidation}$)

$$-E_{red} = E_{ox}$$

Let's try!

1. Is the following reaction a redox reaction? *If* the reaction is redox, identify what was oxidized and reduced!

$$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$$

2. Rank these metals in order of the most easily oxidized to the least easily oxidized: Cu, Ni, Ag, Pb

Directions: Write the half reaction for each of the follow redox reactions, identify what is being reduced and what is being oxidized, then write the correction reduction or oxidation potential next to each half-reaction.

3. $Cd^{2+} + Zn(s) \rightarrow Cd(s) + Zn^{2+}$

4. $2\ Ag^{+} + Ni(s) \rightarrow Ni^{2+} + 2\ Ag(s)$

Multiple Choice Practice FTW!

1. A balanced equation for the reaction of copper metal with nitric acid is shown below. Which of the following represents a true statement about the reaction?

$$3\ \mathrm{Cu}(s) + 8\ \mathrm{H}^{+}(aq) + 2\ \mathrm{NO}_3^{-}(aq) \rightarrow 3\ \mathrm{Cu}^{2+}(aq) + 4\ \mathrm{H_2O}(l) + 2\ \mathrm{NO}(g)$$

   a. The oxidation state of nitrogen changed from +5 to +2.
   b. Hydrogen ions are oxidized to form $H_2O(l)$.
   c. The oxidation state of oxygen changes from −1 to −2.
   d. Copper metal is reduced to a copper (II) ion.

2. A strip of metal X is placed into a solution containing Y^{2+} ions and no reaction occurs. When metal X is placed in a separate solution containing Z^{2+} ions, metal Z starts to form on the strip. Which of the following choices organizes the reduction potentials for metals X, Y, and Z from greatest to least?

a. X > Y > Z   b. Y > Z > X   c. Z > X > Y   d. Y > X > Z

$$2\ H_2O(l) + 4\ MnO_4^-(aq) + 3\ ClO_2^-(aq) \rightarrow 4\ MnO_2(aq) + 3\ ClO_4^-(aq) + 4\ OH^-(aq)$$

3. Which species is reduced in the reaction represented above?

a. MnO_2 b. ClO_2^- c. MnO_4^- d. ClO_4^-

4. In the reaction below, a piece of solid nickel is added to a solution of potassium dichromate.

$$14\ H^+(aq) + Cr_2O_7^{2-}(aq) + 3\ Ni(s) \rightarrow 2\ Cr^{3+}(aq) + 3\ Ni^{2+}(aq) + 7\ H_2O(l)$$

Which species is being oxidized and which is being reduced?

	Oxidized	Reduced
a.	$Cr_2O_7^{2-}(aq)$	$Ni(s)$
b.	$Cr^{3+}(aq)$	$Ni^{2+}(aq)$
c.	$Ni(s)$	$Cr_2O_7^{2-}(aq)$
d.	$Ni^{2+}(aq)$	$Cr^{3+}(aq)$

Balancing Redox Reactions by the Half-Reaction Method

In this method, the reaction is broken down into ______ half-reactions, one for oxidation and another for reduction.

- Each half-reaction includes ______ AND is balanced for its ___________.
- Then the two half-reactions are adjusted so the electrons lost and gained will be ___________ when combined.

How to ACE Redox Balancing

1. Write the oxidation and reduction half-reactions (without electrons).
2. Balance half-reactions using the **ACED** technique:
   - _____________: balance atoms first
   - _______________: balance the charge on both sides of the reaction equal by adding electrons to the side which is more positive
   - _______________: multiply as needed so both half-reactions have the same # of electrons exchanged.
   - _____________: Add the two half-reactions, cancelling anything that appears on both sides of the equation. (*Don't forget to add states in your final reaction!*)

Examples

Example #2: $\mathrm{Cu^{2+}}(aq) + \mathrm{Na}(s) \rightarrow \mathrm{Cu}(s) + \mathrm{Na^{+}}(aq)$

Example #1: $\mathrm{O_2}(g) + \mathrm{Al}(s) \rightarrow \mathrm{O^{2-}}(aq) + \mathrm{Al^{3+}}(aq)$

Yum, Electrifying Practice!

Now You Try! Balance the following redox reactions.

1. $O^{2-}(aq) + F_2(g) \rightarrow O_2(g) + F^-(aq)$

2. $Al(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Al^{3+}(aq)$

3. $Ag^+(aq) + Cu(s) \rightarrow Ag(s) + Cu^{2+}(aq)$

4. $Hg(l) + Au^{3+}(aq) \rightarrow {Hg_2}^{2+}(l) + Au(s)$

5. $Zn^{2+}(aq) + K(s) \rightarrow Zn(s) + K^+(aq)$

Redox Titrations: Electrochemistry Stoichiometry in Action!

Important Caveat: at this point in time, we have only performed and analyzed ________________ titrations, but a redox titration is _______ **AN ACID-BASE TITRATION.** Although much of our language and calculations will be the same, please be cognizant of the difference and DO NOT use acid-base terminology (especially in your lab report 😉).

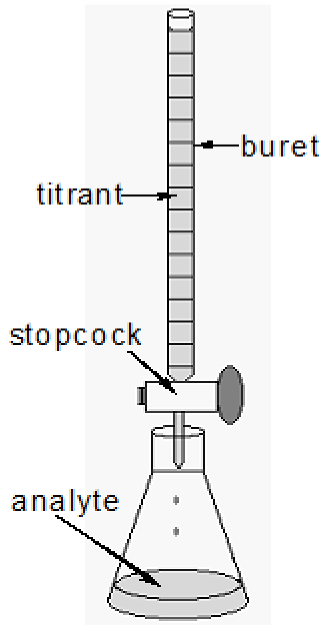
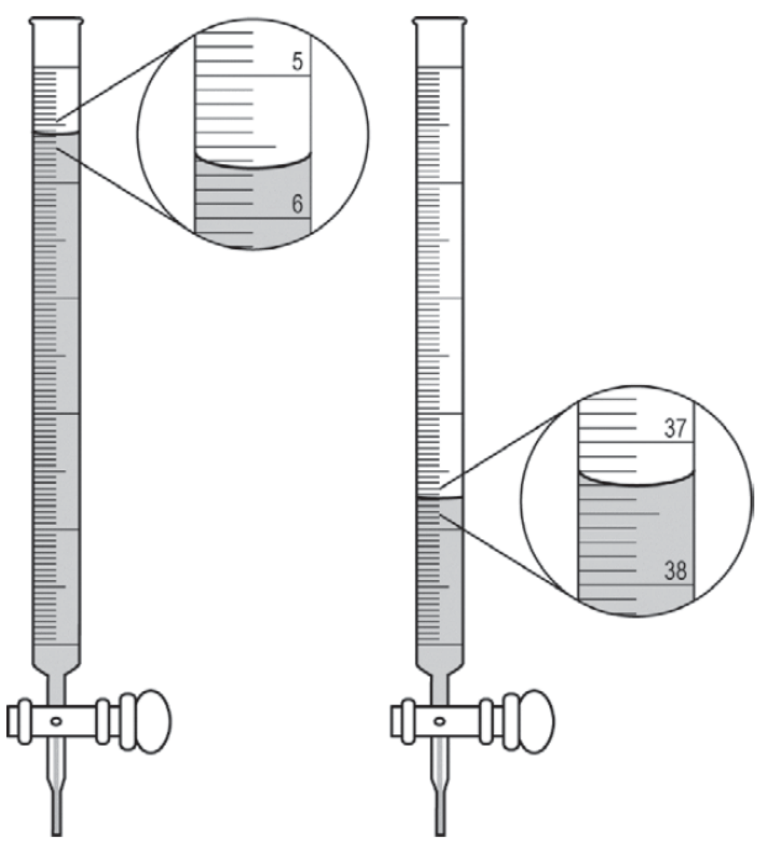


Figure 1: Titration Setup

Goal of Titration: determine the unknown concentration of the analyte!

A substance in a solution of ____________ concentration (the____________, usually in a buret) is reacted with another substance in a solution of ______________ concentration (the **analyte**, usually in a flask or beaker)

- **Equivalence point**: the point at which the moles of each reactant are ***stoichiometrically equal*** to each other in solution. (*However*, this is _______ where the two reactants have neutralized each other: that only happens in an acid-base titration!)
    - **Equivalence point is _________ equilibrium!!!**
- **End point**: the point of the titration where an ________________ changes color (if thoughtfully chosen, this occurs approximately at the equivalence point as a visual indication of the point at which the stoichiometric equivalent has been added).



Quick Reminder: How to Read a Buret

Burets, unlike most glassware used for precise volume measurements, are read from the ________ down, not the bottom up!

Example: The image to the left shows a buret of $KMnO_4$ before titration begins (leftmost) and at the end point (rightmost).

1. What volume of solution was initially present in the buret? ________________
2. What volume was present at the end point of the titration? _______________
3. How much $KMnO_4$ solution was added to the flask?

Let's Practice!

1. A solution of I_3^-(aq) can be standardized by using it to titrate As_4O_6(aq) under acidic conditions. The titration of 0.1021 g of As_4O_6(aq) dissolved in 30.00 mL of water requires 36.55 mL of I_3^-(aq) solution to reach equivalence point. The unbalanced equation is:

$$\mathrm{As_4O_6}(aq) + \mathrm{I_3^-}(aq) \rightarrow \mathrm{As_4O_{10}}(aq) + \mathrm{I^-}(aq)$$

   a. Balance the reaction above (in acidic solution). *Just kidding! This skill is no longer AP tested, so please feel free to fast forward through this part of the video. :D* **However, it would be great practice to take a moment and ID the oxidation state for each species in the reaction above: that's totally AP tested!**

$$\mathrm{As_4O_6}(aq) + 4\ \mathrm{H_2O}(l) + 4\ \mathrm{I_3^-}(aq) \rightarrow \mathrm{As_4O_{10}}(aq) + \mathrm{I^-}(aq)$$

   b. Identify the species oxidized and reduced in this reaction.

   c. Calculate the moles of As_4O_6(aq) present in the initial sample.

   d. Calculate the initial $[As_4O_6]$.

   e. How many moles of I_3^-(aq) were added to reach equivalence point?

   f. Calculate $[I_3^-]$.

Potential Titration Lab Errors:

1. Over-titration
   a. **Cause**: going past the equivalence point by adding too ________ titrant
   b. **Effect**: calculated moles of titrant ____ → calculated moles of analyte ____ → calculated [analyte] is ____ than actual [analyte]
2. Under-titration
   a. **Cause**: not reaching the equivalence point by adding too ___________ titrant
   b. **Effect**: calculated moles of titrant ____ → calculated moles of analyte ____ → calculated [analyte] is ____ than actual [analyte]
3. Water added to titrant (buret)
   a. **Cause**: Buret still wet from rinsing when it is filled with titrant
   b. **Effect**: Actual concentration of analyte is ____ than it is marked → ____ volume added than you should need → calculated moles of analyte are ____ → calculated [analyte] is ____
4. Water added to analyte (flask or beaker)
   a. **Cause**: Flask or beaker still wet from rinsing when the analyte (unknown solid) is added
   b. **Effect**: moles of analyte _________ change → ____ effect on [analyte]!

More practice!

2. The amount of $I_3^-(aq)$ in a solution can be determined by titration with a solution containing a known concentration of thiosulfate, $S_2O_3^{2-}(aq)$. The determination is based on the balanced equation:

$$I_3^-(aq) + 2\ S_2O_3^{2-}(aq) \rightarrow 3\ I^-(aq) + S_4O_6^{2-}(aq)$$

   a. Given that it requires 36.40 mL of 0.3300 M $Na_2S_2O_3(aq)$ to titrate the $I_3^-(aq)$ in a 15.00 mL sample, calculate the molarity of $I_3^-(aq)$ in the solution.

   b. After performing the titration, a student determines that the molarity of $I_3^-(aq)$ that they calculated was larger than the actual molarity of $I_3^-(aq)$. Explain the most likely source of their error.

Redox Titration Lab

In this lab, a redox titration (analogous to an acid/base titration) will be performed to study concepts of redox chemistry. In this redox reaction, Fe^{2+} is converted into Fe^{3+} and MnO_4^- (dark purple) is converted into Mn^{2+} (clear). The change in the oxidation state of manganese causes a change in color. When the reaction is complete and excess MnO_4^- is added to the reaction mixture, the solution turns pink and the titration is done. The net ionic equation that describes this redox reaction is described below:

$$8\ \mathrm{H}^+(aq) + 5\ \mathrm{Fe}^{2+}(aq) + \mathrm{MnO}_4^-(aq) \rightarrow 5\ \mathrm{Fe}^{3+}(aq) + \mathrm{Mn}^{2+}(aq) + 4\ \mathrm{H}_2\mathrm{O}(l)$$

Note about the experimental method: an indicator will NOT be needed for this lab to determine the equivalence point: as described above, MnO_4^- (one of the reactants) has a dark purple color. When reduced by the presence of iron (II) ions in solution (before the equivalence point), it becomes a clear manganese cation. However, once ALL iron (II) has been used up (i.e. the equivalence point is reached), any permanganate added into the solution will remain in excess – the characteristic dark purple color of the permanganate ion will give a light pink hue to the solution, which will get darker and darker the further past the equivalence point you titrate. Thus, one of the reactants acts as an indicator, and no additional indicator is required.

Materials:

Ring Stand
Buret Clamp
Buret
Erlenmeyer Flask (125 mL)
Graduated cylinder (10 mL)

Chemicals:

0.0100 M potassium permanganate, $KMnO_4$

$FeSO_4$ solution of unknown molarity

1.00 M sulfuric acid, H_2SO_4

distilled water

Procedure:

1. Record the initial volume of the $KMnO_4$ in Data Table 2.
2. Measure out approximately 7 mL of $FeSO_4$ in a 10 mL graduated cylinder, and transfer the $FeSO_4$ solution to a 125 Erlenmeyer flask. **Record the exact volume measured in Data Table 1.** (Hint: shining a flashlight through the index card from the back makes it much easier to read the purple meniscus!)
3. Measure out approximately 12 mL of distilled water into the same 10 mL graduated cylinder and transfer the distilled water to the 125 Erlenmeyer flask. Swirl the flask to mix the water with the $FeSO_4$ solution.
4. Add the $KMnO_4$ from the buret approximately 1.00 mL at a time to the flask. **Swirl the flask after each addition: the purple color should disappear when the flask is swirled** (the reaction is taking place). You can make the additions fairly quickly in the initial part of the titration.
5. **After 7 mL has been added, slow down the titration (make additions of 0.25 - 0.50 mL)**. When it takes a few seconds to get rid of the purplish color slow down even more (drop by drop). When the reaction is complete, the solution turns pink and stays pink for at least 30 sec. **Record the final volume of the $KMnO_4$ from the buret in Data Table 2.**
6. Dump the contents of the waste beaker and the 125 Erlenmeyer flask into the specified waste beaker. Clean the glassware used.

Data Table 1: Volume of $FeSO_4$

	Volume Measured (mL)	Volume Used (L)
Trial 1		
Trial 2		

Data Table 2: $KMnO_4$ Buret Readings

	Volume Measured		**Volume Used**	
	Initial (mL)	Final (mL)	mL	L
Trial 1				
Trial 2				

Analysis

1. Identify the oxidation state for each element in the reaction we performed in lab (shown below).

$$8\ \mathrm{H}^+(aq) + 5\ \mathrm{Fe}^{2+}(aq) + \mathrm{MnO}_4^-(aq) \rightarrow 5\ \mathrm{Fe}^{3+}(aq) + \mathrm{Mn}^{2+}(aq) + 4\ \mathrm{H}_2\mathrm{O}(l)$$

2. What was oxidized and what was reduced in this reaction? Justify your answer using the oxidation numbers derived in question #1.

3. How can you tell that all of the iron (II) has reacted? Describe in terms of excess and limiting reactants. (Hint: this is based on something you can see!)

4. What's the difference between the end point and an equivalence point in a titration?

Calculations Table: Fill in **AFTER** completing the calculations questions below

	Moles of Unknown	Volume Used (L)	Experimental $[Fe^{2+}]$	Accepted $[Fe^{2+}]$	% Error
Trial 1					
Trial 2					

Calculations

5. Calculate the moles of $KMnO_4$ added to reach the equivalence point in both trials.

6. Using your answer to #5, calculate the moles of $FeSO_4$ present in your initial solution in both trials. (Hint: yum, stoich!)

7. Using your answer to #6 and the volume of $FeSO_4$ used (from Data Table 1), calculate the initial concentration of your iron (II) solution in both trials.

8. Using the two values from #8, calculate the average value of $[Fe^{2+}]$. This is your "experimental" value.
   - Experimental value of $[Fe^{2+}]$: ______________________
   - Accepted value of $[Fe^{2+}]$: (from teacher) ______________________

   Using the accepted value provided, calculate the percent error in your calculated concentration.

$$\%\ \boldsymbol{Error} = \frac{\boldsymbol{Accepted - Experimental}}{\boldsymbol{Accepted}} \times \mathbf{100}$$

9. What do you think might be the source of the percent error calculated in question #8? How might this error have caused the percent error calculated?

Electrochemical Cells and Cell Potential (E°)

Electrochemistry: a study of the interchange of electrical and chemical energy

- ❖ There are two types of electrochemical cells: galvanic (voltaic) and electrolytic.
  - Voltaic or Galvanic Cell (_________________): a thermodynamically ________________ (i.e. spontaneous) redox reaction which generates useful electrical energy in the form of an electric current
  - Electrolytic Cell: requires electrical energy (direct current or DC power source) to drive a thermodynamically ____________________ (i.e. non-spontaneous) redox reaction.
- In short: galvanic (voltaic) cells ________________ current, while electrolytic cells _______ current!

QUICK REMINDERS

1) Oxidation is LOSS of electrons	2) Reduction is GAIN of electrons
OIL (LEO)	RIG (GER)
Ex: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	Ex: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Cell Potential (______): a measure of the potential difference (how much ____________ exists) between two half cells an electrochemical cell. The potential difference is caused by the ability of ______________ to flow from one half-cell to the other.

➔ The cell potential is a quantitative description of the _____________ __________ behind an electrochemical reaction that pushes electrons through the wire (or external circuit).

Standard Cell Potential (_________): cell potential measured at standard conditions**: 1 atm, 1 M solution, and 25°C**. A 'naught' sign (°) is used to show standard conditions. Usually measured in Volts (1 V = 1 Joule/Colomb = 1 J/C)

The cell potential can be easily calculated by adding the oxidation and reduction potentials.

$$E^o_{cell} = E^o_{ox} + E^o_{red}$$

- E^o_{cell} is the standard cell potential.
- E^o_{ox} is the standard oxidation cell potential for the oxidation half-reaction.
- E^o_{red} is the standard reduction cell potential for the reduction half-reaction.

➔ The more ________ the value of $E°_{cell}$, the greater the driving force of electrons through the system (under standard conditions), thus the more likely the reaction will proceed → more spontaneous/more thermodynamically favorable.

> **When you balance a redox reaction, don't change the cell potential!**
> Cell potential is an intensive property and thus does not depend on how many times a reaction occurs.

To find a table of Standard Electrode Potentials: Tro (p. 873, Section 18.4) or Zumdahl, 5th ed. (p. 843, Section 17.2)

1. Because the values come from a chart of standard reduction potentials, you ***MUST REVERSE*** the sign of the E° of the ***oxidized*** species before adding to the E° of the reduced species.
2. For a spontaneous redox reaction to occur, the overall cell potential must be ___________________.
   a) The metal with the **greater** (more positive) reduction potential will be **reduced**!
3. For a non-spontaneous redox reaction to occur, the overall cell potential must be ___________________.
   a) The metal with the **greater** (more positive) reduction potential will be **oxidized**!
4. A reduction potential table can be used as an activity series: metals with a lower reduction potential are more active and will replace metals with more positive potentials.

Example: Consider the half reactions shown below and the standard electrode reduction potentials that follow.

$Al^{3+}(aq) + 3\ e^- \rightarrow Al(s)$ $\qquad E^o = -1.66$ V

$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s)$ $\qquad E^o = -0.76$ V

1. Write the balanced redox reaction for zinc and aluminum that is thermodynamically favorable (i.e. spontaneous). Calculate the standard cell potential of this reaction.

2. Write the balanced redox reaction for zinc and aluminum that is *not* thermodynamically favorable (i.e. non-spontaneous). Calculate the standard cell potential of this reaction.

Now you try!

3. Identify the standard reduction potential, E^o_{red}, for each half reaction (use reduction potential chart!). Determine which species will be oxidized and which will be reduced for a redox reaction that is thermodynamically favorable (spontaneous). Next, calculate the value of E^o_{cell} for the thermodynamically favorable cell. Justify why your calculated E^o_{cell} represents a thermodynamically favorable reaction.

$Ag^+(aq) + 1e^- \rightarrow Ag(s)$ $\qquad E^o =$ ________

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\qquad E^o =$ ________

4. For the net ionic equation below, determine the standard cell potential, E°, for the reaction. Next, use the calculated value of E° to determine if the reaction is thermodynamically favorable (spontaneous) as written. (Hint: the reduction potential chart is all REDUCTION values. Flipping the rxn changes the sign of E).

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

For the two examples below, use the provided information to find the reduction potential of the missing half-reaction (**without** using the reduction potential chart 😉). Next, use the value of $E°_{cell}$ for the overall reaction to determine whether or not the reaction is thermodynamically favorable (spontaneous) as written.

1.

$$Al^{3+}(aq) + 3\ e^- \rightarrow Al(s) \qquad E^\circ_{red} = -1.66\ \text{V}$$

$$Pb^{2+}(aq) + 2\ e^- \rightarrow Pb(s) \qquad E^\circ_{red} = ???$$

$$2\ Al^{3+}(aq) + 3\ Pb(s) \rightarrow 2\ Al(s) + 3\ Pb^{2+}(aq) \qquad E^\circ_{cell} = -1.53\ \text{V}$$

2.

$$Cl_2(g) + 2\ e^- \rightarrow 2\ Cl^-(aq) \qquad E^\circ_{red} = +1.36\ \text{V}$$

$$Ni^{2+}(aq) + 2\ e^- \rightarrow Ni(s) \qquad E^\circ_{red} = ???$$

$$Ni(s) + Cl_2(g) \rightarrow 2\ Cl^-(aq) + 3\ Ni^{2+}(aq) \qquad E^\circ_{cell} = +1.59\ \text{V}$$

Multiple Choice Practice: Whee!

$$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s) \qquad E^\circ_{red} = +0.34\ V$$

$$Fe^{2+}(aq) + 2\ e^- \rightarrow Fe(s) \qquad E^\circ_{red} = -0.44\ V$$

3. Based on the reduction potentials given above, what is the reaction potential for the following reaction?

$$Fe^{2+}(aq) + Cu(s) \rightarrow Fe(s) + Cu^{2+}(aq)$$

a. – 0.78 V b. – 0.10 V c. + 0.10 V d. + 0.78 V

$$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s) \qquad E^\circ_{red} = +0.34\ V$$

$$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s) \qquad E^\circ_{red} = -0.76\ V$$

$$Mn^{2+}(aq) + 2\ e^- \rightarrow Mn(s) \qquad E^\circ_{red} = -1.18\ V$$

4. Based on the reduction potentials given above, which of the following reactions will be thermodynamically favored?

a. $Mn^{2+}(aq) + Cu(s) \rightarrow Mn(s) + Cu^{2+}(aq)$

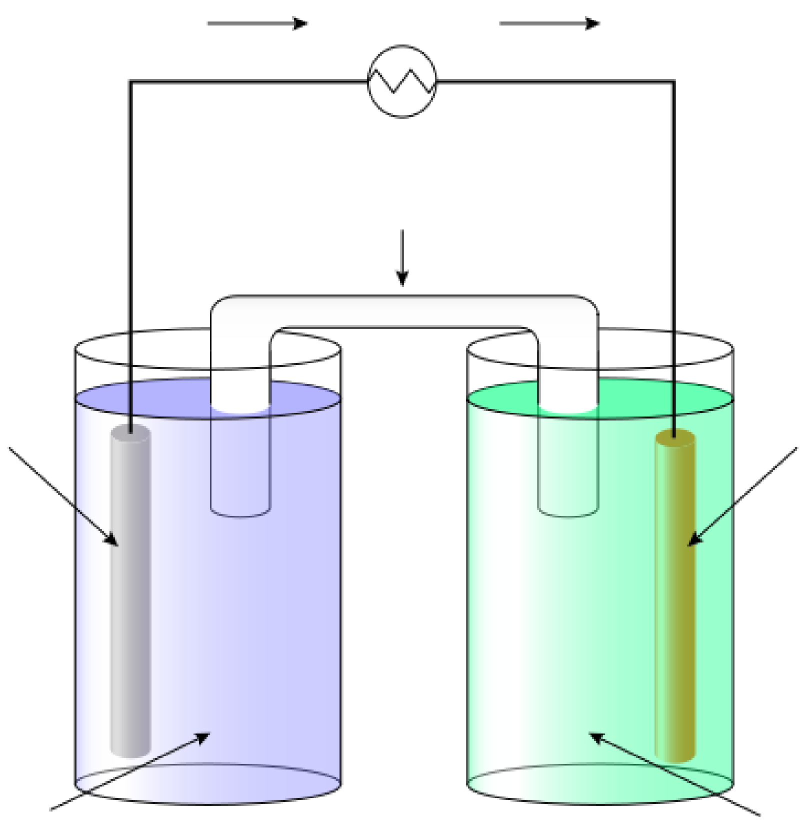
b. $Mn^{2+}(aq) + Zn(s) \rightarrow Mn(s) + Zn^{2+}(aq)$

c. $Zn^{2+}(aq) + Cu(s) \rightarrow Zn(s) + Cu^{2+}(aq)$

d. $Zn^{2+}(aq) + Mn(s) \rightarrow Zn(s) + Mn^{2+}(aq)$

Galvanic/ Voltaic Cells

- Redox reactions involve the transfer of electrons from one substance to another, and thus have the potential to generate an electric current (i.e. flow of ________________).
- To use that current, we need to _________________ the place where oxidation is occurring from the place where reduction is occurring.
  - o Current is the number of electrons that flow through the system per second.
  - o Current is measured in amperes, or _______ (A) = 1 coulomb of charge per second.
- This known as a **voltaic** (or **galvanic**) cell: the most common form of which is a _____________!
- Galvanic (voltaic) cells are always thermodynamically favorable (spontaneous) and thus have a ____ E^{o}_{cell}.



Parts of the Galvanic Cell

1. ______________ (–): the electrode where **oxidation** occurs (loses mass as reaction progresses)
2. ______________ (+): the electrode where **reduction** occurs (gains mass as reaction progresses, metal 'plated')
3. ______________ (or disk): bridge between cells whose purpose is to provide ______ to balance the charge and complete the circuit.
   - **An**ions (–) flow to the ___________; **cat**ions (+) flow to the ______________
   - If the salt bridge is removed, current will slow and then ________ (V = 0) as charge builds up in the half-cells.
4. ______________: measures the cell potential (emf or E°) in volts
   - Over time, voltage in the galvanic cell will decrease as [reactants] ____ and [products] ____

Important Notes about Galvanic Cells

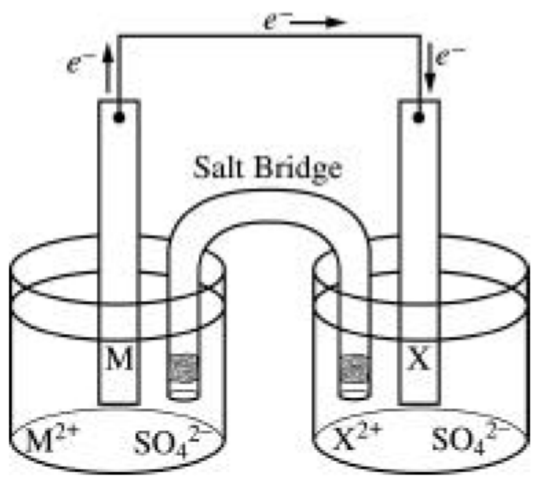
1. **Electron Flow**: ALWAYS through the _________ from anode to cathode (alpha order)
2. **Ion-ion or Ion-Gas Redox**: a voltaic cell can be constructed where the underlying redox reaction involves a gas or the conversion from one ion to another (unlikely a traditional voltaic cell in which reduction is from ion to solid and oxidation is from solid to ion).
   - ❖ Requires an _____________ electrode: this doesn't take part in the redox reaction but provides a surface on which the electrons can transfer; commonly made of platinum (expensive) or graphite (cheap)

Let's Practice!

1. The diagram below shows the experimental setup for a typical electrochemical cell that contains two standard half-cells. The cell operates according to the reaction represented by the following equation. (2002 #7, form B)

$$Zn(s) + Ni^{2+}(aq) \rightarrow Zn^{2+}(aq) + Ni(s)$$

   a. Identify M and M^{2+} in the diagram and specify the initial concentration for M^{2+} in solution.

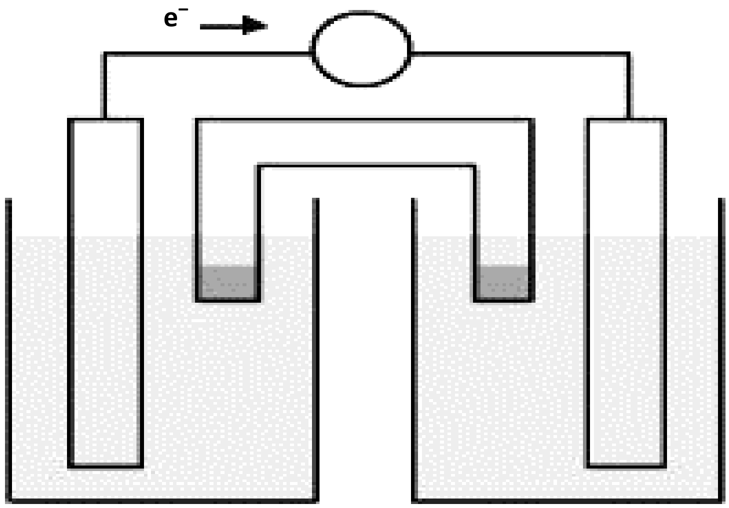


   b. Indicate which of the metal electrodes is the cathode. Write the balanced equation for the reaction that occurs in the half-cell containing the cathode.

   c. Describe what would happen to the cell voltage if the salt bridge was removed. Explain.

$$2\ Ce^{4+}(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + 2\ Ce^{3+}(aq)$$

2. The diagram below shows a galvanic cell based on the reaction represented above.



a. Label the anode and cathode in the diagram.

b. Label each electrode with the appropriate metal.

c. What ions will be found in solution in each half-cell? Fill them in on the diagram.

d. The diagram includes a salt bridge that is filled with a saturated solution of NaBr. Draw what happens in the salt bridge as the cell operates on the diagram above.

e. Circle the electrode that will lose mass as the reaction progresses. As the mass is "lost", where does it go?

f. Put a rectangle around the electrode that will gain mass as the reaction progresses. Where does this mass come from?

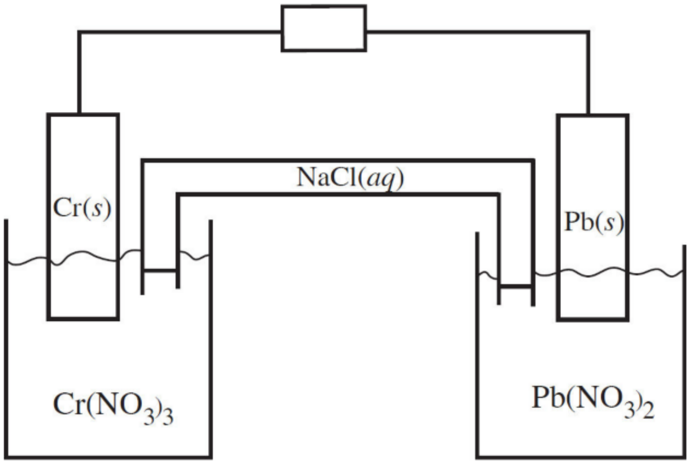
g. Use the following reduction potentials to calculate the voltage measured on the voltmeter shown above.

$$Sn^{2+}(aq) + 2\ e^- \rightarrow Sn(s) \qquad E^\circ_{red} = -0.14\ V$$

$$Ce^{4+}(aq) + 4\ e^- \rightarrow Ce(s) \qquad E^\circ_{red} = -2.34\ V$$

$$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq) \qquad E^\circ_{red} = +1.72\ V$$

A voltaic cell is created using the half-cells shown below. The concentrations of the solutions in each half-cell are 1.0 M.



$Cr^{3+} + 3e^- \rightarrow Cr(s)$ $E° = -0.41$ V

$Pb^{2+} + 2e^- \rightarrow Pb(s)$ $E° = -0.12$ V

3. Which of the following occurs at the cathode?
   a. Cr^{3+} is reduced to Cr(s).
   b. Pb^{2+} is reduced to Pb(s).
   c. Cr(s) is oxidized to Cr^{3+}.
   d. Pb(s) is oxidized to Pb^{2+}.

4. Which of the following best describes the activity in the salt bridge as the reaction progresses?
   a. Electrons flow through the salt bridge from the Pb/Pb^{2+} half-cell to the Cr/Cr^{3+} half-cell.
   b. Pb^{2+} flows to the Cr/Cr^{3+} half-cell, and Cr^{3+} flows to the Pb/Pb^{2+} half-cell.
   c. Na^+ flows to the Cr/Cr^{3+} half-cell, and Cl^- flows to the Pb/Pb^{2+} half-cell.
   d. Na^+ flows to the Pb/Pb^{2+} half-cell, and Cl^- flows to the Cr/Cr^{3+} half-cell.

5. Based on the given reduction potentials, which of the following would lead to a reaction?
   a. Placing some Cr(s) in a solution containing Pb^{2+} ions.
   b. Placing some Pb(s) in a solution containing Cr^{3+} ions.
   c. Placing some Cr(s) in a solution containing Cr^{3+} ions.
   d. Placing some Pb(s) in a solution containing Pb^{2+} ions.

6. Which of the following statements applies to the change in mass of the electrodes involved in this electrochemical cell?
   a. Cr(s) is the cathode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.
   b. Pb(s) is the cathode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.
   c. Cr(s) is the anode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.
   d. Pb(s) is the anode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.

Electrochem Equilibrium

Equilibrium: Recall that many chemical reactions are reversible. We show this with a double-ended arrow: $\rightleftharpoons$

➔ Surprise! The exchange of _____________ in a redox reaction is reversible, and therefore relevant to equilibrium.

<u>At Equilibrium</u>:

- Forward and backward reactions continue at the __________ rate.
- There is NO NET movement between reactants and products, which means there is _________ measurable voltage in an electrochemical cell at equilibrium!
- **Voltaic/Galvanic cell at equilibrium = ______________ battery!**

<u>The Equilibrium Constant, K</u>: ratio of products to reactants at equilibrium (at given temperature).

If you have the following equilibrium reaction, $aA + bB \rightleftharpoons cC + dD$

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b} = \frac{[Products]^{coefficient}}{[Reactants]^{coefficient}}$$

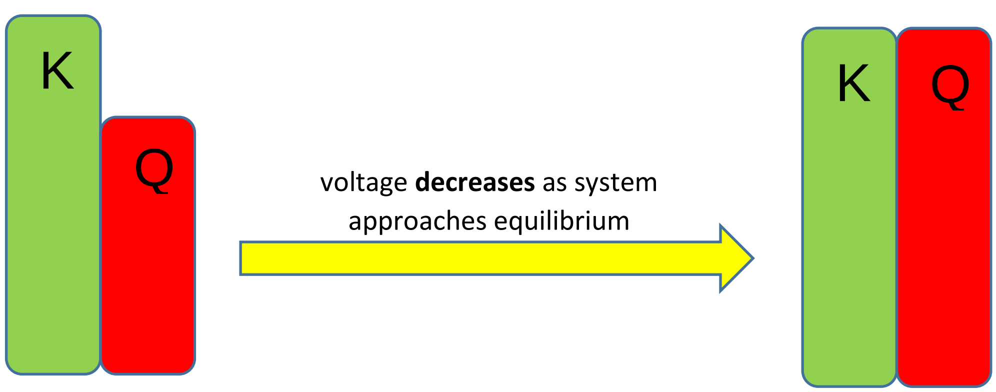
Since pure solids and liquids are not included in the K expression, only __________________ (or rarely, gaseous) reactants and products are included in the equilibrium constant expression, K, for electrochemical cells.

<u>The Reaction Quotient, Q</u>

In a standard galvanic/voltaic cell:

- K >> 1 and Q = 1 (since [reactants] = [products] = _________)
- Thus, K > Q and the reaction spontaneously makes products
    - o Since this is a redox reaction, as products are made, electrons are transferred = _____________ flows!
- Once the reaction has run long enough that K = Q, equilibrium is reached and no current flows (i.e. dead battery)

Voltaic Cell, Standard Conditions (1.0 M) **Cell at Equilibrium**



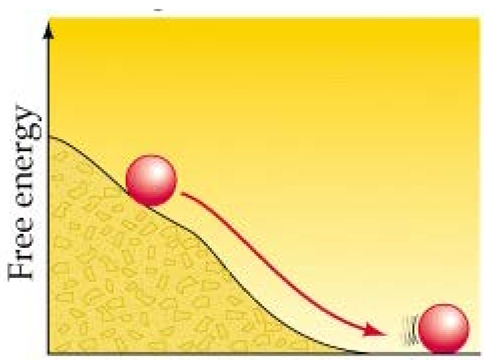
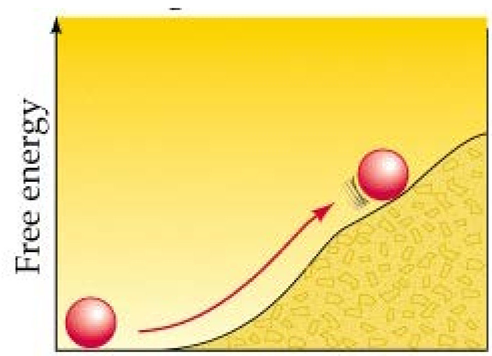
K >> Q
Products Form/Current Flows
V = E^o_{cell}

K = Q
No Products Form = No Current
Dead Battery (0 V)

Gibb's Free Energy (ΔG)

ΔG = change in free energy = ability to do ___________

→ Next unit, we will talk more about the meaning of "free energy", but for this unit you need to be able to use the numerical value of ΔG to predict the spontaneity of a reaction.

Exergonic Reaction (−ΔG)	Endergonic Reaction (+ΔG)
Spontaneous (Thermodynamically Favorable)	Not Spontaneous (Thermodynamically Unfavorable)
$+E^o{}_{cell}$ = voltage created (battery)	$-E^o{}_{cell}$ = external power source needed
	

The relationship between Gibb's free energy and cell potential can be quantified through the following equation:

$$\Delta G^o = -nFE^o_{cell}$$

The variables are: n = number of moles of electrons transferred in a **BALANCED** redox reaction

F = faraday's constant = 96,485 C/ mol e⁻ (charge on one mole of electrons)

Remember: E° is measured in volts, and 1 V = 1 J/C

Note about Units:

ΔG is *usually* measured in kJ/mol or J/mol: Mastering Chem only uses units of kJ or J!

Let's Practice!

1. Identify n (the number of moles of electrons transferred) in each of the following reactions (hint: think about the half-reactions).

   a. $2\ Ce(s) + 3\ Ni^{2+}(aq) \rightleftharpoons 2\ Ce^{3+}(aq) + 3\ Ni(s)$ n = ______

   b. $Sn(s) + 2\ Ag^{+}(aq) \rightleftharpoons Sn^{2+}(aq) + 2\ Ag(s)$ n = ______

2. Calculate ΔG° (in kJ/mol) for this reaction: $Fe(s) + Cu^{2+}(aq) \rightleftharpoons Fe^{2+}(aq) + Cu(s)$, where $E^o{}_{cell} = 0.79$ V. Is this reaction thermodynamically favorable? Explain.

ΔG and K are related by the following equation at standard conditions (1 M, 1 atm, 298 K)

$$\Delta G^o = -RT \ln K$$

The variables are: R = universal gas constant = 8.314 J mol^{-1} K^{-1}
T = temperature (in Kelvin)
K = equilibrium constant

- This equation can be rewritten to give: $K = \mathrm{e}^{-\Delta G^o/RT}$
- The units for $\Delta G^o = \dfrac{\text{joules}}{\text{moles}_{\text{reaction}}} = \dfrac{\text{J}}{\text{mol}_{\text{rxn}}}$ ***but Mastering Chem uses J***

Summary

E^o_{cell}	ΔG^o	K	**Thermodynamically favorable?**
+	–	K > 1	Favorable
–	+	K < 1	Not favorable
= 0	= 0	K = Q	n/a (at equilibrium)

Let's Practice!

1. The standard cell potential, E^o_{cell}, is +0.67 V for the balanced oxidation-reduction reaction shown below:

$$S_4O_6^{2-}(aq) + 2\ Cr^{2+}(aq) \rightarrow 2\ Cr^{3+}(aq) + 2\ S_2O_3^{2-}\ (aq)$$

a. Calculate the free energy change for the cell (in kJ/mol_{rxn}).

b. Calculate the equilibrium constant for this reaction (at 25°C).

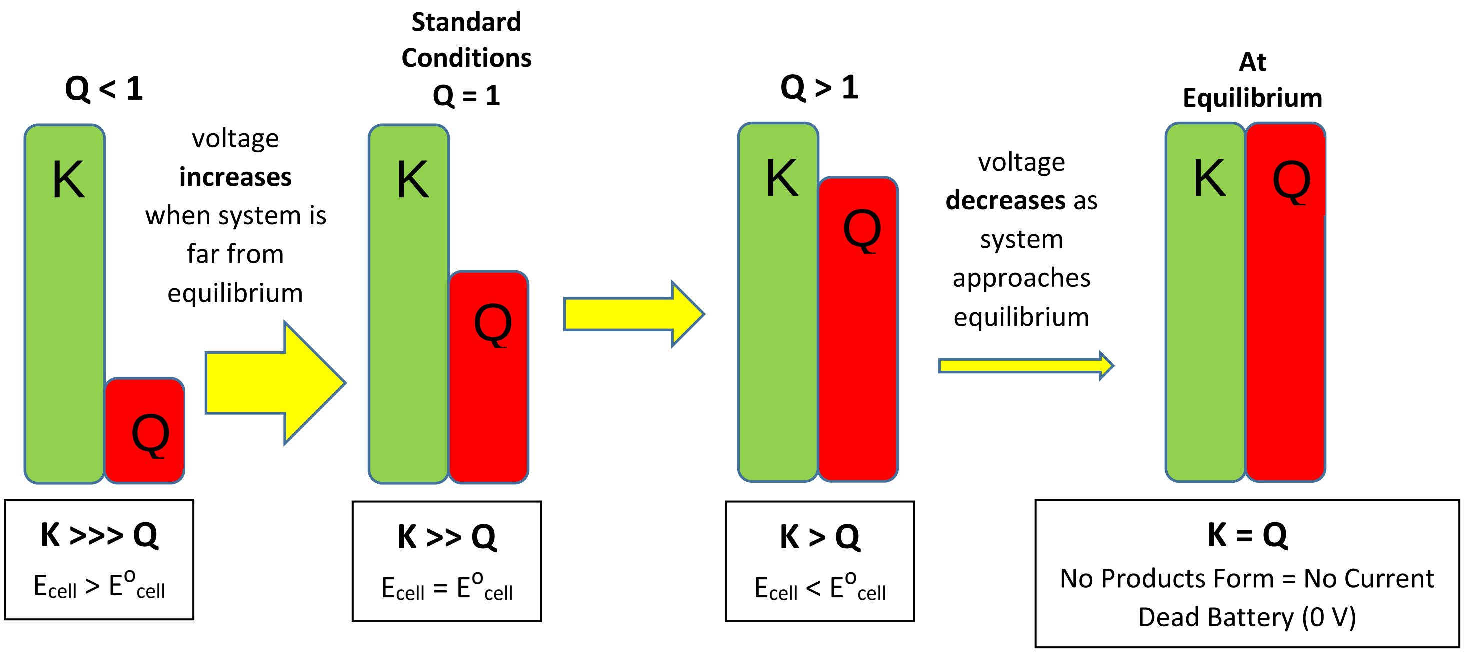
Electrochem Equilibrium: Non-standard Conditions

Standard Cell Potential (E^o_{cell}): cell potential measured at standard conditions of **1 atm, 1.0 M solution, and 25°C**.

- As a voltaic cell discharges, reactants are consumed and products are generated. Therefore, as the cell operates, the conditions become non-standard and the voltage _______________!

Non-Standard Cell Potential (E_{cell}): cell potential measured at nonstandard conditions (concentration ≠ 1.0 M)

- In a standard galvanic/voltaic cell: K >> 1 and Q = 1, voltage measured = E^o_{cell}
  - If concentrations change to make **Q larger** (______________ to K), **voltage decreases**! $E_{cell} < E^o_{cell}$
  - If concentrations change to make **Q smaller** (______________ from K), **voltage increases**! $E_{cell} > E^o_{cell}$



Conditions that Increase Q, Decrease Voltage ($E_{cell} < E^o_{cell}$)

- Increase [Products] *relative* to [Reactants]

Conditions that Decrease Q, Increase Voltage ($E_{cell} > E^o_{cell}$)

- Increase __________________
- Decrease __________________

<u>Note about thermodynamic favorability</u>:

- If $E_{cell} > E^o_{cell}$, cell voltage increases, and the cell is ___________ thermodynamically favorable.
- If $E_{cell} < E^o_{cell}$, cell voltage decreases, and the cell is ___________ thermodynamically favorable.

Let's Practice!

1. Given the information below, answer the following questions.

$$Ni(s) + Cu^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cu(s), \quad E^{o}_{cell} = +0.59\ V$$

   a. A galvanic cell is constructed where $[Cu^{2+}] = 0.50$ M and $[Ni^{2+}] = 1.0$ M. Is the voltage of this cell higher, lower, or equal to the standard cell potential? Why?

   b. A galvanic cell is constructed with $[Cu^{2+}] = 2.0$ M, $[Ni^{2+}] = 2.0$ M. Is the voltage of this cell higher, lower, or equal to the standard cell potential? What happens to the amount of time this cell could operate? Explain.

   c. A standard galvanic cell is constructed (using the balanced equation provided), and the voltage is measured to be 0.59 V. After two hours, the voltage is re-measured: would you expect the new cell potential to be higher, lower, or equal to the standard cell potential? Justify your answer.

In Summary

- Increasing [reactants] or decreasing [products] will cause voltage to <u>**increase**</u>. ($E_{cell} > E^{o}_{cell}$)
- Decreasing [reactants] or increasing [products]will cause voltage to <u>**decrease**</u>. ($E_{cell} < E^{o}_{cell}$)
- Increasing or decreasing <u>**both**</u> [reactants] and [products] <u>**by the same factor**</u> will cause voltage to <u>stay the same</u> **if and only if** they have the same stoichiometric coefficient in the balanced equation.
- Removing <u>**ALL**</u> of one or more reactants will cause voltage to <u>drop to zero</u>.
- Removing the solid product (cathode electrode) has <u>**NO**</u> effect on voltage.
- Increasing <u>reactants</u> will increase amount of time the cell can operate (regardless of change to cell potential).

The Nernst Equation

How to Calculate Cell Potential Under Non-standard Conditions

The voltage of the cell is dependent on the concentration of aqueous reactants and products.

$$E_{cell} = E^{\circ}_{cell} - \frac{RT}{nF} \ln Q$$

E_{cell} = cell potential under nonstandard conditions (V)
E°_{cell} = cell potential under standard conditions (V)
R = gas constant ($8.314 \text{ J mol}^{-1}\text{K}^{-1}$)
T = absolute temperature (K)
n = number of moles of electrons transferred in the reaction (mol)
F = Faraday's constant ($96{,}485 \text{ C mol}^{-1}$)
Q = reaction quotient

The Nernst equation is often used when conditions are at ____________ temperature (25°C or 298K) and can be simplified to the following equation, where 0.052 is a ____________________ that incorporates Faraday's constant, F, the universal gas constant, R, and a temperature of 298K.

$$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n} \ln Q$$

	$E_{cell} > E^{\circ}_{cell}$	$E_{cell} = E^{\circ}_{cell}$	$E_{cell} < E^{\circ}_{cell}$
Q	less than 1	equal to 1	greater than 1
$\ln Q$	− (negative)	0 (zero)	+ (positive)

<u>Note about voltage</u>:

- If $E_{cell} > E^{o}_{cell}$, cell voltage ___________________, and the cell is __________ thermodynamically favorable.
- If $E_{cell} < E^{o}_{cell}$, cell voltage ___________________, and the cell is __________ thermodynamically favorable.

Example: Given the information below, a non-standard cell is made using the reaction shown below, with $[Cu^{+}]$ = 0.50 M and $[Al^{3+}]$ = 1.0 M.

$$3\ Cu^{+}(aq) + Al(s) \rightarrow 3\ Cu(s) + Al^{3+}(aq),\ E^{o}_{cell} = 2.18\ V$$

a. Predict: will the voltage of the cell increase or decrease when compared to the standard cell? Why?

b. Calculate the value of E_{cell} under these conditions. Do your results match your prediction?

How to Answer FR Questions about Nonstandard Cell Conditions

You do not need to use the Nernst equation on the AP Chem test, but you can if you want to. ;D
You will only be asked to make qualitative predictions about E_{cell} or voltage (does it increase, decrease, or remain the same relative to standard?). ***However***, you WILL need to explain how the nonstandard concentrations affect the value of the reaction quotient, *Q*, and what that means for voltage or the cell potential.

- It can be very helpful, on FR, to write out how you calculate the value of the reaction quotient, *Q*, and compare that value to when *Q* = 1 (standard conditions).

Example FR Explanations

1. [Products] > [reactants]: Q > 1, voltage ↓
   - Because the concentration of products is greater than the concentration of reactants, Q > 1, and therefore the cell voltage decreases relative to standard conditions ($E_{cell} < E^{o}_{cell}$).
2. [Reactants] > [products]: Q < 1, voltage ↑
   - Because the concentration of products is less than the concentration of reactants, Q < 1, and therefore the cell voltage increases relative to standard conditions ($E_{cell} > E^{o}_{cell}$).
3. [Reactants] **and** [products] multiplied by same amount: Q = 1, voltage stays the same!
   - Because the ratio of products to reactants is the same as standard conditions, Q =1, and therefore the cell voltage is the same as the voltage at standard conditions ($E_{cell} = E^{o}_{cell}$).

Let's Practice!

1. A nonstandard galvanic cell is constructed based on the following half-reactions at 25°C, where $[Cd^{2+}]$ = 0.010 M and $[Pb^{2+}]$ = 0.100 M.

$$\mathrm{Cd^{2+}}(aq) + 2\mathrm{e^-} \rightarrow \mathrm{Cd}(s) \qquad \mathrm{E^o} = -0.40\ \mathrm{V}$$
$$\mathrm{Pb^{2+}}(aq) + 2\mathrm{e^-} \rightarrow \mathrm{Pb}(s) \qquad \mathrm{E^o} = -0.13\ \mathrm{V}$$

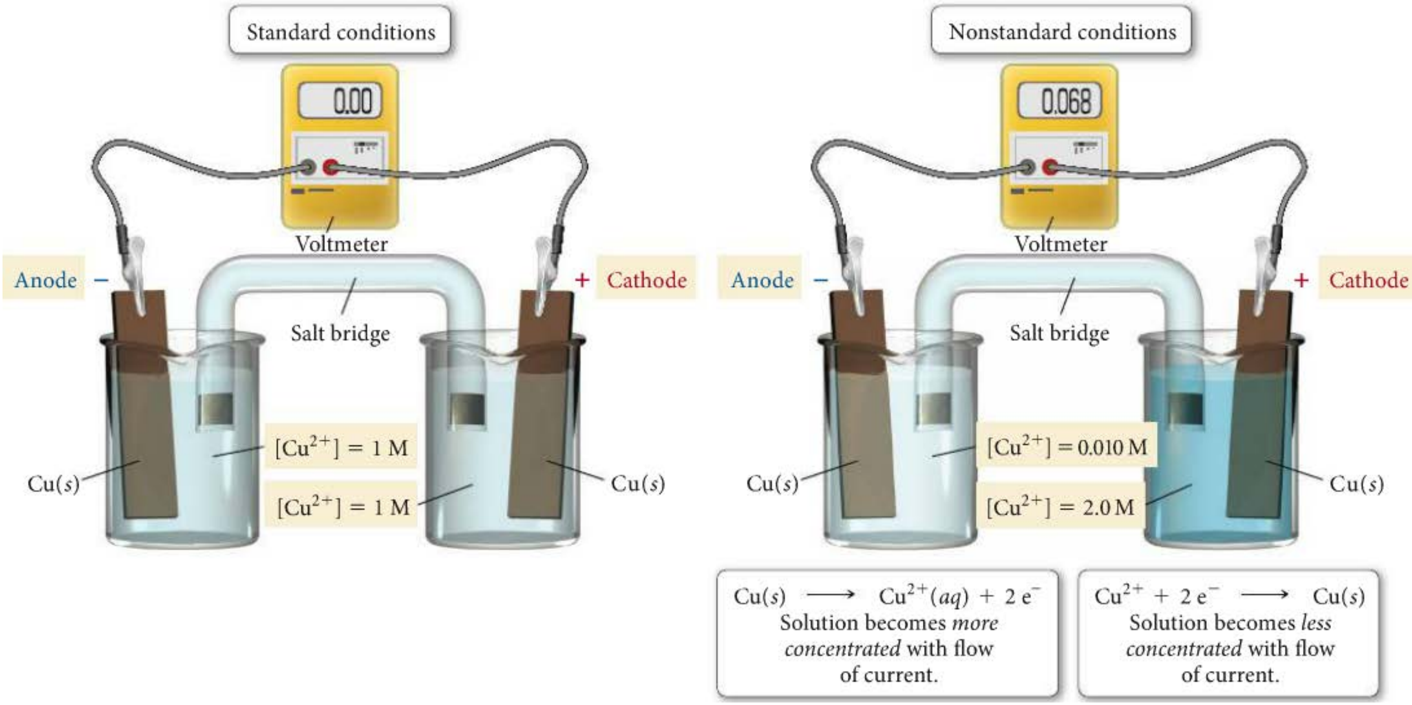
   a. Identify and write the equation for the half-reaction which occurs at the anode and the half-reaction that occurs at the cathode.

   b. Write the balanced reaction occurring in this galvanic cell.

   c. Under these conditions would the cell potential be greater than, less than, or equal to the cell potential under standard conditions? Justify your answer.

Concentration Cells

You can make a concentration cell by using two half cells with the same metal/ion combination, as long as the ion concentrations differ! **Electrons will flow from the lower concentration half-cell to the higher concentration half-cell to achieve equilibrium.**

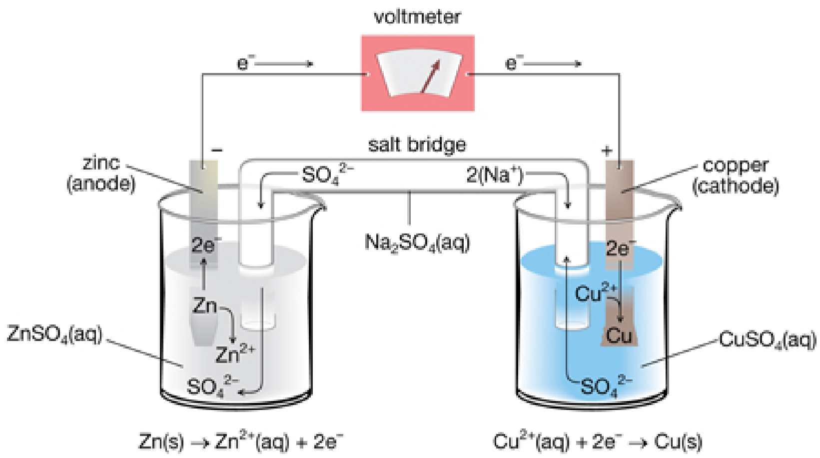


Let's Practice!

2. A cell has on its left side a 0.20 M Zn^{2+} solution. The right side has a 0.050 M Zn^{2+} solution. The compartments are connected by Zn electrodes and a salt bridge. Draw this galvanic cell and designate the cathode, anode, and direction of current.

Let's Practice!

3. A voltaic cell is constructed based on the following reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$



a. A student accidentally adds additional $ZnSO_4$ to the $Zn(s)/Zn^{2+}(aq)$ half-cell. What happens to the magnitude of the cell voltage (relative to the standard cell)? Justify your answer.

b. Is the value of the equilibrium constant for the cell reaction greater than 1, less than 1, or equal to 1? Explain.

c. What must be true about the standard free energy change of this reaction, ΔG°? Justify.

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

4. A galvanic cell based on the reaction represented above was constructed from zinc and copper half-cells. The observed voltage was found to be 1.22 volt instead of the standard cell potential, E°, of 1.10 volts. Which of the following could correctly account for this observation?

A. The cell had been running for a period of time.
B. The standard free energy of the cell, ΔG°, is negative.
C. The Cu^{2+} solution was less concentrated than the Zn^{2+} solution.
D. The Zn^{2+} solution was less concentrated than the Cu^{2+} solution.

5. Which of the following statements is true about the reaction below?

$$2\,Ag(s) + Cu^{2+}(aq) \rightleftharpoons 2\,Ag^{+}(aq) + Cu(s) \qquad K_{eq} = 2.7 \times 10^{-16}$$

a. E° and ΔG° are both positive.
b. E° and ΔG° are both negative.
c. E° is positive and ΔG° is negative.
d. E° is negative and ΔG° is positive.

6. In the reaction below, a piece of solid nickel is added to a solution of potassium dichromate, $K_2Cr_2O_7$.

$$14\ H^+(aq) + Cr_2O_7^{2-}(aq) + 3\ Ni(s) \rightarrow 2\ Cr^{3+}(aq) + 3\ Ni^{2+}(aq) + 7\ H_2O(l)$$

How many moles of electrons are transferred when 1 mole of potassium dichromate is mixed with 3 mol of nickel?

a. 2 moles of electrons
b. 3 moles of electrons
c. 5 moles of electrons
d. 6 moles of electrons

7. Calculate the standard free energy of the following reaction at 25°C.

$$Fe^{2+}(aq) + Mg(s) \rightarrow Fe(s) + Mg^{2+}(aq) \qquad E^\circ_{cell} = 1.92\ V$$

a. 3.7×10^5 J/mol$_{rxn}$
b. 1.6×10^3 J/mol$_{rxn}$
c. -3.7×10^5 J/mol$_{rxn}$
d. -1.6×10^3 J/mol$_{rxn}$

$$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$$

8. An electrolytic cell based on the reaction represented above was constructed from iron and copper half-cells. The observed voltage was found to be 0.59 volt instead of the standard cell potential, E°, of 0.78 volts. Which of the following could correctly account for this observation?

A. The copper electrode was larger than the iron electrode.
B. The solutions in the half-cells had different volumes.
C. The Cu^{2+} solution was more concentrated than the Fe^{2+} solution.
D. The Fe^{2+} solution was more concentrated than the Cu^{2+} solution.

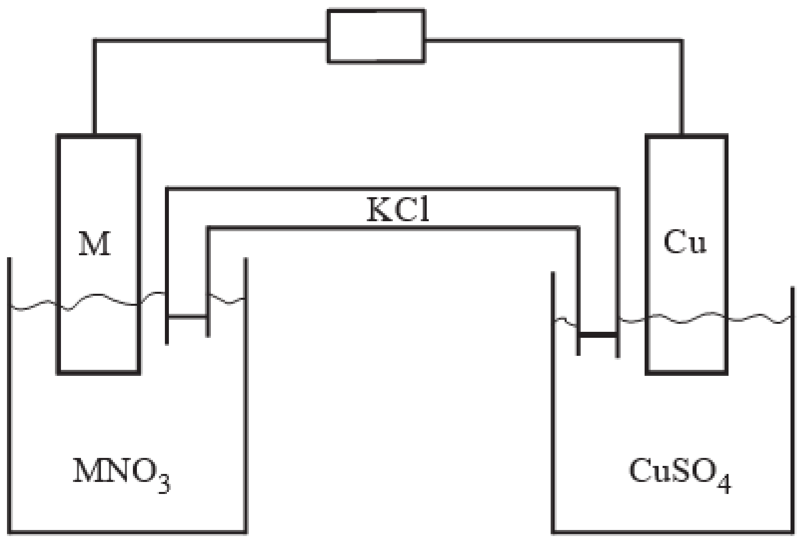
$$2\ Al(s) + 3\ Mn^{2+}(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ Mn(s)$$

9. A thermodynamically favorable cell, utilizing the reaction shown above, ran for 45 minutes. What happens to the measured voltage and why?

A. The measured voltage decreases over time because deviations in concentration that bring the cell closer to equilibrium will decrease the magnitude of the cell potential.
B. The measured voltage increases over time because deviations in concentration that bring the cell closer to equilibrium will increase the magnitude of the cell potential.
C. The measured voltage increases over time because $[Mn^{2+}]$ increases as the cell runs.
D. The measured voltage remains constant because E°_{cell} is an intensive property.

Unit 3: Free Response Practice (7 points)

1. A student performs an experiment in which a bar of unknown metal M is placed in a solution with the formula MNO_3. The metal is then hooked up to a copper bar in a solution of $CuSO_4$ as shown below. A salt bridge that contains aqueous KCl links the cell together.



The standard cell potential is found to be +0.74 V. Separately, when a bar of metal M is placed in the copper sulfate solution, solid copper starts to form on the bar. When a bar of copper is placed in the MNO_3 solution, no visible reaction occurs.

The following gives some reduction potentials for copper:

Half-reaction	$E°$
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34 V
$Cu^{2+} + e^- \rightarrow Cu^+$	0.15 V
$Cu^+ + e^- \rightarrow Cu(s)$	0.52 V

a. Write the net ionic equation that takes place in the Cu/M cell. [1 point]

b. What is the standard reduction potential for metal M? [2 points]

c. Which metal acted as the anode and which as the cathode? Justify your answer. [1 point]

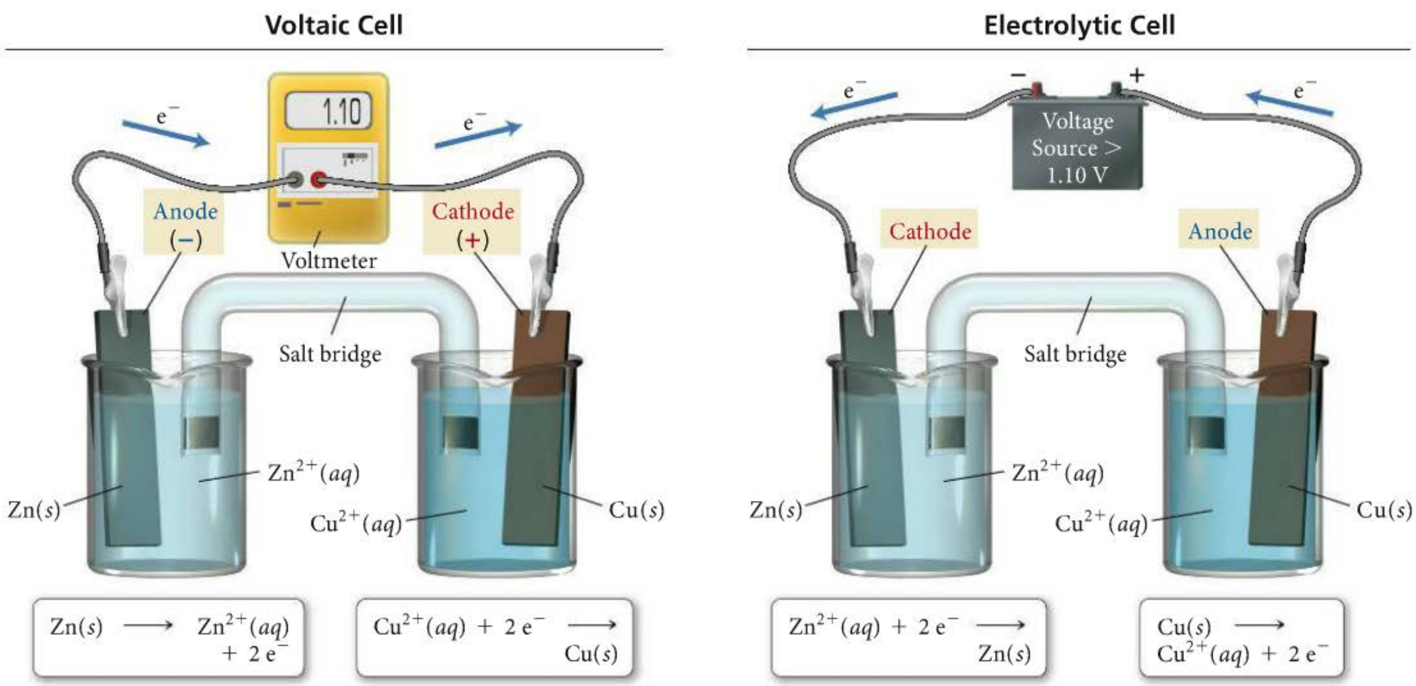
d. On the diagram of the cell, indicate which way the electrons are flowing in the wire. Additionally, indicate any ionic movement occurring in the salt bridge. [2 points]

e. What would happen to the voltage of the reaction in the Cu/M cell if the concentration of the $CuSO_4$ increased while the concentration of the MNO_3 remained constant? Justify your answer. [1 point]

Electrolytic Cells: You have the *power*!

<u>Electrolytic cells</u>: thermodynamically unfavorable, therefore ____ **ΔG^0** and ____ **ΔE^0**

- Since an electrolytic cell is NOT spontaneous, it will undergo a redox reaction only if _____________ is applied!



<u>Note</u>: In electrolytic cells, **An Ox** and **Red Cat** still work (yay!)

Differences between Galvanic/Voltaic Cells and Electrolytic Cells

Galvanic/ Voltaic Cells	**Electrolytic Cells**
___ΔG, ___E^o_{cell}, K ____ 1	___ΔG, ___E^o_{cell}, K ____ 1
Thermodynamically favorable	Thermodynamically Unfavorable
spontaneous in the _____________ direction	spontaneous in the _____________ direction
<u>Separated</u> into two half cells to generate electricity	Usually occurs in a __________ container (but can be set up in two containers)
____ a battery (turns chemical energy into electrical energy)	__________ a battery (turns electrical energy into chemical energy)
Often electrodes made of metal used in half-reactions	Usually use _____________ electrodes (such as Pt or graphite)
Electrons supplied by species being oxidized	Electrons supplied by external battery at cathode
Cathode ____, Anode ____	Anode ____, Cathode ____

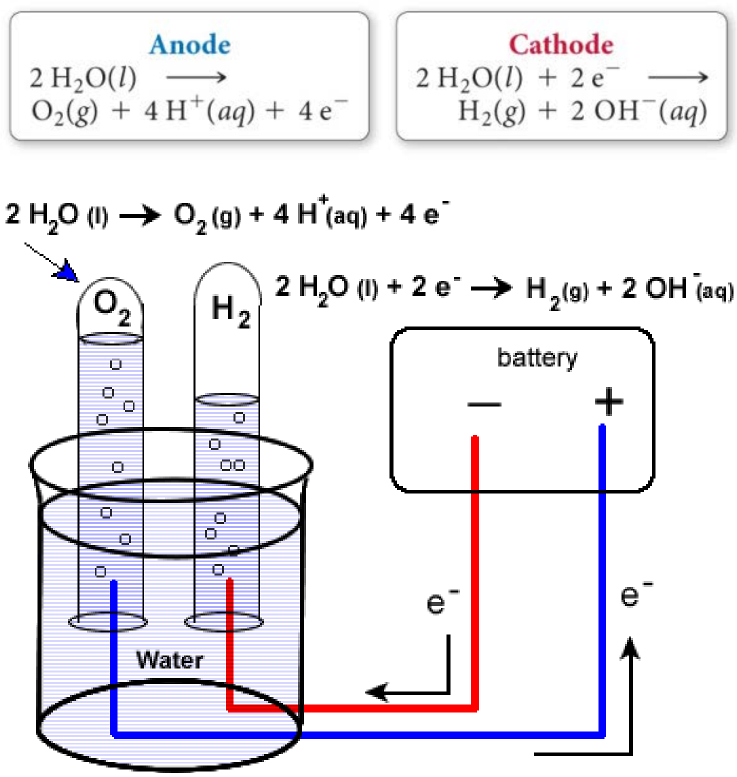
Electrolysis: the process of using electrical energy to ___________ apart a compound.

→ Can be used to separate compounds into their component ___________________.

Applications of Electrolysis

1. Electrolysis of water → used to generate hydrogen for _______ cells!
   - As water dissociates, ____ moles of hydrogen gas are formed for every ____ mole of oxygen gas, according to the balanced reaction shown below.
   - Note: water is being dissociated at the ___________ rate in both test tubes, but _______________ as much gas is being produced in the hydrogen gas test tube!

$$2\ \mathrm{H_2O}(l) \rightarrow 2\ \mathrm{H_2}(g) + \mathrm{O_2}(g)$$



Let's Practice!

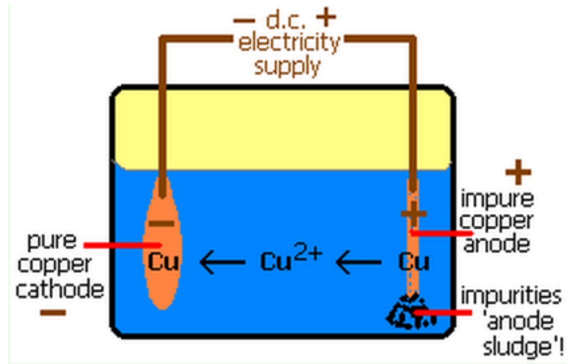
1. In the dissociation of water, what is being reduced? Justify your answer using oxidation states.

2. In the dissociation of water, what is being oxidized? Justify your answer using oxidation states.

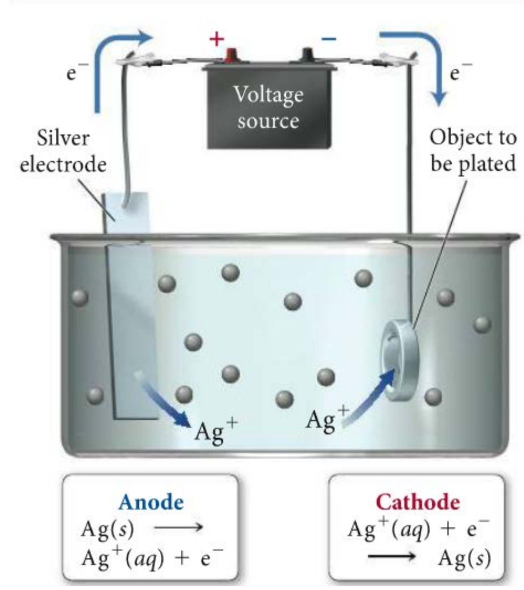
3. During a lab demonstrating the electrolysis of water, 2.7 mL of $O_2(g)$ was collected. What volume of $H_2(g)$ should be collected? Explain.

2. **<u>Electrorefining</u>:** Purification of metals (often _______) through electrolysis
    - The <u>anode</u> is the _____________ metal (i.e. ore) to be purified.
    - The <u>cathode</u> is the electrode at which the _________ metal will be deposited (often made of a thin sheet of the pure metal).
    - The electrolyte (solution) contains the ___________ of the metal to be purified.
    - This process can be used to purify multiple metals, including copper, nickel, cobalt, lead, and tin.

Electrolytic Cell for Copper Ore Purification



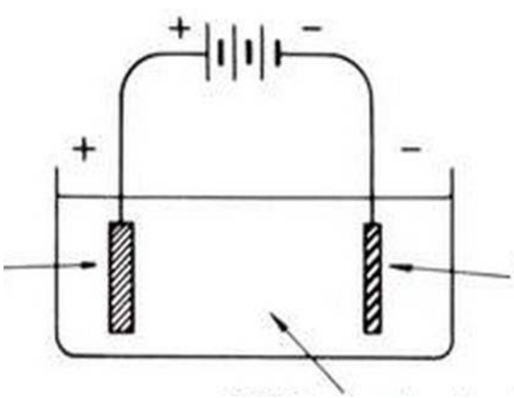
Electrolytic Cell for Silver Plating



3. **<u>Electroplating</u>:** Applying a _______ layer of an expensive metal onto a less expensive one for structural or cosmetic reasons
    - The <u>object to be plated</u> is the ________________.
        - o Remember: Fat Cat!
    - The electrolyte (solution) contains the cation of the metal to be plated on the object.
    - The best anode is made of the metal to be _____________ onto the object.

Example: Gold ore, when discovered in nature, often contains impurities. If a sample of gold ore contains some silver impurity, the ore can be purified by electrolysis. (Assume Au(s) will form a Au^{3+}(aq) cation.)

a. On the diagram below, identify the anode (and what it's made of), the cathode (and what it's made of), and the direction of electron flow.



b. Where will the purified gold be found? Add it into the diagram on the left.

c. What might be a possible electrolyte for use in this solution?

<u>Quantitative Electrolysis</u>: The Math (Or How I Learned to Stop Worrying and Love Dimensional Analysis)

In an electrolytic cell, the amount of product made is related to the number of electrons transferred. Essentially, the electrons are a reactant. To solve, use the formula for current from the periodic table and follow the following steps:

1. Convert your time, t, into seconds.

2. Multiply your time (sec) by the current (Amp) to solve for the charge (q) in Coulombs.

$$I = \frac{q}{t} \rightarrow q = t \cdot I$$

I = current (amperes, A)
q = charge (coulombs, C)
t = time (sec)

$$\text{Amp} = \frac{C}{sec}$$

3. Divide by Faraday's constant (given on the formula chart) to convert Coulombs into moles of electrons.

q / F = mol e⁻

Faraday's constant, F, is the charge on 1 mole of electrons.

F = 96, 485 C/mol e^-

Use F = 100,000 (1E5) for multiple choice!

*<u>Note</u>: a "faraday" can be used to refer to the number of moles of electrons being transferred:

- "3 faradays" → 3 moles of e^- of transferred
- "0.25 faradays" → 0.25 moles of e^- of transferred

4. Use the appropriate half-reaction to determine the moles of the metal formed.

For example,

$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$

This indicates for every 3 moles of electrons gained, 1 mole of solid gold is formed.

5. Use molar mass to convert from moles to grams of metal.

Example: Gold can be plated out of a solution containing Au^{3+} according to the half-reaction:

$$Au^{3+}(aq) + 3\ e^- \rightarrow Au(s)$$

What mass of gold (in grams) is plated by a 25 minute flow of 5.5 A current?

Solution: $$25\ \cancel{\text{min}} \times \frac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}} \times \frac{5.5\ \cancel{\text{C}}}{1\ \cancel{\text{s}}} \times \frac{1\ \cancel{\text{mol e}^-}}{96{,}485\ \cancel{\text{C}}} \times \frac{1\ \cancel{\text{mol Au}}}{3\ \cancel{\text{mol e}^-}} \times \frac{196.97\ \text{g Au}}{1\ \cancel{\text{mol Au}}} = 5.6\ \text{g Au}$$

You will need to be able to do two basic calculations for quantitative electrolysis:

1. Given time (sec) and current (A), calculate mass (g).
2. Given mass (g) and current (A), calculate time required (sec).

Of course, there are endless variaties of these two calculation types we can give you! 😉

Quick Trick to remember the order of steps to calculate mass of a metal produced (given time and current):

Are you **SAF**e? **MMM.**

Let's Try!

1. If 3.30 faraday of charge is passed through a solution of $Al_2(SO_4)_3$, what mass of aluminum is deposited?

2. How long must a current of 5.00 A be applied to a solution of AgCN to produce 10.5 g silver metal?

3. Copper may be used for electroplating, starting with a solution of $Cu(NO_3)_2(aq)$.
   a. If a current of 10.0 amp is applied to the $Cu(NO_3)_2(aq)$ solution for 60.0 minutes, what mass of copper will be plated out? (Assume excess $Cu(NO_3)_2(aq)$).

   b. How many moles of electrons must be transferred in this reaction to produce 5.16 g of copper metal?

4. Molten $AlCl_3$ is electrolyzed with a constant current of 5.00 amperes over a period of 600.0 seconds. Which of the following expressions is equal to the maximum mass of Al(s) that plates out? (1 faraday = 96,500 coulombs)

   a. $\frac{(600)(5.00)}{(96{,}500)(3)(27.0)}$ grams

   b. $\frac{(600)(5.00)(3)(27.0)}{(96{,}500)}$ grams

   c. $\frac{(600)(5.00)(27.0)}{(96{,}500)(3)}$ grams

   d. $\frac{(96{,}500)(3)(27.0)}{(600)(5.00)}$ grams

5. A chemist wants to plate out 1.00 g of solid iron from a solution containing aqueous Fe^{2+} ions. Which of the following expressions will equal the amount of time, in seconds, it takes if a current of 4.00 A is applied?

   a. $\frac{(2)(55.85)(4.00)}{(96{,}500)}$ seconds

   b. $\frac{(2)(96{,}500)}{(55.85)(4.00)}$ seconds

   c. $\frac{(55.85)(96{,}500)}{(2)(4.00)}$ seconds

   d. $\frac{(2)(55.85)(96{,}500)}{(4.00)}$ seconds

6. If 0.060 faraday is passed through an electrolytic cell containing a solution of In^{3+} ions, the maximum number of moles of In that could be deposited at the cathode is

   (A) 0.010 mole
   (B) 0.020 mole
   (C) 0.030 mole
   (D) 0.060 mole
   (E) 0.18 mole

7. If a copper sample containing some zinc impurity is to be purified by electrolysis, the anode and the cathode must be which of the following?

	Anode	Cathode
(A)	Pure copper	Pure zinc
(B)	Pure zinc	Pure copper
(C)	Pure copper	Impure copper sample
(D)	Impure copper sample	Pure copper
(E)	Impure copper sample	Pure zinc

Unit 3 All Electrochem 2-page Review

OIL RIG – oxidation is loss, reduction is gain (of electrons)

Rules for Assigning Oxidation Numbers (in order of priority)

If any rules are in conflict, follow the rule that is **higher** on the list!

Oxidation Rules:

1. Free elements = 0
2. All atoms in a neutral compound add up to 0.
3. All atoms in a polyatomic ion add up to the ion's charge.
4. *The rules below apply to bonded elements*:
   a. Group 1A metals = +1
   b. Group 2A = +2
   c. Non-metals usually follow the chart to the right, in order:
   d. Note: the carbon family (4A) isn't mentioned – you will ALWAYS have to solve for the oxidation number of group 4A elements in a compound.

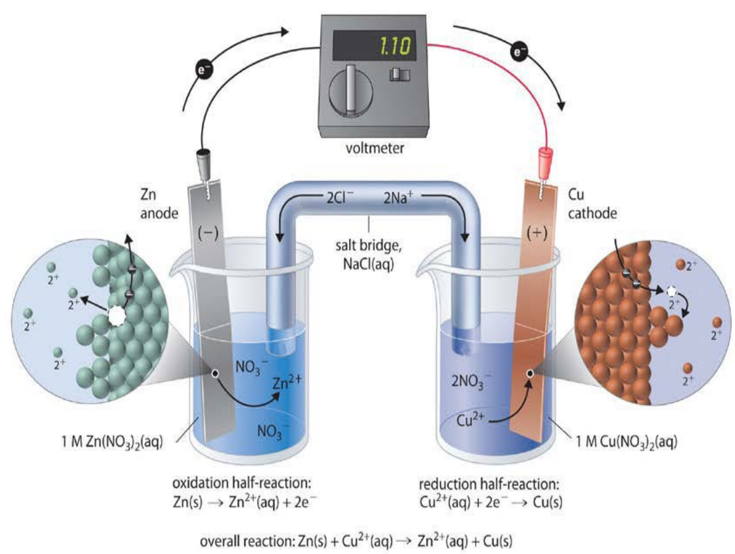
Nonmetal	Oxidation State	Example
Fluorine	−1	MgF_2 −1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	−2	CO_2 −2 ox state
Group 7A	−1	CCl_4 −1 ox state
Group 6A	−2	H_2S −2 ox state
Group 5A	−3	NH_3 −3 ox state

Electrochemistry Involves TWO MAIN TYPES Of Electrochemical Cells:

1. Galvanic (voltaic) cells – thermodynamically favorable → **battery** (+E_{cell} , −ΔG, K > 1)
2. Electrolytic cells –thermodynamically unfavorable and require external power source (−E_{cell} , +ΔG, K < 1)

GALVANIC or VOLTAIC CELL "ANATOMY"

- Anode – (An Ox) the electrode where oxidation occurs, loses mass into solution as the cell runs ("anode"-rexic)
- Cathode – (Red Cat) the electrode where reduction occurs, gains mass from solution as the cell runs (fat cat)
- Inert electrodes – used when a gas is involved OR ion to ion; Pt (expensive) or graphite (cheap)
- Salt bridge – used to maintain electrical neutrality in a galvanic cell: **anions to anode, cations to cathode**
- Electron flow – ALWAYS through the wire from anode to cathode (alphabetical order) **FAT CAT



(a)

CELL POTENTIAL, E_{cell}

- it is a measure of the electromotive force or the "pull" of the electrons as they travel from the anode to the cathode, measured in Volts (V)
- E_{cell} becomes E^o_{cell} when measurements are taken at standard conditions (1 atm, 1.0 *M,* and 25^{o}C)

Standard Electrode Potentials (will be given as needed for each problem!)

- Galvanic/voltaic cells: **The MORE POSITIVE reduction potential is reduced**

Calculating Standard Cell Potential Symbolized by E^o_{cell}

1. The Metal with the MORE POSITIVE REDUCTION POTENITAL is be REDUCED, so the other is oxidized.
2. Reverse the equation that will be oxidized and change the sign of its voltage: this is now E^o_{ox}.

$$E^o_{cell} = E^o_{ox} + E^o_{red}$$

DEPENDENCE OF CELL POTENTIAL ON CONCENTRATION

- Voltaic cells at NONstandard conditions: Le Chatlier's principle can be applied. An increase in the concentration of a reactant will favor the forward reaction and the cell potential will increase. The converse is also true!
  - ↑ [reactants] or ↓ [products]: $E_{cell} > E^o_{cell}$, ↑ voltage
  - ↓ [reactants] or ↑ [products]: $E_{cell} < E^o_{cell}$, ↓ voltage
- For free response, write out how you calculate the value of the reaction quotient, *Q*, and compare that value to when *Q* = 1 (standard conditions).
  - **Q > 1**: As the concentration of the products of a redox reaction increases, the value of the reaction quotient, *Q*, increases and therefore the voltage decreases as the reaction shifts to the left decreasing the driving force.
  - **Q < 1**: As the concentration of the reactants of a redox reaction increases, the value of the reaction quotient, *Q*, decreases and therefore the voltage increases as the reaction shifts to the right increasing the driving force.

ELECTROLYTIC CELLS

- Thermodynamically unfavorable, therefore + ΔG^0 and - ΔE^0
- Since an electrolytic cell is NOT spontaneous, it will undergo a redox reaction only if an outside power source is applied!
- AN OX and RED CAT still work
- Used to separate ores or plate out metals.

In an electrolytic cell, the amount of product made is related to the number of electrons transferred. Essentially, the electrons are a reactant.

You will need to be able to do two basic calculations for quantitative electrolysis:

1. Given time (sec) and current (A), calculate mass (g).
2. Given mass (g) and current (A), calculate time required (sec).

Quick Trick to remember the order of steps to calculate mass of a metal produced (given time and current):

Are you **SAF**e? **MMM.**

1. Convert your time, t, into seconds.

⇩

2. Multiply your time (sec) by the current (Amp) to solve for the charge (q) in Coulombs.

$$I = \frac{q}{t} \rightarrow q = t \cdot I$$

I = current (amperes, A)
q = charge (coulombs, C)
t = time (sec)

$\text{Amp} = \frac{C}{sec}$

⇩

3. Divide by Faraday's constant (given on the formula chart) to convert Coulombs into moles of electrons.

q / F = mol e⁻

Faraday's constant, F, is the charge on 1 mole of electrons.

F = 96, 485 C/mol e^-

Use F = 100,000 (1E5) for multiple choice!

*Note: a "faraday" can be used to refer to the number of moles of electrons being transferred:

- "3 faradays" → 3 moles of e^- of transferred
- "0.25 faradays" → 0.25 moles of e^- of transferred

⇩

4. Use the appropriate half-reaction to determine the moles of the metal formed.

For example,

$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$

This indicates for every 3 moles of electrons gained, 1 mole of solid gold is formed.

⇩

5. Use molar mass to convert from moles to grams of metal.

Unit 3 Multiple Choice Practice

1. According to the balanced equation above, how many moles of the permanganate ion are required to react completely with 25.0 ml of 0.100 M hydrogen peroxide?

$$6\ H^+ + 5\ H_2O_2 + 2\ MnO_4^- \rightarrow 5\ O_2 + 2\ Mn^{2+} + 8\ H_2O$$

   a. 0.000500 mol
   b. 0.00100 mol
   c. 0.00500 mol
   d. 0.00625 mol

2. A chemist wants to plate out 29 g of solid nickel from a solution containing aqueous $Ni(NO_3)_2$. Approximately how many moles of electrons must be transferred to produce that mass of solid nickel?

   a. 0.25 mol e^-
   c. 0.50 mol e^-
   b. 1.0 mol e^-
   d. 1.5 mol e^-

3. A balanced equation for the reaction of copper metal with nitric acid is shown below. Which of the following represents a true statement about the reaction?

$$3\ \mathrm{Cu}(s) + 8\ \mathrm{H}^+(aq) + 2\ \mathrm{NO}_3^-(aq) \rightarrow 3\ \mathrm{Cu}^{2+}(aq) + 4\ \mathrm{H_2O}(l) + 2\ \mathrm{NO}(g)$$

   a. The oxidation state of nitrogen changed from +5 to +2.
   b. Hydrogen ions are oxidized to form H_2O(l).
   c. The oxidation state of oxygen changes from −1 to −2.
   d. Copper metal is reduced to a copper (II) ion.

4. Molten $GaCl_3$ is electrolyzed with a constant current of 1.30 amperes over a period of 2.00 minutes. Which of the following expressions is equal to the maximum mass of Ga(s) that plates out? (1 faraday = 96,500 coulombs)

   a. $\dfrac{(120)(1.30)}{(96{,}500)(3)(69.7)}$ grams
   b. $\dfrac{(120)(1.30)(3)(69.7)}{(96{,}500)}$ grams
   c. $\dfrac{(120)(1.30)(69.7)}{(96{,}500)(3)}$ grams
   d. $\dfrac{(96{,}500)(3)(69.7)}{(120)(1.30)}$ grams

$$Cr^{3+}(aq) + 3\,e^- \rightarrow Cr(s) \qquad E^\circ_{red} = -0.74\text{ V}$$

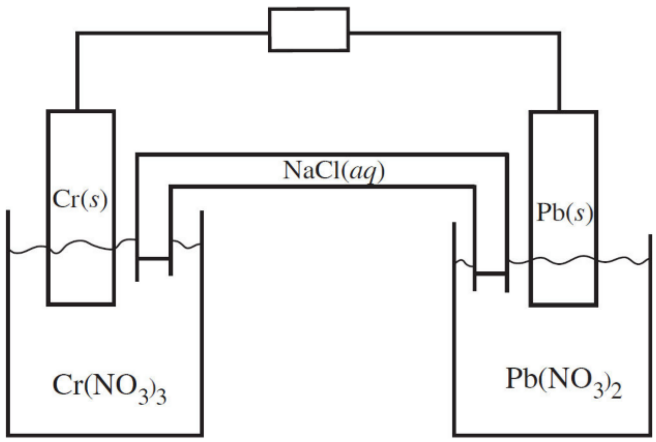
$$Fe^{2+}(aq) + 2\,e^- \rightarrow Fe(s) \qquad E^\circ_{red} = -0.44\text{ V}$$

5. Based on the reduction potentials given above, what is the reaction potential for the following reaction?

$$3\,Fe(s) + 2\,Cr^{3+}(aq) \rightarrow 3\,Fe^{2+}(aq) + 2\,Cr(s)$$

a. – 0.16 V b. – 0.30 V c. + 0.16 V d. + 0.30 V

Use the diagram below to answer the questions 6 - 8. A voltaic cell is created using the half-cells shown below. The concentrations of the solutions in each half-cell are 1.0 M.



$Cr^{3+} + 3e^- \rightarrow Cr(s)$ $E^\circ = -0.41$ V
$Pb^{2+} + 2e^- \rightarrow Pb(s)$ $E^\circ = -0.12$ V

6. Which of the following occurs at the cathode?
   a. Cr^{3+} is reduced to Cr(s).
   b. Pb^{2+} is reduced to Pb(s).
   c. Cr(s) is oxidized to Cr^{3+}.
   d. Pb(s) is oxidized to Pb^{2+}.

7. Which of the following best describes the activity in the salt bridge as the reaction progresses?
   a. Electrons flow through the salt bridge from the Pb/Pb^{2+} half-cell to the Cr/Cr^{3+} half-cell.
   b. Pb^{2+} flows to the Cr/Cr^{3+} half-cell, and Cr^{3+} flows to the Pb/Pb^{2+} half-cell.
   c. Na^+ flows to the Cr/Cr^{3+} half-cell, and Cl^- flows to the Pb/Pb^{2+} half-cell.
   d. Na^+ flows to the Pb/Pb^{2+} half-cell, and Cl^- flows to the Cr/Cr^{3+} half-cell.

8. Which of the following statements applies to the change in mass of the electrodes involved in this electrochemical cell?
   a. Cr(s) is the cathode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.
   b. Pb(s) is the cathode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.
   c. Cr(s) is the anode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.
   d. Pb(s) is the anode and it gains mass since metal ions are being converted to metal atoms which often adhere to the electrode.

$$Cu^{2+}(aq) + 2\ e^- \rightarrow Cu(s) \qquad E^\circ_{red} = +0.34\ V$$

$$Zn^{2+}(aq) + 2\ e^- \rightarrow Zn(s) \qquad E^\circ_{red} = -0.76\ V$$

$$Mn^{2+}(aq) + 2\ e^- \rightarrow Mn(s) \qquad E^\circ_{red} = -1.18\ V$$

9. Based on the reduction potentials given above, which of the following reactions will be thermodynamically favored?

   a. $Mn^{2+}(aq) + Cu(s) \rightarrow Mn(s) + Cu^{2+}(aq)$

   b. $Mn^{2+}(aq) + Zn(s) \rightarrow Mn(s) + Zn^{2+}(aq)$

   c. $Zn^{2+}(aq) + Cu(s) \rightarrow Zn(s) + Cu^{2+}(aq)$

   d. $Zn^{2+}(aq) + Mn(s) \rightarrow Zn(s) + Mn^{2+}(aq)$

10. A chemist wants to plate out 98 g of solid titanium from a solution containing Ti_2S_3. Approximately how many moles of electrons must be transferred to produce that much solid titanium?

   a. 2.0 mol e^- c. 4.0 mol e^-

   b. 3.0 mol e^- d. 6.0 mol e^-

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

11. A galvanic cell based on the reaction represented above was constructed from zinc and copper half-cells. The observed voltage was found to be 0.22 volt instead of the standard cell potential, E°, of 1.25 volts. Which of the following could correctly account for this observation?

   A. The cell had been running for a period of time.
   B. The standard free energy of the cell, ΔG°, is negative.
   C. The Cu^{2+} solution was less concentrated than the Zn^{2+} solution.
   D. The Zn^{2+} solution was less concentrated than the Cu^{2+} solution.

$$2\ H_2O(l) + 4\ MnO_4^-(aq) + 3\ ClO_2^-(aq) \rightarrow 4\ MnO_2(aq) + 3\ ClO_4^-(aq) + 4\ OH^-(aq)$$

12. Which species is reduced in the reaction represented above?

   a. MnO_2 b. ClO_2^- c. MnO_4^- d. ClO_4^-

13. In the reaction below, a piece of solid nickel is added to a solution of potassium dichromate.

$$14\ H^+(aq) + Cr_2O_7^{2-}(aq) + 3\ Ni(s) \rightarrow 2\ Cr^{3+}(aq) + 3\ Ni^{2+}(aq) + 7\ H_2O(l)$$

How many moles of electrons are transferred when 1 mole of potassium dichromate is mixed with 3 mol of nickel?

   a. 2 moles of electrons c. 5 moles of electrons
   b. 3 moles of electrons d. 6 moles of electrons

14. Calculate the standard free energy of the following reaction at 25°C.

$$Fe^{2+}(aq) + Mg(s) \rightarrow Fe(s) + Mg^{2+}(aq) \qquad E^{o}_{cell} = 1.92\ V$$

a. 3.7×10^{5} J/mol$_{rxn}$
b. 1.6×10^{3} J/mol$_{rxn}$
c. -3.7×10^{5} J/mol$_{rxn}$
d. -1.6×10^{3} J/mol$_{rxn}$

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

15. A thermodynamically favorable cell, utilizing the reaction shown above, ran for 45 minutes. What happens to the measured voltage and why?

A. The measured voltage decreases over time because deviations in concentration that bring the cell closer to equilibrium will decrease the magnitude of the cell potential.
B. The measured voltage increases over time because deviations in concentration that bring the cell closer to equilibrium will increase the magnitude of the cell potential.
C. The measured voltage increases over time because $[Mn^{2+}]$ increases as the cell runs.
D. The measured voltage remains constant because E^{o}_{cell} is an intensive property.

16. Given the following half reactions:

$$Sn^{4+} + 2e^{-} \rightarrow Sn^{2+} \qquad E^{o} = 0.15\ V$$
$$Fe^{3+} + 1e^{-} \rightarrow Fe^{2+} \qquad E^{o} = 0.77\ V$$

Determine the standard cell potential ($E°_{cell}$) for the voltaic cell based on the reaction

$$Sn^{2+} + 2\ Fe^{3+} \rightarrow 2\ Fe^{2+} + Sn^{4+}$$

a. +0.62 V
b. +0.92 V
c. −0.62 V
d. −0.92 V

17. A chemist wants to plate out 1.00 g of solid iron from a solution containing aqueous $Fe(NO_3)_3$. Which of the following expressions will equal the amount of time, in seconds, it takes if a current of 2.00 A is applied?

a. $\frac{(3)(55.85)(2.00)}{(96{,}500)}$ seconds

b. $\frac{(3)(96{,}500)}{(55.85)(2.00)}$ seconds

c. $\frac{(55.85)(96{,}500)}{(3)(2.00)}$ seconds

d. $\frac{(3)(55.85)(96{,}500)}{(2.00)}$ seconds

18. For this reaction, $E°_{cell}$ = 0.79 V.

$$6\ I^- + Cr_2O_7^{2-} + 14\ H^+ \rightarrow 3\ I_2 + 2\ Cr^{3+} + 7\ H_2O$$

Given that the standard reduction potential for $Cr_2O_7^{2-} \rightarrow 2\ Cr^{3+}$ is 1.33 V, what is $E°_{red}$ for $I_{2(aq)}$?

a. +0.54 V

b. −0.54 V

c. +2.12 V

d. −2.12 V

19. If 0.060 faraday is passed through an electrolytic cell containing a solution of Ni^{2+} ions, the maximum number of moles of Ni that could be deposited at the cathode is

a. 0.020 mol
b. 0.030 mol
c. 0.060 mol
d. 0.12 mol

20. If a gold sample containing some silver impurity is to be purified by electrolysis, the anode and the cathode must be which of the following?

	Anode	**Cathode**
a.	Pure gold	Pure silver
b.	Pure silver	Pure gold
c.	Pure gold	Impure gold sample
d.	Impure gold sample	Pure gold
e.	Impure gold sample	Pure silver

21. Which expression below should be used to calculate the mass of copper that can be plated out of a 1.0 M $Cu(NO_3)_2$ solution using a current of 0.75 A for 5.0 min?

a. $\frac{(5.0)(60)(0.75)(63.55)}{(96{,}500)(2)}$

b. $\frac{(5.0)(60)(63.55)(2)}{(0.75)(96{,}500)}$

c. $\frac{(5.0)(60)(96{,}500)(0.75)}{(63.55)(2)}$

d. $\frac{(5.0)(60)(96{,}500)(63.55)}{(0.75)(2)}$

22. Which of the following statements is true about the reaction below?

$$2\ Ag^{+}(aq) + Cu(s) \rightleftharpoons 2\ Ag(s) + Cu^{2+}(aq) \qquad K_{eq} = 3.7 \times 10^{15}$$

a. E° and ΔG° are both positive.

b. E° and ΔG° are both negative.

c. E° is positive and ΔG° is negative.

d. E° is negative and ΔG° is positive.

23. In the reaction

$$SO_2 + 2\ H_2S \rightarrow 3\ S + 2\ H_2O$$

a. sulfur is oxidized and hydrogen is reduced
b. sulfur is reduced and there is no oxidation
c. sulfur is oxidized and sulfur is reduced
d. sulfur is reduced and hydrogen is oxidized

24. Which of these ions is most easily oxidized?

Standard Reduction Potentials, E°	
$Fe^{3+} + e^{-} \rightarrow Fe^{2+}$	+ 0.77 V
$Cu^{2+} + e^{-} \rightarrow Cu^{+}$	+ 0.15 V

a. Fe^{2+}

b. Fe^{3+}

c. Cu^{2+}

d. Cu^{+}

25. What is the oxidation number of manganese in the $KMnO_4$?

a. +1 b. +2 c. +5 d. +7

26. When this reaction is balanced, the coefficient on the Sn^{2+} is.

$$Sn^{4+} + Cr \rightarrow Cr^{3+} + Sn^{2+}$$

a. 1
b. 2
c. 3
d. 4

Unit 3: AP Free Response Practice #1 [2010B #3, 10 points]

A sample of ore containing the mineral tellurite, TeO_2, was dissolved in acid. The resulting solution was then reacted with a solution of $K_2Cr_2O_7$ to form telluric acid, H_2TeO_4. The unbalanced chemical equation for the reaction is given below.

$$3\,\mathrm{TeO_2}(s) + \mathrm{Cr_2O_7^{2-}}(aq) + 8\,\mathrm{H^+}(aq) \rightarrow 3\,\mathrm{H_2TeO_4}(aq) + 2\,\mathrm{Cr^{3+}}(aq) + \mathrm{H_2O}(l)$$

a. Identify the molecule or ion that is being oxidized in the reaction.

b. Give the oxidation number of Cr in the $Cr_2O_7^{2-}$(aq) ion.

In the procedure described above, 46.00 mL of 0.03109 M $K_2Cr_2O_7$ was added to the ore sample after it was dissolved in acid. When the chemical reaction had progressed as completely as possible, the amount of unreacted (excess) $Cr_2O_7^{2-}$(aq) was determined by titrating the solution with 0.110 M $Fe(NO_3)_2$. The reaction that occurred during the titration is represented by the following balanced equation.

$$6\,\mathrm{Fe^{2+}}(aq) + \mathrm{Cr_2O_7^{2-}}(aq) + 14\,\mathrm{H^+}(aq) \rightarrow 2\,\mathrm{Cr^{3+}}(aq) + 6\,\mathrm{Fe^{3+}}(aq) + 7\,\mathrm{H_2O}(l)$$

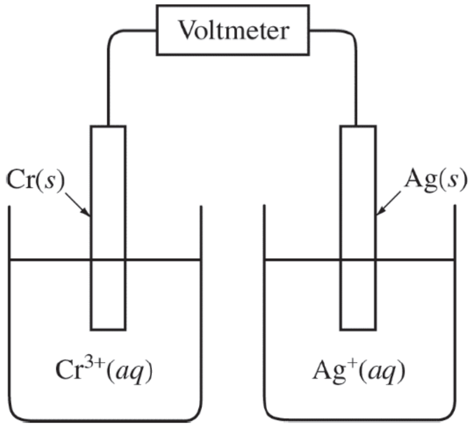
A volume of 9.85 mL of 0.110 M $Fe(NO_3)_2$ was required to reach the equivalence point.

c. Calculate the number of moles of excess $Cr_2O_7^{2-}$(aq) that was titrated.

d. Calculate the number of moles of $Cr_2O_7^{2-}$(aq) that reacted with the tellurite.

e. Calculate the mass, in grams, of tellurite that was in the ore sample.

AP Free Response Practice #2 (2018 #6, 4 points)



Half-Reaction	$E°$ (V)
$Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$	+ 0.80
$Cr^{3+}(aq) + 3\,e^{-} \rightarrow Cr(s)$	?

A student sets up a galvanic cell at 298 K that has an electrode of Ag(s) immersed in a 1.0 M solution of Ag^{+}(aq) and an electrode of Cr(s) immersed in a 1.0 M solution of Cr^{3+}(aq), as shown in the diagram above.

a. The student measures the voltage of the cell shown above and discovers that it is zero. Identify the missing component of the cell, and explain its importance for obtaining a nonzero voltage.

b. The student adds the missing component to the cell and measures $E°_{cell}$ to be +1.54 V. As the cell operates, Ag^{+} ions are reduced. Use this information and the information in the table above to do the following.

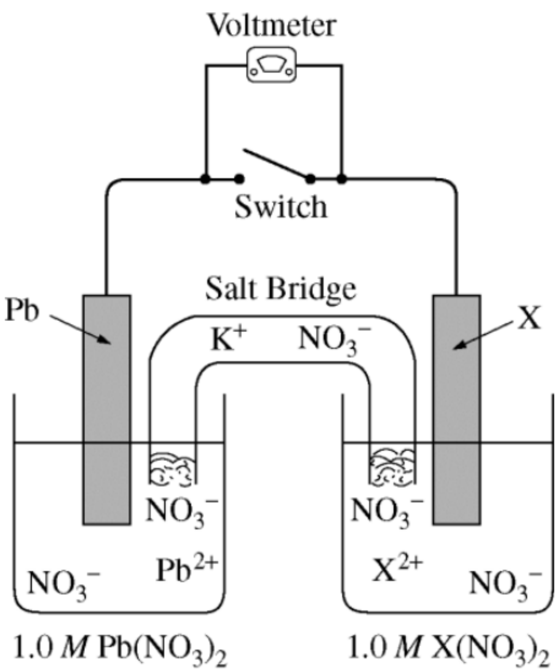
   i. Calculate the value of E° for the half-reaction $Cr^{3+} + 3\,e^{-} \rightarrow Cr(s)$.

   ii. Write the balanced net-ionic equation for the overall reaction that occurs as the cell operates.

   iii. Calculate the value of ΔG° for the overall cell reaction in J/mol_{rxn}.

Unit 3: AP Free Response Practice #3 (2012 #6, modified) [8 points]

The diagram below shows an electrochemical cell that is constructed with a Pb electrode immersed in 100. mL of 1.0 M $Pb(NO_3)_2$(aq) and an electrode made of metal X immersed in 100. mL of 1.0 M $X(NO_3)_2$(aq). A salt bride containing saturated aqueous KNO_3 connects the anode compartment to the cathode compartment. The electrodes are connected to an external circuit containing a switch, which is open. When a voltmeter is connected to the circuit as shown, the reading on the voltmeter is 0.47 V. When the switch is closed, electrons flow through the switch from the Pb electrode toward the X electrode.



Half-reaction	Standard Reduction Potential
$Cu^+ + e^- \rightarrow Cu(s)$	+ 0.52 V
$Cu^{2+} + 2\,e^- \rightarrow Cu(s)$	+ 0.34 V
$Pb^{2+} + 2\,e^- \rightarrow Pb(s)$	− 0.13 V
$Ni^{2+} + 2\,e^- \rightarrow Ni(s)$	− 0.25 V

a) Write the equation for the half-reaction that occurs at the anode.

b) The value of the standard potential for the cell, E°, is 0.47 V. Use the table of reduction potentials provided to answer the following.

i. Determine the standard reduction potential for the half reaction that occurs at the cathode.

ii. Determine the identity of metal X.

c) Describe what happens to the mass of each electrode as the cell operates.

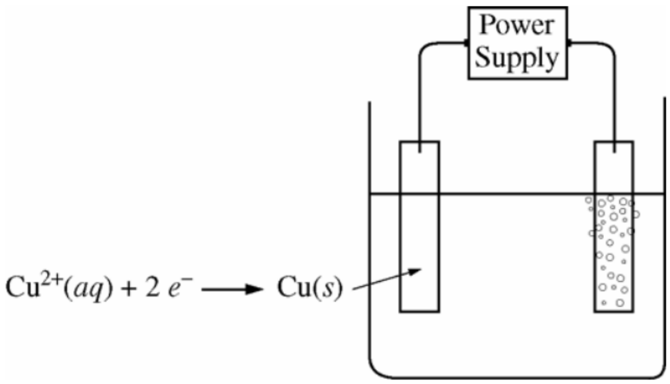
d) During a laboratory session, students set up the electrochemical cell shown above. For each of the following three scenarios, choose the correct value of the cell voltage and justify your answer.

i. A student bumps the cell setup, resulting in the salt bridge losing contact with the solution in the cathode compartment. Is V equal to 0.47 or is V equal to 0? Justify your choice.

ii. A student spills a small amount of 0.5 M $Na_2SO_4(aq)$ into the compartment with the Pb electrode, resulting in the formation of a precipitate. Is V greater than 0.47, equal to 0.47, or less than 0.47? Justify your choice.

iii. After the laboratory session is over, a student leaves the switch closed. The next day, the student opens the switch and reads the voltmeter. Is V greater than 0.47, equal to 0.47, or less than 0.47? Justify your choice.

Unit 3: AP Free Response Practice #4 (2007 FR #3, modified) [10 points]



3. An external direct-current power supply is connected to two platinum electrodes immersed in a beaker containing 1.0 M $CuSO_4$(aq) at 25°C, as shown in the diagram above. As the cell operates, copper metal is deposited onto one electrode and O_2(g) is produced at the other electrode. The two reduction half-reactions for the overall reaction that occurs in the cell are shown in the table below.

Half-Reaction	$E°$(V)
$O_2(g) + 4\,H^+(aq) + 4\,e^- \rightarrow 2\,H_2O(l)$	+1.23
$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$	+0.34

a. On the diagram, indicate the direction of electron flow in the wire.

b. Write the balanced net ionic equation for the electrolysis reaction that occurs in the cell.

c. Predict the algebraic sign of ΔG° for the reaction. Justify your prediction.

d. Calculate the value of ΔG° for the reaction.

An electric current of 1.50 amps passes through the cell for 40.0 minutes.

e. Calculate the mass, in grams, of the Cu(s) that is deposited on the electrode.

f. Calculate the dry volume, in liters measured at 25°C and 1.16 atm, of the O_2(g) that is produced.

AP Free Response Practice #5 (2015 #1, Form B)

Half-Reaction	$E°$ (V)
$2\ CO_2(g) + 12\ H^+(aq) + 12\ e^- \rightarrow C_2H_5OH(aq) + 3\ H_2O(l)$	– 0.085
$O_2(g) + 4\ H^+(aq) + 4\ e^- \rightarrow 2\ H_2O(l)$	1.229

2. A student used a galvanic cell to determine the concentration of ethanol, C_2H_5OH, in an aqueous solution. The cell is based on the half-cell reactions represented in the table above.

   a. Write a balanced equation for the overall reaction that occurs in the cell.

   b. Calculate E° for the overall reaction that occurs in the cell.

   c. A 10.0 mL sample of $C_2H_5OH(aq)$ is put into the electrochemical cell. The cell produces an average current of 0.10 amp for 20. seconds, at which point the $C_2H_5OH(aq)$ has been totally consumed.

      i. Calculate the charge, in coulombs, that passed through the cell.

      ii. Calculate the initial $[C_2H_5OH]$ in the solution.

Polyatomic Ions

Br–Based Ions

BrO^-	hypobromite
BrO_2^-	bromite
BrO_3^-	bromate
BrO_4^-	perbromate

Cr–Based Ions

CrO_4^{2-}	chromate
$Cr_2O_7^{2-}$	dichromate

I–Based Ions

IO^-	hypoiodite
IO_2^-	iodite
IO_3^-	iodate
IO_4^-	periodate

P–Based Ions

PO_3^{3-}	phosphite
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate

Other Ions

CN^-	cyanide
O_2^{2-}	peroxide
MnO_4^-	permanganate

C–Based Ions

CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate *or* bicarbonate
$C_2H_3O_2^-$ *or* CH_3COO^-	acetate
$C_4H_4O_6^{2-}$	tartrate
$C_2O_4^{2-}$	oxalate

Cl–Based Ions

ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate

N–Based Ions

NO_2^-	nitrite
NO_3^-	nitrate
NH_4^+	ammonium

S–Based Ions

SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate
$S_2O_3^{2-}$	thiosulfate

Acid & Base Ions

H_3O^+	hydronium
OH^-	hydroxide

PERIODIC TABLE OF THE ELEMENTS

Atomic number — 14
Symbol — Si
Atomic mass — 28.09
Name — Silicon

	1	2											3	4	5	6	7	8
1	1 **H** 1.008 Hydrogen																	2 **He** 4.00 Helium
2	3 **Li** 6.94 Lithium	4 **Be** 9.01 Beryllium											5 **B** 10.81 Boron	6 **C** 12.01 Carbon	7 **N** 14.01 Nitrogen	8 **O** 16.00 Oxygen	9 **F** 19.00 Fluorine	10 **Ne** 20.18 Neon
3	11 **Na** 22.99 Sodium	12 **Mg** 24.30 Magnesium											13 **Al** 26.98 Aluminum	14 **Si** 28.09 Silicon	15 **P** 30.97 Phosphorus	16 **S** 32.06 Sulfur	17 **Cl** 35.45 Chlorine	18 **Ar** 39.95 Argon
4	19 **K** 39.10 Potassium	20 **Ca** 40.08 Calcium	21 **Sc** 44.96 Scandium	22 **Ti** 47.90 Titanium	23 **V** 50.94 Vanadium	24 **Cr** 52.00 Chromium	25 **Mn** 54.94 Manganese	26 **Fe** 55.85 Iron	27 **Co** 58.93 Cobalt	28 **Ni** 58.69 Nickel	29 **Cu** 63.55 Copper	30 **Zn** 65.39 Zinc	31 **Ga** 69.72 Gallium	32 **Ge** 72.59 Germanium	33 **As** 74.92 Arsenic	34 **Se** 78.96 Selenium	35 **Br** 79.90 Bromine	36 **Kr** 83.80 Krypton
5	37 **Rb** 85.47 Rubidium	38 **Sr** 87.62 Strontium	39 **Y** 88.91 Yttrium	40 **Zr** 91.22 Zirconium	41 **Nb** 92.91 Niobium	42 **Mo** 95.94 Molybdenum	43 **Tc** (98) Technetium	44 **Ru** 101.10 Ruthenium	45 **Rh** 102.91 Rhodium	46 **Pd** 106.42 Palladium	47 **Ag** 107.87 Silver	48 **Cd** 112.41 Cadmium	49 **In** 114.82 Indium	50 **Sn** 118.71 Tin	51 **Sb** 121.75 Antimony	52 **Te** 127.60 Tellurium	53 **I** 126.91 Iodine	54 **Xe** 131.29 Xenon
6	55 **Cs** 132.91 Cesium	56 **Ba** 137.33 Barium	57 ***La** 138.91 Lanthanum	72 **Hf** 178.49 Hafnium	73 **Ta** 180.95 Tantalum	74 **W** 183.85 Tungsten	75 **Re** 186.21 Rhenium	76 **Os** 190.2 Osmium	77 **Ir** 192.2 Iridium	78 **Pt** 195.08 Platinum	79 **Au** 196.97 Gold	80 **Hg** 200.59 Mercury	81 **Tl** 204.38 Thallium	82 **Pb** 207.2 Lead	83 **Bi** 208.98 Bismuth	84 **Po** (209) Polonium	85 **At** (210) Astatine	86 **Rn** (222) Radon
7	87 **Fr** (223) Francium	88 **Ra** 226.02 Radium	89 **†Ac** 227.03 Actinium	104 **Rf** (261) Rutherfordium	105 **Db** (262) Dubnium	106 **Sg** (266) Seaborgium	107 **Bh** (264) Bohrium	108 **Hs** (277) Hassium	109 **Mt** (268) Meitnerium	110 **Ds** (271) Darmstadtium	111 **Rg** (272) Roentgenium							

***Lanthanide Series**	58 **Ce** 140.12 Cerium	59 **Pr** 140.91 Praseodymium	60 **Nd** 144.24 Neodymium	61 **Pm** (145) Promethium	62 **Sm** 150.4 Samarium	63 **Eu** 151.97 Europium	64 **Gd** 157.25 Gadolinium	65 **Tb** 158.93 Terbium	66 **Dy** 162.5 Dysprosium	67 **Ho** 164.93 Holmium	68 **Er** 167.26 Erbium	69 **Tm** 168.93 Thulium	70 **Yb** 173.04 Ytterbium	71 **Lu** 174.97 Lutetium
†Actinide Series	90 **Th** 232.04 Thorium	91 **Pa** 231.04 Protactnium	92 **U** 238.03 Uranium	93 **Np** (237) Neptumium	94 **Pu** (244) Plutonium	95 **Am** (243) Americium	96 **Cm** (247) Curium	97 **Bk** (247) Berkelium	98 **Cf** (251) Californium	99 **Es** (252) Einsteinium	100 **Fm** (257) Fermium	101 **Md** (258) Mendelevium	102 **No** (259) Nobelium	103 **Lr** (262) Lawrencium

STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			$E°$(V)
$F_2(g) + 2e^-$	→	$2F^-$	2.87
$Co^{3+} + e^-$	→	Co^{2+}	1.82
$Au^{3+} + 3e^-$	→	$Au(s)$	1.50
$Cl_2(g) + 2e^-$	→	$2Cl^-$	1.36
$O_2(g) + 4H^+ + 4e^-$	→	$2H_2O(l)$	1.23
$Br_2(l) + 2e^-$	→	$2Br^-$	1.07
$2Hg^{2+} + 2e^-$	→	Hg_2^{2+}	0.92
$Hg^{2+} + 2e^-$	→	$Hg(l)$	0.85
$Ag^+ + e^-$	→	$Ag(s)$	0.80
$Hg_2^{2+} + 2e^-$	→	$2Hg(l)$	0.79
$Fe^{3+} + e^-$	→	Fe^{2+}	0.77
$I_2(s) + 2e^-$	→	$2I^-$	0.53
$Cu^+ + e^-$	→	$Cu(s)$	0.52
$Cu^{2+} + 2e^-$	→	$Cu(s)$	0.34
$Cu^{2+} + e^-$	→	Cu^+	0.15
$Sn^{4+} + 2e^-$	→	Sn^{2+}	0.15
$S(s) + 2H^+ + 2e^-$	→	$H_2S(g)$	0.14
$2H^+ + 2e^-$	→	$H_2(g)$	0.00
$Pb^{2+} + 2e^-$	→	$Pb(s)$	−0.13
$Sn^{2+} + 2e^-$	→	$Sn(s)$	−0.14
$Ni^{2+} + 2e^-$	→	$Ni(s)$	−0.25
$Co^{2+} + 2e^-$	→	$Co(s)$	−0.28
$Cd^{2+} + 2e^-$	→	$Cd(s)$	−0.40
$Cr^{3+} + e^-$	→	Cr^{2+}	−0.41
$Fe^{2+} + 2e^-$	→	$Fe(s)$	−0.44
$Cr^{3+} + 3e^-$	→	$Cr(s)$	−0.74
$Zn^{2+} + 2e^-$	→	$Zn(s)$	−0.76
$2H_2O(l) + 2e^-$	→	$H_2(g) + 2OH^-$	−0.83
$Mn^{2+} + 2e^-$	→	$Mn(s)$	−1.18
$Al^{3+} + 3e^-$	→	$Al(s)$	−1.66
$Be^{2+} + 2e^-$	→	$Be(s)$	−1.70
$Mg^{2+} + 2e^-$	→	$Mg(s)$	−2.37
$Na^+ + e^-$	→	$Na(s)$	−2.71
$Ca^{2+} + 2e^-$	→	$Ca(s)$	−2.87
$Sr^{2+} + 2e^-$	→	$Sr(s)$	−2.89
$Ba^{2+} + 2e^-$	→	$Ba(s)$	−2.90
$Rb^+ + e^-$	→	$Rb(s)$	−2.92
$K^+ + e^-$	→	$K(s)$	−2.92
$Cs^+ + e^-$	→	$Cs(s)$	−2.92
$Li^+ + e^-$	→	$Li(s)$	−3.05

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	=	liter(s), milliliter(s)	mm Hg	=	millimeters of mercury
g	=	gram(s)	J, kJ	=	joule(s), kilojoule(s)
nm	=	nanometer(s)	V	=	volt(s)
atm	=	atmosphere(s)	mol	=	mole(s)

ATOMIC STRUCTURE

$E = h\nu$

$c = \lambda\nu$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number $= 6.022 \times 10^{23}$ mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}, \text{ where } a\,\mathrm{A} + b\,\mathrm{B} \rightleftarrows c\,\mathrm{C} + d\,\mathrm{D}$$

$$K_p = \frac{(P_\mathrm{C})^c(P_\mathrm{D})^d}{(P_\mathrm{A})^a(P_\mathrm{B})^b}$$

$$K_a = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

$$K_b = \frac{[\mathrm{OH^-}][\mathrm{HB^+}]}{[\mathrm{B}]}$$

$$K_w = [\mathrm{H^+}][\mathrm{OH^-}] = 1.0 \times 10^{-14} \text{ at } 25^\circ\mathrm{C}$$
$$= K_a \times K_b$$

$$\mathrm{pH} = -\log[\mathrm{H^+}],\ \mathrm{pOH} = -\log[\mathrm{OH^-}]$$

$$14 = \mathrm{pH} + \mathrm{pOH}$$

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{A^-}]}{[\mathrm{HA}]}$$

$$\mathrm{p}K_a = -\log K_a,\ \mathrm{p}K_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$[\mathrm{A}]_t - [\mathrm{A}]_0 = -kt$$

$$\ln[\mathrm{A}]_t - \ln[\mathrm{A}]_0 = -kt$$

$$\frac{1}{[\mathrm{A}]_t} - \frac{1}{[\mathrm{A}]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{\text{total}} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE_{\text{molecule}} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = \varepsilon bc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
ε = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1} K^{-1}
= 0.08206 L atm mol^{-1} K^{-1}
= 62.36 L torr mol^{-1} K^{-1}

1 atm = 760 mm Hg = 760 torr

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS / ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\begin{aligned} \Delta G^\circ &= \Delta H^\circ - T\Delta S^\circ \\ &= -RT \ln K \\ &= -nFE^\circ \end{aligned}$$

$$I = \frac{q}{t}$$

$$E_{cell} = E_{cell}^\circ - \frac{RT}{nF}\ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 **H** 1.008																	2 **He** 4.00
3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
11 **Na** 22.99	12 **Mg** 24.30											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.06	17 **Cl** 35.45	18 **Ar** 39.95
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.10	45 **Rh** 102.91	46 **Pd** 106.42	47 **Ag** 107.87	48 **Cd** 112.41	49 **In** 114.82	50 **Sn** 118.71	51 **Sb** 121.75	52 **Te** 127.60	53 **I** 126.91	54 **Xe** 131.29
55 **Cs** 132.91	56 **Ba** 137.33	57 ***La** 138.91	72 **Hf** 178.49	73 **Ta** 180.95	74 **W** 183.85	75 **Re** 186.21	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.08	79 **Au** 196.97	80 **Hg** 200.59	81 **Tl** 204.38	82 **Pb** 207.2	83 **Bi** 208.98	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
87 **Fr** (223)	88 **Ra** 226.02	89 **†Ac** 227.03	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (271)	111 **Rg** (272)							

***Lanthanide Series**	58 **Ce** 140.12	59 **Pr** 140.91	60 **Nd** 144.24	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 151.97	64 **Gd** 157.25	65 **Tb** 158.93	66 **Dy** 162.5	67 **Ho** 164.93	68 **Er** 167.26	69 **Tm** 168.93	70 **Yb** 173.04	71 **Lu** 174.97
†Actinide Series	90 **Th** 232.04	91 **Pa** 231.04	92 **U** 238.03	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (262)