Pre–AP® Chemistry

Summer Assignment

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Chemistry – the study of matter and energy and more importantly, the changes between them

Why study chemistry? When a car starts do you think about chemistry? When a battery in your cell-phone goes dead do you think about chemistry? In fact, the food you eat for lunch provides energy, through chemical reactions, and you are able to read these sentences and comprehend them because of chemical reactions occurring in your brain. Did you know the theory that a huge meteor hit the earth 65 million years ago, causing the extinction of the dinosaurs, was first recognized as a plausible explanation by chemists, who noticed that rocks from that time period contained amounts of the elements iridium (Ir) and niobium (Nb) that are seen only in meteors. All of these ideas can be understood with a basic foundation in chemistry. Chemistry is around you all the time; you encounter chemistry every waking moment of your life, whether you recognize it or not. You will become a better problem solver in all areas of your life; and better understand all areas of science…

Units of Measurement

A big part of chemistry, and of all sciences, is making observations. Chemists rely upon 2 types of observations

1. **Qualitative:** an observation – of a "quality nature": color changes, the way something looks, smoke was given off, a solid (precipitate) forms, it's raining outside, etc…
2. **Quantitative:** a measurement – consists of 2 parts – both of which MUST BE PRESENT in order for the measurement to have meaning. Number & Unit

   Ex: 250 mL

Metric prefixes – You don't have to know the nature of a unit (gram, liter, meter, etc…) to convert it; for example, from kilo-unit to micro-*unit*. All metric prefixes are powers of 10. The most commonly used prefixes are shaded…

YOU Must be able to convert the common scientific metric prefixes to the desired unit of measurement.

Prefix	Symbol	Factor
yotta	Y	10^{24} = 1,000,000,000,000,000,000,000,000
zetta	Z	10^{21} = 1,000,000,000,000,000,000,000
exa	E	10^{18} = 1,000,000,000,000,000,000
peta	P	10^{15} = 1,000,000,000,000,000
tera	T	10^{12} = 1,000,000,000,000
giga	G	10^{9} = 1,000,000,000
mega	M	10^{6} = 1,000,000
kilo	**k**	**10^{3} = 1,000**
hecto	h	10^{2} = 100
deka	da	10^{1} = 10
deci	**d**	10^{-1} = 0.1
centi	**c**	10^{-2} = 0.01
milli	**m**	10^{-3} = 0.001
micro	**µ**	10^{-6} = 0.000,001
nano	n	10^{-9} = 0.000,000,001
pico	p	10^{-12} = 0,000,000,000,001
femto	f	10^{-15} = 0.000,000,000,000,001

EXAMPLE PROBLEM:

1. 250 mg of salt is the same as 0.250 g of salt
2. 2.5 L is the same as 2500 mL
3. 55.5 kg is the same as 55,500 g

All measurements are made according to the International System of Units – (SI system):

1. Based upon the metric system and its measurements
2. Utilizes prefixes to change the size of the unit

Table 1.1 The Fundamental SI Units

Physical Quantity	*Name of Unit*	*Abbreviation*
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

Table 1.3 Some Examples of Commonly Used Units

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

Mass – amount of matter an object contains. Also defined as the force giving an object acceleration, which is also called weight; however the weight of an object changes with changes in gravity (earth and moon) whereas the mass of an object will remain the same.

Measured in **GRAMS**, or **milligrams, micrograms, kilograms, etc…**

Measured using a laboratory balance; We mass (often misspoken as "weigh") quantities on a balance NOT a scale!!

Mass vs. Weight – chemists are quite guilty of using these terms interchangeably

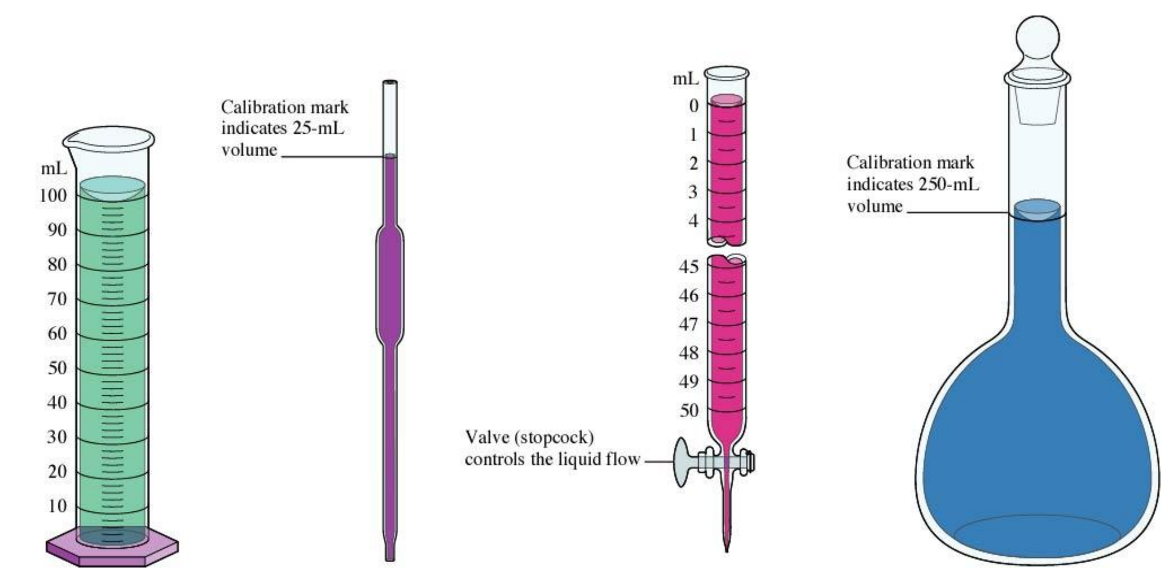
- Mass (g or kg) – a measure of the resistance of an object to a change in its state of motion (i.e. exhibits inertia); the quantity of matter present.
- Weight (a force $\therefore$ Newton's) – the response of mass to gravity; since all of our measurements will be made here on Earth, considered the acceleration due to gravity

Volume – amount of space occupied by an object; this measurement is derived from measurements of length

Measured in **LITERS** – defined as 1 dm $\times$ 1dm $\times$ 1 dm = 1dm^3 or a cubic decimeter

Measured in **MILLILITERS** - defined as 1 cm $\times$ 1cm $\times$ 1 cm = 1cm^3 or a cubic centimeter (also know in the medical industry as a "CC")

Common devices used to measure volume



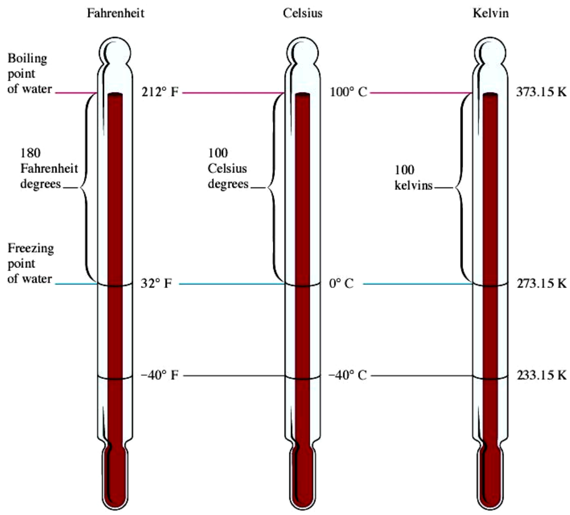
100-mL graduated cylinder | 25-mL pipet | 50-mL buret | 250-mL volumetric flask

moderately accurate | very accurate | very accurate | very accurate

Temperature – measure of the average kinetic energy in a sample of matter

Measured using a thermometer in units of:

1. Celsius – 0°C freezing pt. of water; 100°C is the boiling pt. of water
2. Kelvin – 0 K is absolute zero (coldest possible temperature –absence of any energy); 273 K is the freezing pt. of water and 373K is the boiling point of water



Notice for 1 °C change you also get a change of 1K

$T_K = T_C + 273$ K

$T_C = T_K - 273$ºC

$T_F = T_C \times \frac{9°F}{5°C} + 32$ºF

EXAMPLE PROBLEM:

1. The temperature in a flask used in a gas collection experiment was 22.5°C; determine the temperature in Kelvin.

$T_K = T_C + 273$ K $\qquad T_K = 22.5 + 273$ K $= 295.5$ K

Density – the amount of mass per unit volume; often used to identify substances

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

EXAMPLE PROBLEM:

1. Diamonds are measured in Carats; where 1 carat = 0.200g. The density of a diamond is 3.51 g/cm^3. Determine the volume of a 5.0 carat diamond (1.00 g).

$$\text{Density} = \frac{\text{mass}}{\text{volume}}\text{; } 3.51\ \text{g/cm}^3 = \frac{1.00\ \text{g}}{\text{volume}}\text{; Volume} = \frac{1.00\ \text{g}}{3.51\ \text{g/cm}^3} = 0.285\ \text{cm}^3$$

Significant Figures – a measurement consists of all values of the measurement that are known with certainty, plus one digit that is estimated.

Significant Figures and Calculations
There are two types of numbers you will encounter in science; exact numbers and measured numbers

- Exact numbers are absolutely correct and are determined by counting or defining; i.e. 12 in a dozen; 24 hrs in a day…
- Measured numbers are determined and therefore involve some estimation (depending on the capacity of the measuring device). In measured numbers we MUST count the number of significant digits (the digits believed to be correct by the person making the measurement)

Rules

- (1). Non zero digits are significant
- (2). A zero is significant if it is
  - (a) - "terminating AND right" of the decimal [must be both]
  - (b) - "sandwiched" between significant figures
- Exact or counting numbers have an ∞ amount of significant figures as do constants

EXAMPLE PROBLEMS:

- 3.57 mL has 3 significant figures (digits) [Rule 1]
- 20.80 mL has 4 significant figures [Rules 1, 2a, and 2b]
- 0.01 mL has 1 significant figure [Rule 1]
- 0.0100 mL has 3 significant figures [Rule 1 and 2b]

Rules for calculating with measured values

- × and ÷
  - The term with the least number of *significant figures* (∴ least accurate measurement) determines the number of significant figures in the answer.

$$4.56 \times 1.4 = 6.38 \xrightarrow{\text{corrected}} 6.4$$

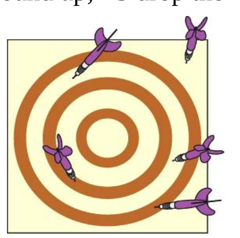
- + and (-) The term with the least number of *decimal places* (∴ least accurate measurement) determines the number of significant figures in the answer.

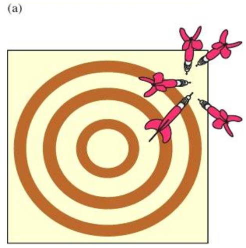
$$\begin{array}{ll} 12.11 & \\ 18.0 & \leftarrow \text{limiting term} \\ \underline{1.013} & \\ 31.123 \xrightarrow{\text{corrected}} 31.1 & \end{array}$$

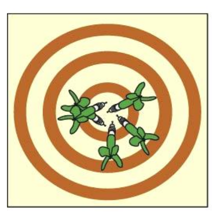
- pH – the number of significant figures in least accurate measurement determines number *decimal places* in the answer.

Rounding Rules:

- Round at the end of all calculations
- Look at the significant figure one place beyond your desired number of significant figures if >5 round up; <5 drop the digit.
- Don't "double round" 4.348 to 2 SF = 4.3 NOT the 8 makes the 4 a 5 then 4.4. [Even though you may have conned an English teacher into this before!]


(a)


(b)


(c)

Precision and Accuracy – when making measurement the precision and accuracy of the measurement are very important

- accuracy – correctness; agreement of a measurement with the true value
- precision – reproducibility; degree of agreement among several measurements.
- random or indeterminate error – equal probability of a measurement being high or low
- systematic or determinate error – occurs in the same direction each time

EXAMPLE PROBLEM: Look at the figure to the right and determine whether each bulls eye is accurate, precise, neither, or both.

- Figure (a) – Neither accurate nor precise (large random errors)
- Figure (b) – Precise but not accurate (small random errors, large systematic error)
- Figure (c) – Bull's-eye! Both precise and accurate (small random errors, no systematic error)

PRACTICE PROBLEMS:

How many significant figures are there in the following measurements?

1. 0.00510 km
2. 2550 g
3. 2.54×10^{-5} L
4. 2.002 seconds
5. 1.14×10^{-3} moles

Convert the following metric measurements; be sure to report your answer with the proper number of significant figures:

2. 4.56 cm x 1.4 cm x 6.038 cm = ____________________ cm^3.
3. 1.050×10^{-3} g / 6.14 mL = ________________________ g/mL
4. (0.1022 x 0..0821 x 273)/1.01 = ____________________atm
5. Convert 250 mL to a measurement in liters:____________________.
6. Convert 0.500 grams to a measurement in milligrams (mg)______________________________.
7. Convert 35.2 cm to kilometers (km)__________________________.
8. Convert 0.025 µL to milliliters (mL)__________________________.

What is Chemistry

Chemistry is chiefly concerned with the composition and structure of matter, how different forms of matter change from one form to another; and the energy changes that these processes undergo.

Matter: anything occupying space and having mass; the material of the universe

Exists in three states:

1. Solid – rigid; definite shape and volume; *molecules close together vibrating about fixed points ∴ virtually incompressible*
2. Liquid – definite volume but takes on the shape of the container; molecules still vibrate but also have rotational and translational motion and can slide past one another BUT are still close together ∴ slightly compressible
3. Gas – no definite volume and takes on the shape of the container; *molecules vibrate, rotate and translate and are independent of each other ∴ VERY far apart ∴ highly compressible*

   - *vapor* – the gas phase of a substance that is normally a solid or liquid at room temperature
   - *fluid* – matter which can flow; gases and liquids

Most of the matter around us consists of **mixtures** of **pure substances**

Ex: wood is a mixture of many substances, as is any living (or previously living) organism

Mixtures**:** have variable composition; 2 types of mixtures

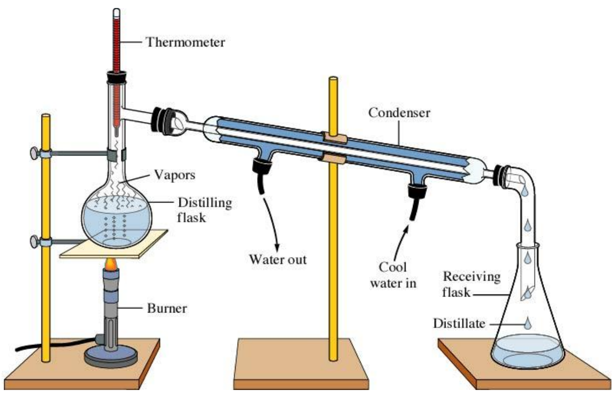
1. *Heterogeneous* (having visibly distinguishable parts)
   a. Sand in water is a heterogeneous mixture
   b. Ice cubes in a soda is a heterogeneous mixture
2. *Homogeneous* (having visibly indistinguishable parts)
   a. Also called **solutions**
   b. Air is a solution of gases (~80% Nitrogen and ~20% Oxygen)
   c. Wine or a soda is a complex solution made of many, equally mixed parts
   d. Brass is a solution of copper and zinc

A **heterogeneous mixture** can be separated into a **homogeneous mixture** by physical methods.

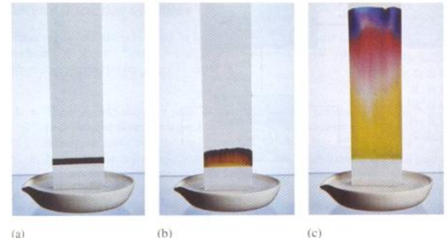
A **homogeneous mixture** (think solution) can be separated into **pure substances** by physical methods

Physical Separation Methods:

1. **Filtration** – using filter paper (like a coffee filter) to separate heterogeneous mixtures that consist of a solid and liquid
2. **Distillation** – used to separate homogeneous mixtures based upon differences in their abilities to change into a gas
   a. Used to purify tap water (tap water is not PURE; it is a solution of water and other dissolved substances)



3. **Chromatography** – separates homogeneous mixtures that consist of solid and liquid phases; i.e. ink from a pen
    a. Paper chromatography can separate ink into its individual substances, where the substances separate based upon how easily they move with water
    b. Tie dying is a rudimentary example of chromatography.



Pure Substances – form of matter with constant composition – in chemistry these are known as

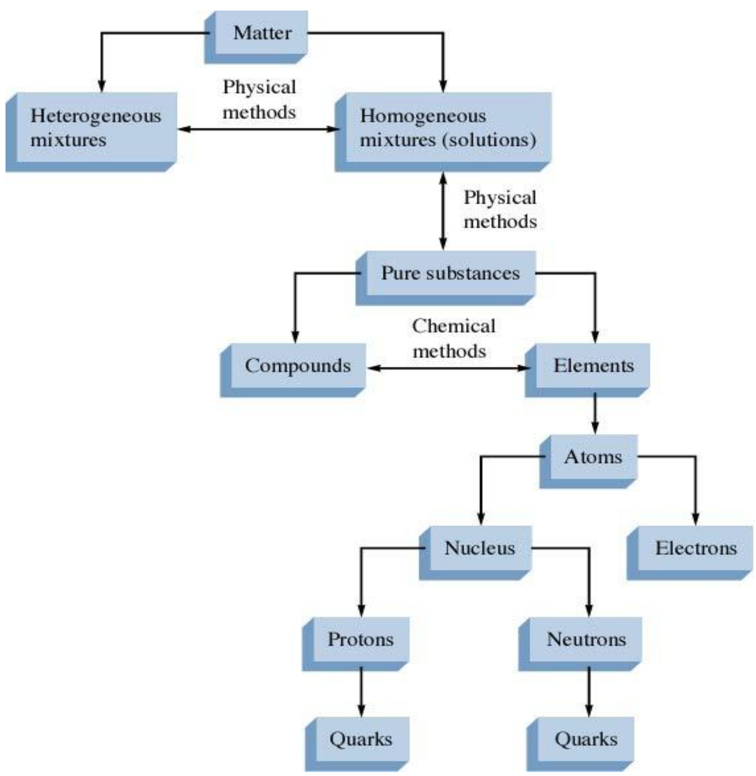
Elements – matter that cannot be broken down into simpler substances by chemical or physical means.

Made of atoms that are uniquely different from each other due to the number and types of sub-atomic particles (protons, electrons, and neutrons)

Compounds – substance with *constant composition* that can be broken down into elements by **chemical processes (called chemical reactions)**

Made of a combination of atoms that result in a substance's properties that are uniquely different from those of the individual atoms

Flow Diagram of Matter



ATOMS

Dalton's Atomic Theory (1808)

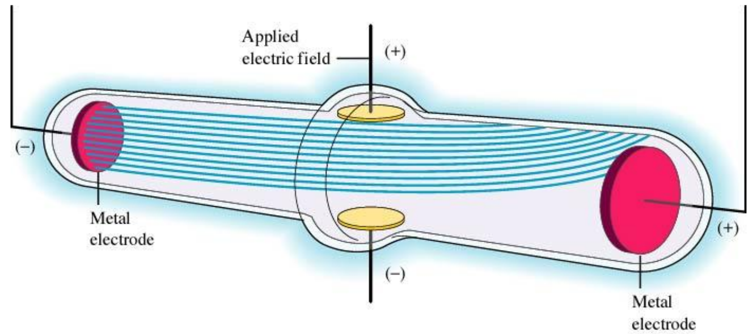
1. Each element is made up of tiny particles called atoms
2. Atoms of a given element are identical; atoms are different elements are different*
3. Chemical compounds are formed when atoms of different elements combine with each other in simple whole number ratios (i.e. water is 2 atoms of Hydrogen and 1 atom of Oxygen – H_2O)
4. Chemical reactions involve reorganization of atoms – i.e. changes in the way they are bound together – atoms are never created or destroyed in chemical reactions.

* - statement adjusted with the discovery of isotopes (discussed below)

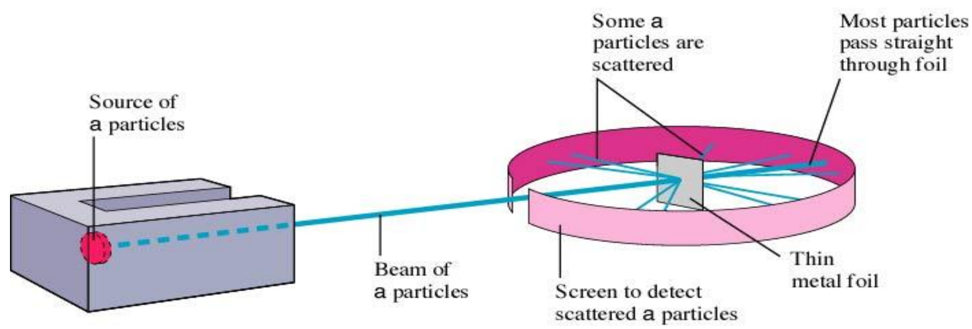
This theory lead to the discoveries of the ***sub-atomic particles*** in the late 1800's and early 1900's, which ultimately determined the structure of the atom

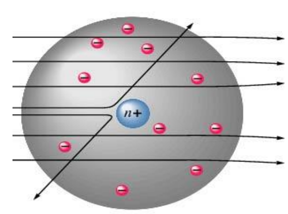
Atomic Structure

Electrons – (~1900) J.J. Thomson discovered a stream of particles that were negatively charged – called the electron (attracted to the positive plate in an electric field – remember opposites attract!)



Nucleus – (~1911) E. Rutherford discovered the nucleus of the atom in the "Gold Foil" Experiment. Tiny particles with a positive charge were ejected towards a thin sheet of gold foil. A number of the particles were deflected back, suggesting that the atom has a high-density core with a positive charge – called this core the nucleus





1. ***Protons*** – positive charge that is equal in magnitude to the electron's negative charge; also is ~1000 times heavier than the electron
2. ***Neutron*** – neutral particle that is equal in mass to the proton, but carries no charge

Notable facts about the nucleus:

- ~10,000 times smaller than the atom itself
- very high density as it accounts for almost all of the atoms mass
  - if an atom's nucleus were the size of a pea, it would have a mass of 250 million tons!
- Protons and neutrons are held together by nuclear binding energy –a lot of energy – think of the atomic bombs dropped in Japan at the end of WWII - that is where the energy came from….

Question for Thought?
"If atoms are all composed of the same components, how are atoms different?"

Dalton was slightly wrong when he theorized that all atoms of the same element are identical – actually all atoms of the same element have EXACTLY the same NUMBER OF PROTONS

- **Atomic number - number of protons in the nucleus of an atom**
  - Which is a whole number on the periodic table, usually just above the symbol of the element

Atomic Number →	6
Element's Symbol ←	**C**
Average Atomic Mass →	**12.011**

- Since Carbon has an atomic number of 6, it has 6 protons
- Since an atom is electrically NEUTRAL, it must have the same number of electrons as protons (negative charges (electrons) cancel the positive charges (protons))
- Atoms of the same element that contain different numbers of neutrons are called **Isotopes**

Example:
The element carbon (C) is actually made from 3 types of atoms (all of which have 6 protons), which are isotopes of each other:

- The three types of Carbon atoms are those with 6 neutrons, with 7 neutrons, and some atoms with 8 neutrons
- Since each have different numbers of neutrons, each type (or isotope) has a different mass, thus a different **Mass Number – number of neutrons and protons**

  - The Mass of the atom is determined solely by the number of PROTONS and NEUTRONS comprising the nucleus – called the Mass number
    - ☑ IMPORTANT: Not the same as the average atomic mass that is given on the periodic table

- 2 ways to identify isotopes of the same element:
  1. Write the name followed by the mass number: Name – mass number
     - ☑ Example: Carbon –12; Carbon – 13; Carbon –14
  2. Write the Atomic Symbol
     - ☑ Mass Number → A; Atomic Number → Z; $^{A}_{Z}X$ ← Element's Symbol

       $^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$

Sample problems: Determine the number of protons (p), electrons (e), and neutrons (n) in the following atoms:

$^{238}_{94}Pu$ Atomic number = 94 $\boxed{\#p = 94}$ #p = #e thus there are $\boxed{\text{94 electrons}}$
Mass number = 238 #p + # n = 238 #n = 238 – 94 thus $\boxed{\#n = 144}$

$^{65}_{29}Cu$ Atomic number = 29 $\boxed{\#p = 29}$ #p = #e thus there are $\boxed{\text{29 electrons}}$
Mass number = 65 #p + # n = 65 #n = 65 – 29 thus $\boxed{\#n = 36}$

$^{52}_{24}Cr$ Atomic number = 24 $\boxed{\#p = 24}$ #p = #e thus there are $\boxed{\text{24 electrons}}$
Mass number = 52 #p + # n = 528 #n = 52 – 24 thus $\boxed{\#n = 28}$

$^{54}_{24}Cr$ Atomic number = 24 $\boxed{\#p = 24}$ #p = #e thus there are $\boxed{\text{24 electrons}}$
Mass number = 54 #p + # n = 54 #n = 54 – 24 thus $\boxed{\#n = 30}$