### Unit 2 Homework Assignments:

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<th>Assignment</th>
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Unit 2 Vocabulary:

1. Alloy: a solution where two metals are dissolved into each other and solidified.
2. Amalgam: a solution where a metal is dissolved into mercury.
3. Aqueous: a solution where a solute is dissolved into water.
4. Atom: the smallest part of an element that still retains the properties of that element.
5. Calorimetry: the measurement of energy change between potential and kinetic energy by measuring the temperature change induced on a measured mass of water in a calorimeter.
6. Change: a transformation from one condition of matter to another.
7. Chemical Property: a property that may only be observed when a chemical change occurs.
8. Compound: matter which results from the bonding of two (or more) different elements to each other; compounds are decomposable back to separate elements.
9. Element: matter which exhibits definite physical and chemical properties unique to itself and different from all other forms of matter. An element may not be decomposed into simpler matter.
10. Endothermic: the conversion of kinetic energy into potential energy.
11. Exothermic: the conversion of potential energy into kinetic energy.
12. Heterogeneous: matter that is unevenly distributed throughout a volume.
13. Homogeneous: matter that is evenly distributed throughout a volume.
Unit 2 Vocabulary (cont’d)

14. Ion: a charged particle formed when an atom gains or loses electron(s).
16. Matter: Anything that exists with mass and volume.
17. Metal: an element that loses electrons when forming chemical bonds.
18. Metalloid: an element that exhibits properties of both metals and nonmetals.
19. Mixture: matter of different types that are physical proximity to each other, yet not chemically combined.
20. Nonmetal: an element that gains electron(s) from metals or shares electron(s) from other nonmetals when forming chemical bonds.
21. Period Table: a chart that lists the elements in order of increasing atomic number and arranges them in groups of similar chemical properties.
22. Physical Property: a change that can be observed without a chemical change occurring.
23. Potential Energy: stored energy, often in form of chemical bonds.
25. Temperature: the average kinetic energy of a sample or system.
26. Tincture a solution where a solute is dissolved into alcohol.
Notes page:
Topic: **Properties & Matter Changes**

Objective: What kinds of matter are there, and may they change?

- Chemistry: the study of **MATTER**, the **CHANGES** matter undergoes, and the **ENERGY** associated with those changes.
- Matter: anything that exists that has definite **MASS** and occupies a **DEFINITIE** volume.
(Elements and compounds) are all HOMOGENEOUS (containing the same composition of material throughout the sample)

1. Elements are substances that cannot be decomposed (broken down) by chemical change. They are made up of ATOMS. Element symbols are either one CAPITALIZED letter, or two letters; the first letter CAPITAL, the second lower-case.

   Examples of elements:

   One letter: O (oxygen)   Two letters: Ni (nickel)

   Note: Cl (chlorine) is C and a lower-case L

2. Compounds are substances that are made of elements chemically bonded to each other, and can be decomposed by chemical change back into separate elements.

   Examples of compounds:

   NaCl (sodium & chlorine)   Cu₃(PO₄)₂ (copper, phosphorous, & oxygen)
Matter and Energy

### Topic: Properties & Matter Changes

**Objective:** What kinds of matter are there, and may they change?

#### Mixtures:

Mixtures are combinations of substances that are not chemically combined together, and they may be broken apart by physical change.

- **Homogenous** mixtures are called **SOLUTIONS**.

#### Examples of homogenous mixtures include:

1. Aqueous solutions: $\text{NaCl}_{(aq)}$ aqueous means that the solid solute (NaCl) was dissolved into WATER. This is an example of a solution. $\text{NaCl}_{(aq)}$ cannot be separated by filtering. To separate the $\text{NaCl}_{(aq)}$ from the water you must evaporate the water. Aqueous solutions are transparent in that you can see through the solution.

2. Tinctures: a solution where the solute is dissolved in alcohol (ethanol). Some solutes that don’t dissolve in water may dissolve in ethanol. This includes tincture of iodine, used for medical disinfection. Tincture of iodine is iodine dissolved into alcohol.

3. Amalgam: a solid solution where a metal is dissolved into liquid mercury. Metals commonly used to make amalgams with mercury are silver (Ag) and gold (Au).

4. Alloys: metals can’t chemically bond with different metals, but they can be mixed together to enhance their combined properties. Iron (Fe) is a strong metal used for structures, but it is soft and rusts. Mixing in carbon (C) forms the alloy steel, and steel can be enhanced by adding in other metals, such as chromium (Cr), molybdenum (Mb), or cadmium (Cd).
Topic: **Properties & Matter Changes**

Objective: What kinds of matter are there, and may they change?

- **Heterogeneous** mixtures have **varying** compositions throughout the sample.

**Examples of heterogeneous mixtures include:**

1. Muddy water. The composition towards the bottom has more sediment, but towards the top is more water. Even if shaken, the particles will not be dispersed evenly to homogeneity.
2. Italian salad dressing. The oil, vinegar, and spices are not in solution. The less dense oil floats, the more dense aqueous vinegar is below the oil, and the solids particulates settle to the bottom.
3. Soil. Soil is visible rock fragments, organic debris, and other materials.

**Particle Diagrams:**

- Particle diagrams show how the forms of matter look in a simple form.

1. Elements: single atoms, not bonded to each other.

![A single atom of an element](image)
2. Diatomic molecule: certain elements are so reactive they are most stable as same-type pairs, forming two-atom diatomic molecules. Elements that form diatomic molecules are Br, I, N, Cl, H, O, and F. Diatomic molecules are written Br₂, I₂, N₂, Cl₂, H₂, O₂, and F₂. They are NOT compounds as they are only one element.

3. Compounds: these are made of two or more different elements chemically bonded in a defined, whole-number ratio.

4. Homogenous mixtures: combinations of elements, compounds, or both, in no fixed ratios, and not bonded together, but evenly dispersed.

5. Heterogeneous mixtures: combinations of elements, compounds, or both, in no fixed ratios, and not bonded together, but unevenly evenly dispersed.
Topic: **Physical Changes & Properties**

Objective: What are physical properties, and may they change (Δ)?

- **Physical** Changes: altering only the appearance of a substance, and not its chemical identity.
- **Physical** Properties: properties that may be observed through physical change.

<table>
<thead>
<tr>
<th>Physical Change</th>
<th>Physical Property</th>
</tr>
</thead>
<tbody>
<tr>
<td>Melting</td>
<td>Melting point (temperature @ solid to liquid Δ) MP</td>
</tr>
<tr>
<td>Boiling</td>
<td>Boiling point (temperature @ liquid to gaseous Δ) BP</td>
</tr>
<tr>
<td>Dissolving</td>
<td>Solubility (amount of solute dissolvable into solvent)</td>
</tr>
<tr>
<td>Evaporating</td>
<td>Vapor pressure (pressure of vapor @ gas-liquid equilibrium)</td>
</tr>
<tr>
<td>Crushing</td>
<td>Malleability (hammered or rolled into sheets)</td>
</tr>
<tr>
<td>Stretching</td>
<td>Ductility (stretched into wire)</td>
</tr>
<tr>
<td>No physical Δ</td>
<td>Specific heat (heat to Δ temperature of 1 g by 1 K)</td>
</tr>
<tr>
<td></td>
<td>Heat of Fusion (heat to melt 1 g of solid @ MP)</td>
</tr>
<tr>
<td></td>
<td>Heat of Vaporization (heat to boil 1 g of liquid @ BP)</td>
</tr>
<tr>
<td>Varying amounts</td>
<td>Density (mass per unit volume)</td>
</tr>
<tr>
<td>of physical Δ</td>
<td>Conductivity (ability for heat &amp; current to pass)</td>
</tr>
</tbody>
</table>
Topic: **Chemical Changes & Properties**

Objective: What are chemical properties, and may they change (Δ)?

- Chemical changes: **alterations** in the chemical composition of a substance that may be **reversed** only by another **chemical** change.
- Chemical properties: properties that may **only** be observed through **chemical** change.

<table>
<thead>
<tr>
<th>Chemical Change</th>
<th>Chemical Property</th>
</tr>
</thead>
<tbody>
<tr>
<td>Corrosion of metals; flammability</td>
<td>Reactivity (likelihood of one substance to undergo a chemical reaction with another substance)</td>
</tr>
<tr>
<td>Chemical decomposition (e.g. hydrogen peroxide, H₂O₂, decomposes to form water and oxygen, but water, H₂O, does not decompose spontaneously)</td>
<td>Stability (likelihood that a substance will not decompose)</td>
</tr>
<tr>
<td>Combustion (exothermic release of heat)</td>
<td>Heat of Reaction (energy absorbed or released by a chemical reaction)</td>
</tr>
<tr>
<td>Rare chemical changes absorb heat (endothermic)</td>
<td></td>
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</tbody>
</table>

**Watch Bozeman Science Chemistry Matter video**

https://www.youtube.com/watch?v=jEoQ6TNLJl8&list=PL43285691048DAD00&index=1
Topic: **Law of Conservation of Mass**

Objective: What are the properties of the Law of Conservation?

Law of Conservation of Mass:

“Matter cannot be created nor destroyed by physical or chemical change, but only converted from one form to another form.”

Examples of conservation of Mass:

1. If 40 g of substance ‘A’ are reacted with 20 g of substance ‘B’ to form substance ‘C’, what will the mass of substance ‘C’ be?
   a. Answer: 60 grams. As the total combined mass of reactants (‘A’ & ‘B’) is 60 g, the total mass of the product ‘C’ must be equal.

2. 35 g of liquid water are evaporated off in a closed container. How many grams of water vapor will be in the container after evaporation?
   a. Answer: 35 g. The mass does not change regardless of the phase.

3. Magnesium metal is reacted with oxygen to form solid magnesium oxide product. How will the mass of the magnesium oxide compare to the mass of the metal and gaseous oxygen reactants?
   a. Answer: They are equal. Mass of the products must be equal to the mass of the reactants.
Properties and Changes of Matter homework

Circle your answer for questions 1 and 3.

1. Which of the following **CAN** be decomposed by chemical changes?
   - a) \( \text{SO}_2 \)
   - b) \( \text{N}_2 \)
   - c) \( \text{Ne} \)
   - d) \( \text{Al} \)

2. Explain your answer choice above:
   \( \text{SO}_2 \) is a compound of \( \text{S} \) and \( \text{O} \); can chemically separate them

3. Which of the following **fictitious** element symbols is correctly written?
   - a) \( \text{Cn} \)
   - b) \( \text{HB} \)
   - c) \( \text{zL} \)
   - d) \( \text{r} \)

4. Explain why **EACH** of the incorrect choices above are wrong:
   - Choice \( \text{HB} \): No double capital letter elements
   - Choice \( \text{zL} \): element names start with a capital letter
   - Choice \( \text{r} \): elements begin with (or are only) capitals
Identify the following as either physical or chemical and explain why.

<table>
<thead>
<tr>
<th>Description</th>
<th>Physical/Chemical</th>
<th>Why?</th>
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<tbody>
<tr>
<td>The steel body of a car begins to rust after the paint is worn away.</td>
<td>Chem</td>
<td>Reaction with elements</td>
</tr>
<tr>
<td>Potassium thiocyanate crystals dissolve easily into water.</td>
<td>Phys</td>
<td>Can evaporate water and gain crystals back</td>
</tr>
<tr>
<td>A 10.0 g sample of lauric acid melts upon heating.</td>
<td>Phys</td>
<td>Solidifies upon cooling</td>
</tr>
<tr>
<td>Paradichlorobenzene in solid form melts at a temperature of 52°C when heat is added.</td>
<td>Phys</td>
<td>Solidifies on cooling</td>
</tr>
<tr>
<td>Sodium metal explosively reacts with water to form hydrogen gas and sodium hydroxide.</td>
<td>Chem</td>
<td>Forms new compounds</td>
</tr>
<tr>
<td>110 g of sodium nitrate crystals can dissolve in 100 g of water at 45°C.</td>
<td>Phys</td>
<td>Can evaporate water and gain crystals back</td>
</tr>
<tr>
<td>Platinum metal conducts electricity because its atoms pass electrons easily from one atom to the next.</td>
<td>Phys</td>
<td>Electrons are not lost; passed along</td>
</tr>
<tr>
<td>Gold metal can be hammered or rolled into very thin sheets.</td>
<td>Phys</td>
<td>Ductability - physical characteristic</td>
</tr>
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</table>
Properties and Changes of Matter homework

1. Draw a particle diagram of the compound CaCl₂, using solid circles to represent Ca and empty circles to represent Cl. Draw at least five molecules of CaCl₂ in the box below.

![Particle Diagram]

2. Name the two parts of the mixture known as tincture of iodine.
   
   (Hint: you’ve read about it!) __iodine (I₂)__ & __alcohol__

3. Vacuumed dirt accumulates in a bag or a container. If you examined the contents of a vacuum, would you expect the contents to be **homogenous** or **heterogeneous**? Explain why you answered that way.

   **Mixed dirt from various rooms having different uses and therefore different types of dirt**
4. Could Br be decomposed into simpler substances?
   No, it is an element

5. What elements can the compound Ca(NO₃)₂ decompose into?
   Ca & N & O

6. Brass is made of two metallic elements, copper (Cu) and zinc (Zn). What type of mixture is this?
   Alloy - mixture (solution) of metals
Energy: the ability to do **WORK**, which is using **FORCE** to move an object a **DISTANCE**.

1. **Kinetic Energy (KE)**: Energy of **MOTION**, contained by anything that **MOVES**.
   - Atoms, molecules, and other particles of similar scale move faster when **TEMPERATURE** (T) is increased.

**Temperature scales:**

I. **Celsius**: Based on melting point (**MP**) of water as **0°C** and the boiling point (**BP**) of water as **100°C**.

II. **Kelvin**: Uses Celsius-sized unit, but starts at (**0 K** or -273°C) where all particle motion stops (**absolute zero**). Note there is no ° symbol used with K.

Converting between °C and K:

![Conversion Chart]

Melting point of water, 0°C, is equal to 273 K (0 + 273 = 273)

Boiling point of water, 100°C, is equal to 373 K (100 + 273 = 373)
2. **Potential Energy (PE):** **STORED** energy, energy that is not currently doing work, but has the **ability** if **released**.

- PE is found in coiled **springs**, chemical **bonds** (batteries, explosives, chemical hot/cold packs), objects at a **height**, and **magnetism** (attraction and repulsion).
- PE cannot be directly measured, and must be **converted** to **KINETIC** energy and then measured.
- Joule (J): metric **unit** for potential **energy**. 1000.0 J is a **kiloJoule** (kJ) and is most often used to measure PE changes in physical and chemical Δ (such as burning, melting, etc.)

3. **Heat Flow:**

- Heat flows from HOT to where it is NOT (**Hot → Cold**).
- Heat flow examples:
  - i. Opening the front door on a cold morning; heat flow OUT of house.
  - ii. Place a hot pack on a sore muscle; heat flows into the muscle.
  - iii. Place warm (45°C) metal into tap water (10°C); heat flows out of metal into the water.
Topic: **Measuring Energy (Ε)**

Objective: How can we measure the energy stored in a substance?

**Law of Conservation of Energy (Ε):**

“Energy cannot be created nor destroyed by physical or chemical change, but only converted from one form to another form.”

How can stored energy in a substance be measured using Calorimetry?

1. **Exothermic** changes: PE stored in a substance is released and converted into KE, which is absorbed by a sample of water whose mass is known. The temperature of the water INCREASES at the rate of 4.18 J/g°C. This SPECIFIC HEAT of water means each added 4.18 J raises each gram of water 1°C.
   - The change in heat (ΔH) is NEGATIVE: (PE \(\rightarrow\) KE; stored energy released as KE)

**Exothermic example:**

- Burning a candle releases heat, which is absorbed by a measured mass of water in a calorimetry cup. The water temperature will INCREASE.
2. **Endothermic** changes: KE from surroundings (water in a calorimetry cup) is **absorbed** by the change and **converted** into PE. The water temperature will **DECREASE**.

- The change in heat ($\Delta H$) is POSITIVE: (KE $\rightarrow$ PE, motion energy from surroundings is captured and stored)

**Endothermic example:**

Place ice into water in a calorimetry cup. As the ice melts, it absorbs heat from the water; the temperature of water **DECREASES**.

**Watch Bozeman Science Chemistry Energy Changing Processes video**

https://www.youtube.com/watch?v=BgcWxlWZ84s
Energy homework

Circle your answers for questions 1, 2, 4, & 6.

1. What unit is used to express the average kinetic energy of a system?
   a) Joule  
   b) **Kelvin**  
   c) Gram  
   d) Meter

2. Overnight the air temperature decreases 30 °C. What is the decrease in Kelvin?
   a) 30 K  
   b) 243 K  
   c) 303 K  
   d) 273 K

3. Explain your answer for #2.
   °C scale interval is the same as interval for Kelvin

4. Which of the following samples has the highest average kinetic energy?
   a) 10 g H₂O @ 20°C  
   c) 30 g H₂O @ 70°C  
   b) 20 g H₂O @ 20 K  
   d) 40 g H₂O @ 200 K

5. Explain your answer for #3.
   70°C is 343 K, higher temperature than other choices

6. Which statement is correct concerning the direction of heat flow between two substances?
   a) Heat will flow from an ice cube at 260 K to water at 280 K
   b) **Heat will flow from a piece of hot glass at 800 K to a 310 K hand that touches it.**
   c) Heat will flow from dry ice at 100 K to air at 300 K.
   d) Heat will flow from a 270 K flagpole to a 310 K tongue that is double-dog-dared to stick upon it.
7. Explain your answer for #6.

Heat flows FROM hot TO cold

Perform the following conversions (show your work!)

<p>| | |</p>
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<tbody>
<tr>
<td>8. 20 kJ =</td>
<td>20000 joules</td>
</tr>
<tr>
<td>9. 100 J =</td>
<td>0.1 kJ</td>
</tr>
<tr>
<td>10. 100 K =</td>
<td>-173 °C</td>
</tr>
<tr>
<td>11. 200°C =</td>
<td>473 K</td>
</tr>
</tbody>
</table>

Describe the following situations as either Potential or Kinetic energy.

<table>
<thead>
<tr>
<th>Situation</th>
<th>PE or KE?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Energy in a compressed spring</td>
<td>12. PE</td>
</tr>
<tr>
<td>Heat given off by burning coal</td>
<td>13. KE</td>
</tr>
<tr>
<td>An earthquake in progress</td>
<td>14. KE</td>
</tr>
<tr>
<td>A rock falling from a cliff</td>
<td>15. KE</td>
</tr>
<tr>
<td>Emitted laser light</td>
<td>16. KE</td>
</tr>
<tr>
<td>A stick of dynamite (in a box)</td>
<td>17. PE</td>
</tr>
</tbody>
</table>

18. Is a temperature of – 300°C possible? Explain your answer.

No; is below absolute zero of -273°C
Energy homework

Consider the reaction A + B \rightarrow C:

If the reactants ‘A’ and ‘B’ have 40 kJ of total Ɛ stored as bonds and the product ‘C’ contains 60 kJ of Ɛ in its bonds, then (gain of 20 kJ)

19. How many kJ of PE must have been absorbed or released in this reaction?
   a) Absorbed 120 kJ       c) Absorbed 20 kJ
   b) Released 120 kJ       d) Released 20 kJ

20. Was this reaction endothermic or exothermic?

21. If this reaction were placed in a calorimeter containing water, would the water temperature increase or decrease?

22. What is the overall change in energy for this reaction, in kilojoules? Use the ‘+’ sign for endothermic change; use the ‘-‘ sign for exothermic.
   + 20 kJ
23. Why is water used as the standard for the Celsius scale? Water is also used for other standards, such as specific heat (4.18 J/g°C), viscosity (water = 1.00 centipoise), and density (water = 1.00 g/cm³ @ 3.98°C). Why use water for these standards?

Water is found all over Earth whereas other materials may be rare
Topic: **Calorimetry**

Objective: How is the energy in your food determined?

**Calorimetry examples:**

- For the reaction of $A + B \rightarrow C$:
  - If the reactants $A$ and $B$ have 80.0 kJ of total $\epsilon$ stored in their bonds and product $C$ contains 20.0 kJ of $\epsilon$ stored in its bonds:

  1. 60.0 kJ of PE must have been released and absorbed in this reaction. Started with 80.0 kJ, ended with 20.0 kJ; loss of 60.0 kJ to surroundings as KE.

  2. Since this reaction **LOST** some PE, this reaction is **EXOTHERMIC**.

  3. If this reaction were in a water calorimeter, the $T$ of the water would **INCREASE** as newly added KE increased the average water molecule speed.

The energy content of food (Calories in the U.S.; kilojoules almost everywhere else) is determined by burning food in a sealed “bomb calorimeter”. The released heat is absorbed by the calorimeter’s water, and knowing the mass of water and the water’s $\Delta T$, you can calculate the energy the body would absorb by consuming the food.
All physical and chemical changes are accompanied by an associated change in energy. This may be referred to in many ways. \( \Delta H \) is called the heat of reaction and is the change in PE of that change.

### Endothermic change:

- A positive change in PE (+\( \Delta H \)) results when KE from the surroundings is absorbed into the change, resulting in greater stored (PE) energy and lower T in the surroundings. This is an **ENDOTHERMIC** change.

Reacting one mole of nitrogen gas with two moles of oxygen gas produces two moles of nitrogen dioxide gas. This reaction requires **absorbing** 66.4 kJ of PE from the environment. The PE is stored as covalent bonds between the nitrogen and oxygen atoms.

If the change was reversed, and the nitrogen dioxide was decomposed back into nitrogen gas and oxygen gas, the reaction would release the 66.4 kJ of heat back into the environment. The +\( \Delta H \) from forming nitrogen dioxide would now be −\( \Delta H \).
Exothermic change:

- A negative change in PE (-ΔH) results when PE stored in substances undergoing change is released into the surroundings as KE, resulting in lower stored (PE) energy and higher T in the surroundings. This is an EXOTHERMIC change.

Dissolving lithium bromide crystals into water results in the breakup of LiBr\(_{(s)}\) into lithium ions and bromide ions. This physical change results in releasing 48.33 kJ of PE, which is converted into KE and released into the surroundings, increasing the T.

If the change were reversed, and the water were evaporated away so the ions rejoined to reform the LiBr\(_{(s)}\) crystal, the reaction would absorb 48.83 kJ of heat that was released into the surroundings. The ΔH, which was −ΔH when dissolved, would now be +H.
Topic: **Calorimetry**

Objective: How is the energy in your food determined?

The **Law of Conservation of Energy** states that energy cannot be created or destroyed, but only converted from one form to another. Therefore, the following must apply:

I. PE lost by \( \Delta = KE \) gained by increased (↑) T of water in calorimeter.

II. PE gained by \( \Delta = KE \) lost by decreased (↓) T of water in calorimeter.

Altogether, this is the calorimetry equation (Reference Table T):

<table>
<thead>
<tr>
<th>Equation</th>
<th>( q )</th>
<th>=</th>
<th>( m )</th>
<th>( C )</th>
<th>( \Delta T )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Units</td>
<td>Joules (J)</td>
<td>=</td>
<td>Grams (g)</td>
<td>4.18 J/g°C</td>
<td>°C</td>
</tr>
<tr>
<td>What each variable means</td>
<td>( q ) is the quantity of heat that is absorbed or released by a physical or chemical change</td>
<td>=</td>
<td>( m ) is the mass of water in the calorimeter cup that absorbs heat from the change or releases heat to the change</td>
<td>( C ) is the specific heat of water, the rate at which water gains or loses heat if energy is absorbed or removed</td>
<td>( \Delta T ) is the temperature change of the water in the calorimeter cup as a result of the physical or chemical change</td>
</tr>
</tbody>
</table>
**Topic:** How does a calorimeter work?

**Objective:** How is the energy change measured in calorimetry?

How a calorimeter works:

A sample that is undergoing a physical or chemical change is placed into the reaction chamber (A). In the example at the right, the change is releasing heat (EXOTHERMIC reaction) that it held as PE inside chemical bonds, intermolecular forces, or both. The heat changes into KE, which is released into the water (B). The water (B) molecules heat up, moving faster and striking the thermometer (C) with more force. The liquid in the thermometer bulb expands, moving upwards, showing an increase in KE. In an ENDOTHERMIC reaction, the opposite occurs as the KE in the water would be absorbed into PE of new bonds and/or intermolecular forces.

[Watch Bozeman Science Chemistry Calorimetry video](https://www.youtube.com/watch?v=BgcWxIWZ84s)
How to do Calorimetry Problems:

There are four things you can calculate using the calorimetry equation (Reference Table T). In each example, the number of sig figs in each measurement has been provided in italics to assist you in rounding.

1. Calculating the number of joules absorbed or released by a change

   Use the calorimetry equation \( q = mC \Delta T \)

   i. Example: How many joules are absorbed by 100.0 g of water if the \( T \) changed from 35.0°C to 50.0°C?

   ii. Solution:

   iii. \( q = mC \Delta T \)

   \[ q = (100.0 \text{ g})(4.18 \text{ J/g°C})(15.0°C) = 6270 = 6270 \text{ J} \]

   4 sig figs 3 sig figs 3 sig figs 3 sig figs

   iv. The temperature change from 35.0°C to 50.0°C is a 15.0°C increase, so that was used for \( \Delta T \).

2. Calculating the mass of water that is undergoing the temperature change

   Rearrange \( q = mC \Delta T \) to solve for \( m \) (divide both sides by \( C\Delta T \)):

   Then, \( m = q/C\Delta T \)

   i. Example: a sample of water is heated by 20.0°C by adding 80.0 J. What is the mass of the water?

   ii. Solution:

   iii. \( m = q/C \Delta T \)

   \[ m = (80.0 \text{ J})/(4.18 \text{ J/g°C} \times 20.0°C) = 0.9569377 = 0.957 \text{ g} \]

   3 sig figs 3 sig figs 3 sig figs 3 sig figs
3. Calculating the temperature change of water from absorption or release of energy

Rearrange \( q = mC \Delta T \) to solve for \( \Delta T \) (divide both sides by \( mC \)):

Then, \( \Delta T = \frac{q}{mC} \)

i. Example: 300.0 J is absorbed by a 50.0 g sample of water in a calorimeter. How much will the temperature change?

ii. Solution:

iii. \( \Delta T = \frac{q}{mC} \)

\[ \Delta T = \frac{300.0 \text{ J}}{50.0 \text{ g} \times 4.18 \text{ J/g}^\circ \text{C}} = 1.4354066 = 1.4^\circ \text{C} \]

4 sig figs 3 sig figs 3 sig figs 3 sig figs

4. Calculating the temperature change of water from absorption or release of energy and using it determine the initial or final temperature of the water

Rearrange \( q = mC \Delta T \) to solve for \( \Delta T \) (divide both sides by \( mC \)):

Then, \( \Delta T = \frac{q}{mC} \)

i. Example: 200.0 J is absorbed by an 80.0 g sample of water in a calorimeter at 25.000 \(^\circ\)C. What will the final T of the water be?

ii. Solution:

iii. \( \Delta T = \frac{q}{mC} \)

\[ \Delta T = \frac{200.0 \text{ J}}{80.0 \text{ g} \times 4.18 \text{ J/g}^\circ \text{C}} = 0.5980861 = 0.598^\circ \text{C} \]

4 sig figs 3 sig figs 3 sig figs

iv. But, the question asked “What will the final temperature be?” Your calculation was \( \Delta T \). You have to add to the original T of 25.000 \(^\circ\)C to the 0.598 \(^\circ\)C temperature to end with 25.6 \(^\circ\)C.
Notes page:
Calorimetry homework

Identify the following changes as indicating an endothermic or exothermic reaction.

1. Forming ammonia from its elements releases 46.2 kJ. **exo**

2. Forming iodine chloride from its elements absorbs 18.1 kJ **endo**

Label the following reactions as endothermic or exothermic. Use Reference Table I to base your answers upon.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Endo or Exo?</th>
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<th>Endo or Exo?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Formation of Al₂O₃</td>
<td>- exo</td>
<td>Dissolving NaCl in H₂O</td>
<td>+ endo</td>
</tr>
<tr>
<td>Decomposition of Al₂O₃</td>
<td>+ endo</td>
<td>Dissolving NaOH in H₂O</td>
<td>- exo</td>
</tr>
<tr>
<td>Formation of HI</td>
<td>+ endo</td>
<td>Burning CH₄ in O₂</td>
<td>- exo</td>
</tr>
<tr>
<td>Decomposition of HI</td>
<td>- exo</td>
<td>Burning C₂H₅OH in O₂</td>
<td>- exo</td>
</tr>
</tbody>
</table>

Solve the following calorimetry problems (show correct setup).

3. How many joules are required to raise the temperature of 100.0 g of water from 30.0°C to 40.0°C?

\[
\Delta T = 40.0^\circ C - 30.0^\circ C = 10.0^\circ C
\]

\[
q = mc\Delta T = (100.0 \text{ g})(4.18 \text{ J/g} \cdot ^\circ C)(10.0^\circ C) = 4180 \text{ J}
\]
4. How many joules are released when 250.0 g of water cools from 60.0°C to 20.0°C?

\[ q = mc\Delta T \]
\[ q = (250.0\text{ g})(4.18\text{ J/g°C})(40.0°C - 20.0°C) = 41800\text{ J} \]

2 s.f.

5. A sample of water is heated from 10.0°C to 15.0°C by adding 30 J of heat energy. What is the mass of the water?

\[ \frac{q}{C \Delta T} = \frac{9}{15.0°C - 10.0°C} = \frac{9}{5.0°C} = 1.85\text{ g} \]

1 s.f.

6. How many grams of water may be heated from 20.0°C to 75.0°C using 3500 J of heat energy?

\[ \frac{q}{C \Delta T} = \frac{3500\text{ J}}{(4.18\text{ J/g°C})(55.0°C - 20.0°C)} = 15.22\text{ g} = 20\text{ g} \]

2 s.f.
7. What temperature change in °C is produced when 800 J of heat energy is absorbed by 100 g of water?

\[ \Delta T = \frac{q}{mC} = \frac{800}{100 \times 4.18 \text{ J/°C}} = 1.91 \text{ °C} \]

8. What temperature change in °C is produced when 600.0 g of water releases 9.60 kJ of heat energy?

\[ \Delta T = \frac{q}{mC} = \frac{9600}{600 \times 4.18 \text{ J/°C}} = 3.83 \text{ °C} \]

9. The temperature of 50 g of water was raised to 50.0 °C by adding 1000 J of heat energy. What was the initial temperature of the water?

\[ \Delta T = \frac{q}{mC} = \frac{1000}{50 \times 4.18 \text{ J/°C}} = 4.785 \text{ °C} \]

\[ \text{Initial Temp} = 50.0 °C - 4.785 °C = 45.215 °C \]
10. What is the final temperature of water after 80.0 J is absorbed by 10.0 g of water at 25.0°C?

\[
\frac{q}{mc} = \Delta T \rightarrow \frac{80.0 \text{ J}}{10.0 \text{ g} \times 4.18 \text{ J/g°C}} = 1.913 \text{ °C} = 1.91 \text{ °C}
\]

\[\text{Initial T (25.0°C) + 1.91°C = 26.9°C Final temp}\]

11. What is the final temperature when 640 J is released by 40.0 g of water at 45.0°C?

\[
\frac{q}{mc} = \Delta T \rightarrow \frac{640 \text{ J}}{40.0 \text{ g} \times 4.18 \text{ J/g°C}} = 3.828 \text{ °C} = 3.8 \text{ °C}
\]

\[\text{Initial T (45.0°C) - 3.8°C = 41.2°C = 41°C}\]
Notes page: